CK-12 FlexBook

WCPSS Chemistry

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CHAPTER 1 Introduction to Chemistry

Chapter Outline

Do these cookies look delicious? Cooking and baking is one way that you may have experienced chemistry in your daily life. Baking powder and baking soda are both chemicals that undergo a reaction during the baking process to produce carbon dioxide. The carbon dioxide is responsible for the familiar "rising" of cookies and cakes. Of course, you know that it is very important to follow a recipe carefully in order to make the cookies turn out the way that you want. A chemist must also carefully follow specific procedures when he or she is attempting to synthesize new chemicals in the laboratory. While you read this book, you will learn about many different types of chemicals and chemical reactions. You will see how chemistry is a part of your life in many varied ways each and every day. This first chapter will introduce you to the basics of chemistry and the steps of the scientific method.

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1.1 What is Chemistry?

Lesson Objectives

- Define chemistry.
- Differentiate between the macroscopic and the microscopic as it relates to chemistry.
- Know the relationship between pure chemistry and applied chemistry.
- Identify and describe the five primary disciplines of chemistry.
- Describe some of the concerns of the modern world in which chemistry has played and will continue to play a role.

Lesson Vocabulary

- analytical chemistry
- applied chemistry
- biochemistry
- chemistry
- inorganic chemistry
- macroscopic
- matter
- microscopic
- organic chemistry
- physical chemistry
- pure chemistry

Check Your Understanding

Recalling Prior Knowledge

- What are some areas of science with which you are familiar?
- List some of the chemicals that influence your everyday life.

Chemistry is a diverse and fascinating science that impacts our lives each and every day. In this first lesson, you will learn what chemistry is and some of the ways in which chemistry influences the modern world.

A Definition of Chemistry

Science is a general term used to describe the principled, rigorous study of the natural world. Many interconnected disciplines fall under this broader concept. For example, physics is the study of motion and forces. Biology is the study of living things. Geology is the study of the Earth and the rocks and minerals of which it is comprised. In this text, you will learn about the field of chemistry. **Chemistry** *is the study of the composition of matter and the changes that matter undergoes*. **Matter** *is anything that has mass and takes up space*. Virtually everything around us is matter, including both living and nonliving things. Chemistry affects nearly everything we see and every action we take. Chemistry explains why the leaves of deciduous trees turn from green in the summer to various shades of red and yellow in the autumn (**Figure** 1.1). Chemistry explains why milk that is left in the refrigerator for too long turns sour. Chemistry explains why certain pollutants called chlorofluorocarbons have done lasting damage to the ozone layer of our planet. Chemistry is very much a central, foundational science, and a basic knowledge of chemistry is essential for students who are pursuing studies in biology, geology, environmental science, medicine, and many other subjects.

FIGURE 1.1

Chemical reactions in the leaves of deciduous trees cause them to change color from green to red, orange, or yellow before dropping to the ground.

Chemists are tasked with looking at the world in two ways, often simultaneously. The two worlds of the chemist are the macroscopic world and the microscopic world. **Macroscopic** *refers to substances and objects that can be seen, touched, and measured directly*. **Microscopic** *refers to the small particles that make up all matter*. Chemists must observe matter and do experiments in the macroscopic world and then use their observations to make generalizations and propose explanations that are microscopic in nature. For example, anyone can observe the physical change in appearance that occurs as an iron object rusts, such as a tractor that is left outside for a long enough period of time. However, a chemist looks at the rusting tractor and thinks about what is going on with the individual atoms that make up the iron and how they are changing as a result of exposure to rainwater or oxygen in the air. Throughout your study of chemistry, you will often need to switch back and forth between the macroscopic and microscopic worlds.

Alchemy

Chemistry is a subject that has its roots in the ancient tradition known as alchemy, from which it derives its name. Alchemy was a combination of philosophy and science that had both practical and mystical aspects. The goals of alchemy were varied and difficult to summarize. In many ways, the alchemists sought to achieve perfection, through such actions as the pursuit of the philosopher's stone and the elixir of life (**Figure** 1.2). The philosopher's stone, it was believed, was a substance that was capable of being used to turn base metals (such as lead) into gold. It was also believed that it could be used to achieve rejuvenation and perhaps immortality. While alchemists did not ultimately succeed in these quests, their work provided the foundation for the modern study of chemistry.

Alchemists laid the groundwork for many chemical processes, such as the refining of ores, the production of gunpowder, the manufacture of glass and ceramics, leather tanning, and the production of inks, dyes, and paints. Alchemists also made the first attempts at organizing and classifying substances so that they could better understand their reactions and be able to predict the products of their experiments. This eventually led to the modern periodic table, which you will learn about in a later chapter. Alchemy began to fully evolve into chemistry in the 17^{th} century, with a greater emphasis on rational thought and experimentation and less emphasis on spirituality and mysticism.

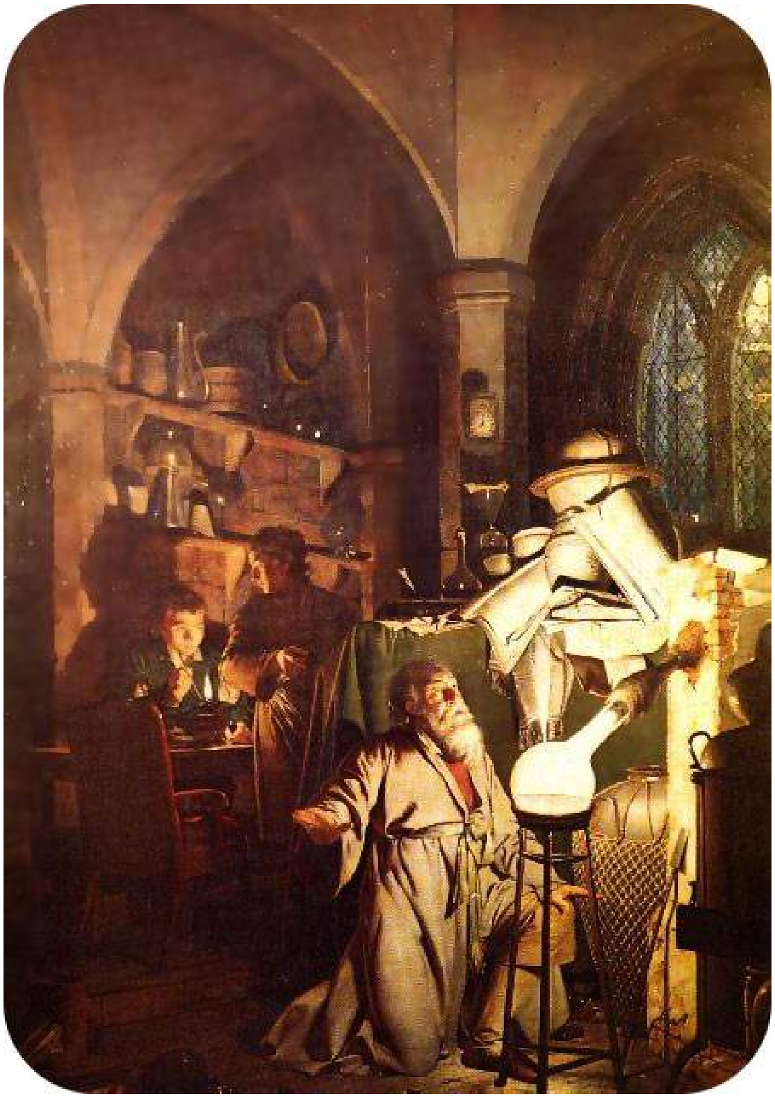

FIGURE 1.2

This painting by Joseph Wright (1771) is titled "The Alchymist, In Search of the Philosopher's Stone."

Pure and Applied Chemistry

The study of modern chemistry can be split into two categories, called pure chemistry and applied chemistry. Chemists who study pure chemistry do research primarily to advance mankind's understanding of chemistry. **Pure chemistry** *is concerned with a greater understanding of the theories behind how matter is changing in chemical reactions*. Pure chemists tend to be less concerned with direct applications of the research that they are doing. That is not to say that pure chemistry can never lead to a real-world application, but rather that a potential application is not the primary motivation for doing the research in the first place. **Applied chemistry** *is chemistry that is directed towards a specific practical goal or application*.

The line between pure chemistry and applied chemistry is not always distinct. For example, in the early 1960s, chemists at DuPont were searching for a new lightweight and strong fiber that could be used in tires. Chemist Stephanie Kwolek discovered, somewhat by accident, that a certain solution she had made displayed unique characteristics that were unlike those of other previously developed substances. The field of polymer chemistry arose in part out of her research. Polymers are very, very large molecules comprised of smaller subunits that are repeated over and over again in extremely long chains. The polymer that was discovered at DuPont was eventually given the name Kevlar. It is used not only in tires but also in bulletproof body armor because of its high strength and light weight. Polymer chemistry continues to be an active and vibrant field of chemistry, as both a pure and an applied discipline.

Chemistry Disciplines

The study of modern chemistry has many, many branches, but it can generally be broken down into five main disciplines, or areas of study. **Physical chemistry** *is the study of both macroscopic and atomic properties and phenomena in chemical systems*. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur during a reaction, or the physical structure of materials at the molecular level. **Organic chemistry** *is the study of carbon-containing chemicals*. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals. Most of the chemicals found in living organisms are based on carbon. **Inorganic chemistry** *is the study of chemicals that do not, in general, contain carbon*. Inorganic chemicals are commonly found in rocks and minerals. **Analytical chemistry** *is the study of the composition of matter, with a focus on separating, identifying, and quantifying chemical samples*. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components (**Figure** 1.3). **Biochemistry** *is the study of chemical processes that occur in living things*. Due to the importance of carbon-containing compounds in living organisms, biochemistry and organic chemistry are often closely interrelated.

In practice, chemical research is often not limited to just one of the five major disciplines. For example, a chemist might use biochemistry techniques to isolate a particular chemical from the human body, such as hemoglobin, the oxygen-carrying component of red blood cells. He or she may then proceed to analyze the hemoglobin using methods that would pertain to physical or analytical chemistry. Many chemists specialize in areas that are combinations of these primary disciplines, such as bioinorganic chemistry or physical organic chemistry.

Chemistry in the World Today

Chemistry has played and will continue to play a central role in all areas of science and technology. Before delving into the specifics of chemistry, it is worthwhile to examine a few of the frontiers that chemists are currently exploring and to review how chemistry has helped shape these fields.

Energy

Virtually all chemical reactions involve a transfer of energy, and the worldwide demand for energy continues to increase every year. The majority of our energy needs since the Industrial Revolution have been met by the burning of fossil fuels. Fossil fuels (coal, petroleum, natural gas) were formed millions of years ago from the remains of plants and animals. Fossil fuels are a nonrenewable energy source, meaning that they cannot be produced or regenerated, so the global supply will eventually be exhausted. Chemists are actively involved in finding ways to conserve energy and to utilize alternative energy sources.

Nuclear energy is obtained by splitting large atoms such as uranium into smaller atoms in a process called nuclear fission. Nuclear fission releases tremendous amounts of energy, much more than burning an equivalent amount of fossil fuels. While nuclear energy is widely used to supplement the energy demands of modern society, it comes with risks and drawbacks. For example, nuclear power generates waste that is dangerously radioactive and difficult

FIGURE 1.3

Nuclear magnetic resonance (NMR) is an analytical technique that uses very large magnets in order to learn about the specific makeup of complex chemicals. Pictured is a 900 MHz NMR spectrometer located in Washington, USA.

to store safely. Chemists are currently searching for better ways to store nuclear waste. Nuclear fusion, the process that occurs in the Sun and other stars, is an alternative that does not generate the same type of waste as nuclear fission. Scientists are actively trying to make nuclear fusion a practical source of energy.

Biofuels are an alternative energy source that has gained a great deal of attention in recent years (**Figure** 1.4). Various types of biofuels are obtained from vegetable oils, animal fats, and the decomposition of corn or other crops. Although the use of biofuels lessens our dependence on fossil fuels, it still involves the burning of carbon-based chemicals, which contributes to climate change.

Finding alternative energy sources that are renewable and do not harm the environment is a primary goal for scientists working in this field. Renewable energy sources are those that either will not run out or that can be used over and over again. Chemists and other scientists are actively involved in the search for these energy sources and for ways to make their use widespread and practical. Solar energy involves the direct transformation of the energy from the Sun into electricity by the use of devices called photovoltaic cells. Energy from wind can be obtained from the use of giant windmills. Wind power is an energy source that has been on the rise in recent years, with an increasing number of "wind farms" being built both on land and offshore. Geothermal energy is energy that is generated and stored deep within the Earth. This energy source can also be tapped and used to generate electricity.

FIGURE 1.4
This city bus is powered by biofuel derived from soybean oil.

Medicine

Advances in medicine over the last century have dramatically increased lifespans and improved our quality of life. Chemists are at the forefront of many of these medical breakthroughs and improvements. Medical researchers work with doctors to gain a better understanding of biological processes that occur in the human body. Today, thousands of prescription and over-the-counter drugs are available to treat everything from a simple headache to major infections, high blood pressure, depression, and a host of other ailments. Drugs work because of their interactions with other chemicals in the body. Chemists are constantly working to improve existing drugs, to design new and better drugs, and to further their understanding of the chemical mechanisms by which drugs perform their functions.

Chemists are also involved in designing materials that are used to repair or replace body parts. Artificial hips and knees can be inserted into a person's body so that he or she can walk without pain (**Figure** 1.5). Burn victims benefit from replacement skin. Diseased arteries can be replaced with plastic tubing that must be able to withstand the harsh environment found inside a human body without degrading or malfunctioning.

FIGURE 1.5
A prosthetic hip made of titanium, with a ceramic head and a polyethylene cup

Biotechnology is an emerging field that involves the manipulation of DNA, the chemicals that store the genetic information that gets passed on from one generation to the next. Genes are the pieces of DNA that contain this genetic information. Biochemists can transfer genes from one organism to another in order to take advantage of useful traits. For example, genes from a certain bacteria produce a chemical that confers resistance to certain crop-eating pests. Inserting these genes into the genetic material of corn plants results in genetically modified corn that is able to resist the pest without having to be sprayed with a chemical pesticide. Bacteria can also be used as "chemical factories" to produce useful drugs such as insulin, which is used to treat some forms of diabetes.

Materials

Chemists are also involved in the design and production of new materials. Some of the materials that chemists have helped discover or develop in recent years include polymers, ceramics, adhesives, coatings, and liquid crystals. Liquid crystals are used in electronic displays, such as those found in watches and calculators. Silicon-based computer chips have revolutionized modern society, and chemists have played a key role in their design and continued development. Many chemists are currently working in the field of superconductivity. Superconductors are materials that can conduct electricity with 100% efficiency. Unlike conventional conducting materials, like copper cables, no energy is lost during electrical transmission through superconductors. Unfortunately, most known superconductors only exhibit this property at extremely low temperatures. A current challenge in this field is to design materials that can act as superconductors at normal temperatures.

The Environment

There is a great deal of concern about the deleterious effects that humankind has had and continues to have on the natural environment. The burning of fossil fuels pumps millions of tons of carbon dioxide into the atmosphere. Carbon dioxide is a greenhouse gas, which means that it traps the heat from the Sun and prevents it from being released back out into space in the form of infrared radiation. As a result, the temperature of the Earth is slowly climbing. Chemists, along with environmental scientists, are searching for ways to slow or reverse the effects of global warming. One way, as mentioned previously, is to develop alternative energy sources such as wind power that do not emit carbon dioxide.

Many other chemicals that were once commonly used were later found to be harmful to the environment, to human health, or both. The element lead was once a common additive to gasoline and to paint. Plumbing pipes were once typically made of lead. Only since the 1970s has the danger of lead become apparent. Too much exposure to lead can cause brain damage, and small children (who often chewed on objects painted with lead-based paint) are particularly susceptible to this effect. The use of lead in gasoline, paint, and plumbing pipes is now banned. Mercury is another toxic element whose use was once far more widespread than it is today, especially in many industrial processes. Chemists continue to look for threats to our health and the environment and to search for substitutes so that harmful chemicals can be replaced with equally effective but less harmful alternatives.

Agriculture

The world's population has increased dramatically over the last century. It is becoming challenging to feed this population, since the amount of land available to grow food is limited. Even worse, the amount of usable land is declining as some of it is being converted to living areas. Chemists are involved in efforts to make the land that is used for agriculture as efficient and productive as possible. Some factors that decrease productivity include poor soil quality, lack of sufficient water, crop-eating pests, weeds, and diseases. Chemists are involved in the study of soils, including how to improve existing soils and to develop techniques so that soils do not lose their nutrients. Fertilizers are used to provide crop plants with vital nutrients, but the excessive use of fertilizers is harmful to the environment, particularly when the fertilizer runs off the land and into the water supply (**Figure** 1.6). Herbicides and insecticides have traditionally been applied to crops in order to reduce losses from weeds and insects. However, pesticides contain toxic chemicals that are potentially harmful to other organisms, including humans. Chemists are involved in developing alternatives to traditional pesticides so that the use of these harmful chemicals can be reduced. Biotechnology was mentioned previously as a way to increase the resistance of crops to pests. Pests can also be attacked by using natural chemicals to disrupt their mating and reproductive mechanisms so that the overall number of pests declines.

FIGURE 1.6

Runoff from excessive fertilizer usage can cause an algal bloom due to the large increase in nutrient levels that causes algae populations to explode. Algal blooms are harmful to the natural processes of the aquatic ecosystem.

Lesson Summary

- Chemistry is the study of matter and its changes. Chemistry is closely interrelated with other scientific fields, such as biology and geology.
- Chemists examine and explain their observations on both macroscopic and microscopic levels.
- Pure chemistry is devoted to the acquisition of chemical knowledge, while applied chemistry is focused on a specific goal.
- The five main disciplines of chemistry are physical, organic, inorganic, analytical, and biochemistry.
- Chemistry has modern applications in the fields of energy, medicine, materials science, the environment, and agriculture.

Lesson Review Questions

Reviewing Concepts

1. Which of the following scientific inquiries involve chemistry? For the others, what field of science is being described?
   a. how long it takes for an object to fall from a rooftop to the ground
   b. how fast a flammable liquid burns
   c. the amount of pollutants in a lake

   d. the structure of a plant cell

2. The same event is described in two ways below. Which description focuses on the macroscopic level and which focuses on the microscopic level?
   a. Chemical bonds between oxygen atoms and hydrogen atoms are broken.
   b. When an electric current passes through water, a gas is produced.
3. State whether each of the chemists below is practicing pure chemistry or applied chemistry. Explain.
   a. He is attempting to formulate a new adhesive that works well in extremely cold temperatures.
   b. She is studying reactions of the element sodium with different gases.
4. Which of the five main chemistry disciplines is most closely related to each of the actions described below?
   a. Studying the digestion of certain types of foods
   b. Investigating the iron content of different samples of ore
   c. Measuring how much energy is released in a series of reactions
   d. Attempting to produce a new carbon-based chemical
   e. Measuring the amount of mercury contamination in a sample of river water
5. What is the difference between renewable and nonrenewable energy?
6. Identify the following energy sources as renewable or nonrenewable.
   a. fossil fuels
   b. solar energy
   c. nuclear power
   d. wind
   e. geothermal energy
7. Name a process that is involved in biotechnology.
8. What are greenhouse gases and why is their production considered to be an environmental problem?
9. What are two alternatives to the traditional use of chemically-based pesticides in agriculture?

Further Reading / Supplemental Links

- Branches of Chemistry: Overview of the Branches of Chemistry, (http://chemistry.about.com/od/branchesofchemistry/a/bra

Points to Consider

Chemists, like all scientists, rely on a rigorous and systematic approach to studying the world called the scientific method.

- What are the steps of the scientific method?
- How do chemists and other scientists communicate their findings to the public?

1.2 The Scientific Method

Lesson Objectives

- Describe how the Renaissance period in history changed the approach to science.
- Identify the steps of the scientific method.
- Differentiate between the independent variable and the dependent variable in an experiment.
- Describe how scientists work in research groups and communicate their results.

Lesson Vocabulary

- control group
- dependent variable
- hypothesis
- independent variable
- scientific law
- scientific method
- theory

Check Your Understanding

Recalling Prior Knowledge

- How did the ancient practice of alchemy differ from the modern science of chemistry?
- What are the properties of a well-designed scientific experiment?

Advancements in science are not always planned. Many groundbreaking discoveries in the history of science were made almost completely by accident. However, science follows a rigorous, time-tested process called the scientific method. By applying the scientific method, scientists can be ready for whatever results their research gives them, and many new and important discoveries can be made.

Experimental Science

It may surprise you to know that science is, relatively speaking, a very young subject. The Renaissance, which occurred roughly between the 14^{th} and 17^{th} centuries, was a time of great upheaval in Europe. The Middle Ages were dominated by a dogmatic approach to knowledge. The Renaissance ushered in new ways of thinking and great advancements in art, music, and culture. There also came a new emphasis on obtaining knowledge through the processes of observation and experimentation, as opposed to a reliance on older ideas that were based on a less rigorous scientific process. Two individuals who personified the Renaissance approach to science are Leonardo da Vinci and Nicolaus Copernicus. Many credit da Vinci with being the father of modern science. He is renowned for his observations and drawings of human anatomy (**Figure** 1.7). He also set up controlled experiments to investigate water flow, medical dissection, and the systematic study of movement and aerodynamics. Copernicus was an astronomer who suggested, based on observations, a heliocentric model of the solar system, meaning that the Earth and other planets traveled around the Sun. The prevailing doctrine had been that the Earth was the center of the universe and that all celestial bodies must travel around it.

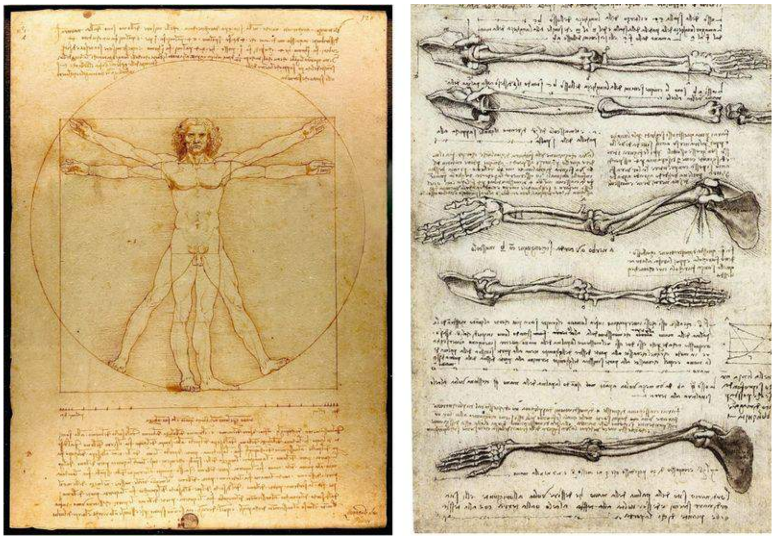

FIGURE 1.7

Leonardo da Vinci was a Renaissance artist and scientist who made great advancements in the field of anatomy. On the left is the famous Vitruvian Man drawing. On the right is a document showing the anatomy of the human arm.

More than any one particular piece of knowledge, the Renaissance represented a change in the way human beings thought about the world around them. Emphasis shifted from a philosophical explanation of the world's natural processes to one that was instead based on observation and experimentation. The process of scientific discovery that we now call the scientific method originated during the Renaissance. This new method led to great advancements in the fields of astronomy, physics, anatomy, biology, and chemistry.

The Scientific Method

The **scientific method** *is a systematic and logical approach to the acquisition of knowledge* (see **Figure** 1.8). It uses specific techniques to observe and investigate natural phenomena. The scientific method differs from other methods of inquiry in that scientists seek to let reality speak for itself rather than approaching inquiry with any preconceived biases towards one explanation or another. The steps of the scientific method are: (1) formulating a question, (2) making a hypothesis, (3) testing the hypothesis through experimentation, (4) analyzing results, and (5) developing theories or revising the hypothesis.

Formulating a Question

The formulation of a question represents the first step in scientific inquiry. The type of question can be very specific or more general. Questions are often based on observations. For example, suppose that you observe that a lamp does not light up when you turn on the switch. The natural question is to ask why the lamp does not work. This in turn would lead to an investigation of the possible reasons why the lamp is not working. However, the process begins with a question that needs to be answered.

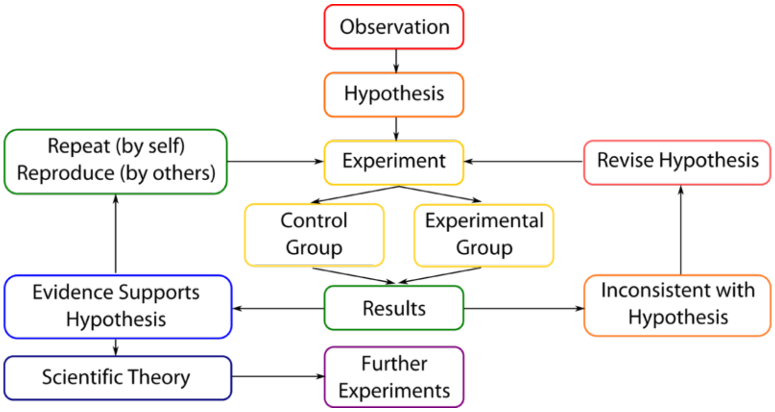

FIGURE 1.8

The steps of the scientific method are easily represented using a flowchart.

Making a Hypothesis

A **hypothesis** *is a testable conjecture that proposes an answer to the question*. A hypothesis for the nonfunctioning lamp could be that the light bulb is burned out. In order to be a valid hypothesis, it must be possible to test it.

Testing the Hypothesis and Analyzing Results

The light bulb could be easily removed and replaced with another one. If the new light bulb works, then the original hypothesis is confirmed. However, if the lamp still does not work even when the light bulb has been replaced, the hypothesis is now in question. Further experiments with more light bulbs could be done. Eventually, if the lamp continues not to work, the original hypothesis must be discarded. At this point, the scientist may decide to either make a slight change in the hypothesis or abandon the experiment entirely. Perhaps the change could be that there is a problem with the switch for the lamp that is preventing it from working. A new set of experiments could be performed to determine if the original switch was responsible for the lamp not working.

The lamp example oversimplifies the usual process in a real scientific experiment. In most experiments, multiple experimental groups are created in order to test the effect of one variable on another. For example, suppose that you wanted to know how elevated air temperature affects the germination rate of a certain type of bean seed. You would set up multiple pots planted in identical fashion with the same bean seeds. You would establish experimental constants such as the amount of water and the amount of light that each planting would receive. The **independent variable** *in an experiment is the variable that is changed during an experiment*. In this case, the temperature is the independent variable. You would figure out a way to place each of the experimental groups at a different elevated temperature. One planting would be left at room temperature. This is called the control group. A **control group** *is a group that experiences the same conditions as the experimental groups with the exception of the variable that is being tested*. The bean plantings would be observed at regular intervals, perhaps every 12 hours. Data would be collected as to how many bean seeds germinated in each planting and when germination took place. The **dependent variable** *in an experiment is the variable that is observed during an experiment*. In this experiment, germination of the bean seeds is the dependent variable.

In order to be confident about the results of an experiment, it needs to be repeatable. Professional scientists normally repeat experiments many times before they make their findings known to other scientists or to the public.

Developing Theories

If many experiments are performed and each and every one supports the hypothesis, it may be that the hypothesis can be accepted. A **theory** *is an explanation that has been repeatedly tested and confirmed by multiple researchers and* *has gained wide acceptance*. The term theory is one that is often misunderstood by people in the general population who are unfamiliar with the scientific method. Though a scientist may state that a theory cannot be proven correct, he or she is certainly not saying that it is in doubt. A scientific theory has generally been confirmed over and over and over again, although small details within a theory may be adjusted from time to time as new discoveries are made. In a later chapter you will learn about atomic theory and how it has changed over the last few centuries. While the existence of atoms has not been in question for some time, new details about how the atom functions have resulted in many updates and alterations of the essential theory.

Scientific Laws

A **scientific law** *is a statement that summarizes the results of many observations and experiments and to which there are no known exceptions*. A familiar scientific law is the law of gravity. The force of attraction between two objects is dependent in a very specific way on the masses of those objects and on the distance between them. The law of gravity has been tested and proven countless times. Scientific laws often have a mathematical basis, though that is not necessarily a requirement.

Scientific Communication

Today's scientists rarely work alone. Rather, most scientists collaborate with one another in various settings as part of a group effort. The majority of research scientists work either for a company, such as the DuPont Chemical Company, or for one of many universities, such as the California Institute of Technology. Working as part of a group has many advantages. Most scientific problems are so complex and time-consuming that one person could not hope to address all of the issues by himself or herself. Instead, different members of a research group are each tasked with a particular small aspect of a larger research problem. Collaboration between members of the group is frequent. This occurs informally in the laboratory on an everyday basis. Research groups also typically have regular meetings where one or more members of the group may give a presentation to the others on the status of the research that they are doing. Progress normally occurs in small steps rather than grand, sweeping discoveries, and that progress is helped along by the teamwork that comes from working as part of a group.

Modern scientific research is usually expensive. Lab equipment, chemicals, research space, and the upkeep of technical instrumentation all cost money. Research groups need to raise money in order to continue their work. For research done at universities, much of that money comes from government sources, such as the National Science Foundation or the National Institute of Health. Private companies can fund their own research, but they may seek outside funding as well. Scientists write grant proposals explaining the goals and projected costs of their research, and funding agencies make decisions about which research projects they would like to fund. The long-term viability of most research labs depends on the ability to get and maintain funding.

Communicating Results

Suppose that your research is a success. What now? Scientists communicate their results to one another and to the public at large in several ways. One is to publish their research findings in publications called scientific journals. There are many hundreds of scientific journals covering every imaginable field of science. A few of the major chemistry journals include the "Journal of the American Chemical Society," the "Journal of Physical Chemistry," and "Inorganic Chemistry." Some journals have a very narrow scope while others publish articles from many different scientific disciplines and appeal to a wider audience. Examples of the latter include the journals "Science" and "Nature." Journal articles are often very complex and detailed. They must be accurate, since the research field as a whole uses these journal articles as a way to make scientific progress. Therefore, journal articles are only published after having been extensively reviewed by other professional scientists in the same field. Reviewers have the power to make suggestions about the research or possibly question the validity of the author's conclusions. Only when the reviewers are satisfied that the research is correct will the journal publish the article. This peer-review process allows scientists to trust the research findings that they read about in journals.

Scientists also communicate with one another by presenting their findings at international conferences. Some scientists are chosen to give lectures at conferences, typically about research that has already been published. Many other scientists at the same conference will present their work at poster sessions (see **Figure** 1.9). These poster sessions are more informal and may often represent research that is still in progress.

FIGURE 1.9
Scientists at a poster session

Lesson Summary

- The Renaissance, which took place in Europe between the 14^{th} and 17^{th} centuries, was a time of great change in thought and culture. Along with many other changes, science began to follow a more systematic approach based on careful observation and experimentation.
- The steps of the scientific method are to formulate a question, make a hypothesis, perform experiments, analyze the results, and develop theories.
- Professional scientists collaborate with one another by working in research groups and present their results in scientific journals and at international conferences.

Lesson Review Questions

Reviewing Concepts

1. What were some of the scientific contributions of da Vinci and Copernicus?
2. What is a key requirement for a hypothesis to be a valid part of the scientific method?
3. Which of the following is not a part of the scientific method?
    a. experiment
    b. hypothesis
    c. chance
    d. theory
4. What is the next step if a particular experiment is inconsistent with the hypothesis?

5. Why do scientists repeat experiments?
6. What is the difference between the independent variable and the dependent variable in an experiment?
7. What makes a theory different from a hypothesis?
8. What is a scientific journal and what is meant by the term “peer review”?

Problems

9. A student decides to test the effect of added salt on the boiling point of water. He collects the following data (**Table** 1.1):

TABLE 1.1: **short caption**

Amount of added salt (g)	Boiling Temperature (°C)
0	100
20	101
40	102
100	105

a. What is the independent variable?
b. What is the dependent variable?
c. What is the purpose of the first experiment, when no salt was added to the water?
d. What needs to be kept constant during the experiment in order for the results to be valid?

Further Reading / Supplemental Links

- Scientific Method History, (http://scientificmethod.com/sm5_smhistory.html
- Steps of the Scientific Method: An Explanation, (http://scientificmethod.com/sm4_stepsofsm.html

Points to Consider

Chemistry is the study of matter and the changes that it undergoes. Before beginning to examine those changes, it is important to be able to classify matter according to certain common characteristics and be able to recognize when a chemical change is occurring.

- What properties of matter allow it to be classified?
- What clues allow a chemist to recognize a chemical reaction?

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4. Courtesy of the US Department of Energy. http://commons.wikimedia.org/wiki/File:Soybeanbus.jpg . Public Domain
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7. Leonardo da Vinci (photographer for image on left: Luc Viatour). (”left’’) http://commons.wikimedia.org/wiki/File:Da_Vinci_Vitruve_Luc_Viatour.jpg; (”right”) http://commons.wikimedia.org/wiki/File:Studies_of_the_Arm_showing_the_Movements_made_by_the_Biceps.jpg . Public Domain
8. CK-12 Foundation - Joy Sheng. . CC-BY-NC-SA 3.0
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CHAPTER 2 Measurements

Chapter Outline

Measurements, measurements, measurements! They affect your life in many ways each and every day. When cooking, it is very important to carefully measure each ingredient so that your recipe turns out exactly the way you intend. We make time measurements frequently, from cooking times, to estimating travel times and deciding how much time will be required for various tasks. The builders of your house or apartment hopefully made good measurements of the various wooden, plastic, and metal materials that went into its construction. A chemist also must make accurate measurements in the laboratory when conducting his or her experiments. The image above shows some of the common glassware that a chemist frequently uses in measuring volumes of liquids. In this chapter you will learn about the critical skills of making, reporting, and performing calculations with measurements.

Opening image courtesy of Tweenk. CC–BY 3.0.

2.1 The International System of Units

Lesson Objectives

- Identify the seven base units of the International System of Units.
- Know the commonly used metric prefixes.
- Convert between three temperature scales.
- Understand volume and energy as combinations of SI Units.
- Distinguish between mass and weight.

Lesson Vocabulary

- energy
- International System of Units (SI)
- joule
- kinetic energy
- liter
- measurement
- scientific notation
- temperature
- weight

Check Your Understanding

Recalling Prior Knowledge

- Why is the metric system easier to use than the English system of units?
- How is scientific notation used to represent very large or very small numbers?
- What units are used to measure length, mass, and volume in the metric system?

The temperature outside is 52 degrees Fahrenheit. Your height is 67 inches and your weight is 145 pounds. All are examples of measurements. A **measurement** *is a quantity that includes both a number and a unit*. If someone were to describe the height of a building as 85, that would be meaningless. 85 meters? 85 feet? Without a unit, a measurement does not convey enough information to be useful. In this lesson, we will begin an exploration of the units that are typically used in chemistry.

SI Base Units

All measurements depend on the use of units that are well known and understood. The English system of measurement units (inches, feet, ounces, etc.) is not used in science because of the difficulty in converting from one unit to another. The metric system is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. The **International System of Units** *is a system of measurement based on the metric system*. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units (see **Table** 2.1).

TABLE 2.1: **SI Base Units of Measurement**

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first five units are frequently encountered in chemistry. The amount of a substance, the mole, will be discussed in detail in a later chapter. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

You can learn more about base units at www.nist.gov/pml/wmd/metric/si-units.cfm .

Metric Prefixes and Scientific Notation

As stated earlier, conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. Prefixes are used to distinguish between units of different size. **Table** 2.2 lists the most common metric prefixes and their relationship to the central unit, which has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

TABLE 2.2: **SI Prefixes**

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
giga	G	10^9	1,000,000,000	1 gigameter (Gm) = 10^9 m
mega	M	10^6	1,000,000	1 megameter (Mm) = 10^6 m
kilo	k	10^3	1000	1 kilometer (km) = 1000 m
hecto	h	10^2	100	1 hectometer (hm) = 100 m
deka	da	10^1	10	1 dekameter (dam) = 10 m
		$\mathbf{10^0}$	**1**	**1 meter (m)**
deci	d	10^{-1}	1/10	1 decimeter (dm) = 0.1 m

TABLE 2.2: (continued)

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
centi	c	10^{-2}	1/100	1 centimeter (cm) = 0.01 m
milli	m	10^{-3}	1/1000	1 millimeter (mm) = 0.001 m
micro	μ	10^{-6}	1/1,000,000	1 micrometer (μm) = 10^{-6} m
nano	n	10^{-9}	1/1,000,000,000	1 nanometer (nm) = 10^{-9} m
pico	p	10^{-12}	1/1,000,000,000,000	1 picometer (pm) = 10^{-12} m

There are more prefixes, although some of them rarely used. Have you ever heard of a zeptometer? You can learn more about metric prefixes at www.nist.gov/pml/wmd/metric/prefixes.cfm .

Table 2.2 introduces a very useful tool for working with numbers that are either very large or very small. **Scientific notation** *is a way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power*. As an example, the distance from Earth to the Sun is about 150,000,000,000 meters –a very large distance indeed. In scientific notation, the distance is written as 1.5×10^{11} m. The coefficient is 1.5 and must be a number greater than or equal to 1 and less than 10. The power of 10, or exponent, is 11. See **Figure** 2.1 for two more examples of scientific notation. Scientific notation is sometimes referred to as exponential notation.

FIGURE 2.1

The Sun is very large and very distant, so solar data is better expressed in scientific notation. The mass of the Sun is 2.0×10^{30} kg and its diameter is 1.4×10^{9} m.

Very small numbers can also be expressed using scientific notation. The mass of an electron in decimal notation is 0.000000000000000000000000000911 grams. In scientific notation, the mass is expressed as 9.11×10^{-28} g. Notice that the value of the exponent is chosen so that the coefficient is between 1 and 10.

Typical Units in Chemistry

Length and Volume

The SI basic unit of length, or linear measure, is the meter (m). All measurements of length may be made in meters, though the prefixes listed in **Table** 2.2 will often be more convenient. The width of a room may be expressed as about 5 meters (m), whereas a large distance such as the distance between New York City and Chicago is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 20 micrometers (μm).

Volume is the amount of space occupied by a sample of matter (see examples in **Figure** 2.2). The volume of a regular object can be calculated by multiplying its length by its width by its height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter (m^3), which is the volume occupied by a cube that measures 1 m on each side. This very large volume is not very convenient for typical use in a chemistry laboratory. A **liter (L)** *is the volume of a cube that measures 10 cm (1 dm) on each side*. A liter is thus equal to both 1000 cm^3 (10 cm × 10 cm × 10 cm) and to 1 dm^3. A smaller unit of volume that is commonly used is the milliliter (mL). A milliliter is the volume of a cube that measures 1 cm on each side. Therefore, a milliliter is equal to a cubic centimeter (cm^3). There are 1000 mL in 1 L, which is the same as saying that there are 1000 cm^3 in 1 dm^3.

FIGURE 2.2

(A) A typical water bottle is 1 liter in volume. (B) These dice measure 1 cm on each side, so each die has a volume of 1 cm^3 or 1 mL. (C) Volume in the laboratory is often measured with graduated cylinders, which come in a variety of sizes.

You can watch a video about measuring volume using graduated cylinders at http://www.benchfly.com/video/153/how-to-use-graduated-cylinders/ .

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The mass of an object is made in comparison to the standard mass of 1 kilogram. The kilogram was originally defined as the mass of 1 L of liquid water at 4°C (the volume of a liquid changes slightly with temperature). In the laboratory, mass is measured with a balance (**Figure** 2.3), which must be calibrated with a standard mass so that its measurements are accurate.

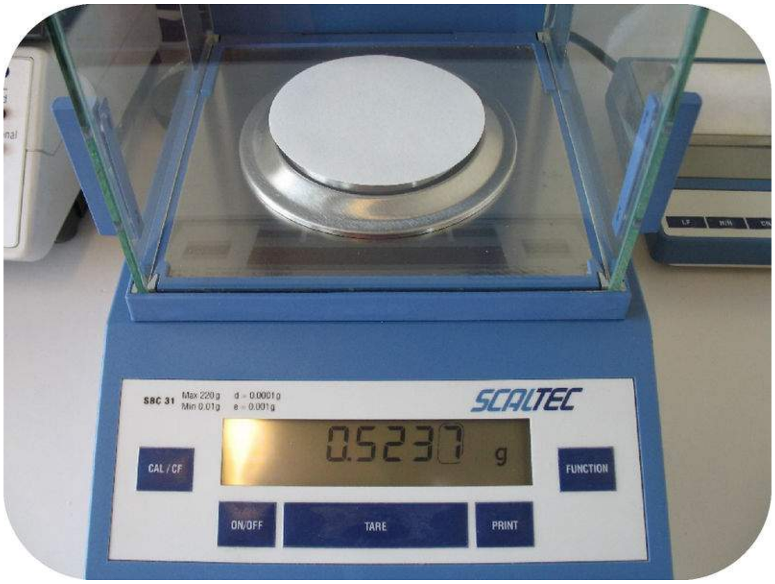


FIGURE 2.3

An analytical balance in the laboratory takes very sensitive measurements of mass, usually in grams.

You can watch a short video about using an analytical balance at http://www.benchfly.com/video/54/how-to-weigh-small-amounts/ .

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are 1000 g in 1 kg. A milligram is 1/1000th of a gram, so there are 1000 mg in 1 g.

Mass is often confused with the term weight. **Weight** *is a measure of force that is equal to the gravitational pull on an object*. The weight of an object is dependent on its location. On the moon, the force due to gravity is about one sixth that of the gravitational force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale by reading the distance that a certain object pulls down and stretches a spring.

Temperature and Energy

Touch the top of the stove after it has been on and it feels hot. Hold an ice cube in your hand and it feels cold. Why? The particles of matter in a hot object are moving much faster than the particles of matter in a cold object. An object's **kinetic energy** *is the energy due to motion*. The particles of matter that make up the hot stove have a greater amount of kinetic energy than those in the ice cube (see **Figure** 2.4). **Temperature** *is a measure of the average kinetic energy of the particles in matter*. In everyday usage, temperature is how hot or cold an object is. Temperature determines the direction of heat transfer. When two objects at different temperatures are brought into contact with one another, heat flows from the object at the higher temperature to the object at the lower temperature. This occurs until their temperatures are the same.

FIGURE 2.4

The glowing charcoal on the left is composed of particles with a high level of kinetic energy, while the snow and ice on the right are made of particles with much less kinetic energy.

Temperature can be measured with several different scales. The Fahrenheit scale is typically not used for scientific purposes. The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at 0°C and 100°C, respectively. The distance between those two points is divided into 100 equal intervals, each of which is referred to as one degree.

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824-1907). It is based on molecular motion, with the temperature of 0 K, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is 273.15 K, while the boiling point is 373.15 K. As can be seen by the 100 kelvin difference between the two, a change of one degree on the Celsius scale is equivalent to the change of one kelvin on the Kelvin scale. Converting from one scale to another is easy, as you simply add or subtract 273 (**Figure** 2.5).

$$^{\circ}\text{C} = \text{K} - 273$$
$$\text{K} = {}^{\circ}\text{C} + 273$$

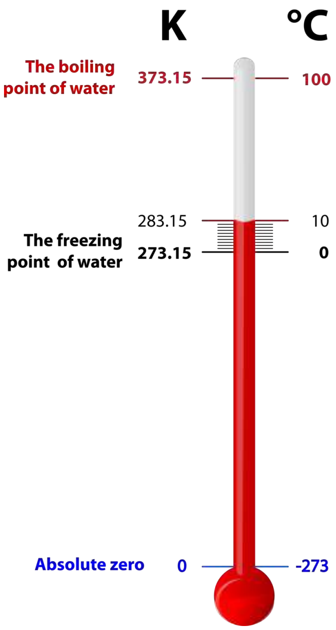


FIGURE 2.5

A comparison of the Kelvin (left) and Celsius (right) temperature scales. The two scales differ from one another by 273.15 degrees.

Energy *is defined as the capacity to do work or to produce heat*. As discussed previously, kinetic energy is one type of energy and is associated with motion. Another frequently encountered form of energy is potential energy, which is a type of energy that is stored in matter. The **joule (J)** *is the SI unit of energy* and is named after English physicist James Prescott Joule (1818-1889). In terms of SI base units, a joule is equal to a kilogram times a meter squared divided by a second squared ($kg•m^2/s^2$). A common non-SI unit of energy that is often used is the calorie (cal), which is equal to 4.184 J.

Lesson Summary

- Measurements are critical to any field of science and must consist of a quantity and an appropriate unit. The International System of Units consists of seven base units.
- The metric system utilizes prefixes and powers of 10 to make conversions between units easy.
- Length (m), mass (kg), temperature (K), time (s), and amount (mol) are the base units that are most frequently used in chemistry. Quantities such as volume and energy can be derived from combinations of the base units.

Lesson Review Questions

Reviewing Concepts

1. Give the SI base unit of measurement for each of the following quantities.
    a. mass
    b. length
    c. time
    d. temperature
2. Convert the following numbers into scientific notation.

   a. 85,000,000
   b. 0.00019

3. Put the following into decimal notation.
   a. 8.72×10^{-8}
   b. 3×10^{4}
4. Place the following units of mass in order from smallest to largest: g, kg, μg, g, pg, Mg, ng, cg, dg.
5. Explain what is wrong with the following statement: "This rock weighs 8 kilograms."
6. What is absolute zero on the Celsius temperature scale?

Problems

7. Calculate the volume in mL of a cube that is 2.20 cm on each side.
8. A rectangular solid has a volume of 80 cm^3. Its length is 2.0 cm and its width is 8.0 cm. What is the height of the solid?
9. Convert the following Celsius temperatures to Kelvin.
   a. 36°C
   b. $-104°$
10. Convert the following Kelvin temperatures to degrees Celsius.
    a. 188 K
    b. 631 K
11. Temperature in degrees Fahrenheit can be converted to Celsius by first subtracting 32, then dividing by 1.8. What is the Celsius temperature outside on a warm 88°F day?
12. Two samples of water are at different temperatures. A 2 L sample is at 40°C, while a 1 L sample is at 70°C.
13. The particles of which sample have a larger average kinetic energy?
14. The water samples are mixed. Assuming no heat loss, what will be the temperature of the 3 L of water?

Further Reading / Supplemental Links

- E.A. Mechtly, *International System of Units: Fundamental Constants and Conversion Factors*. Stipes Pub Llc, 1977.
- SI Metric System: (http://www.simetric.co.uk/sibasis.htm
- You can do an online metric system crossword puzzle at http://education.jlab.org/sciencecrossword/metricsystem_01.html .
- Take a museum tour of weights and measurements at museum.nist.gov/exhibits/ex1/index.html .
- You can view a comparison of the sizes of viruses, DNA, and biological molecules, along with information about DNA-based computers, at http://publications.nigms.nih.gov/chemhealth/cool.htm .
- Time is standardized by an atomic clock. You can view a short video about the Amazing Atomic Clock at http://video.pbs.org/video/2167682634 .
- Test your metric skills by playing Metric System Hangman. There are two Hangman games.
  - The first is at http://education.jlab.org/vocabhangman/metric_system_01/8.html .
  - The second is at http://education.jlab.org/vocabhangman/metric_system_02/8.html .

Points to Consider

Conversions between units of the metric system are made easy because they are related by powers of ten and because the prefixes are consistent across various types of measurement (length, volume, mass, etc.).

- What is the mass in grams of a 2.50 kg book?
- What is the length in cm of a field that is 0.65 km?

2.2 Unit Conversions

Lesson Objectives

- Identify and use conversion factors.
- Use the method of dimensional analysis to convert between units.
- Understand density as a physical property of matter.
- Perform calculations with derived units, including density.

Lesson Vocabulary

- conversion factor
- density
- derived unit
- dimensional analysis

Check Your Understanding

Recalling Prior Knowledge

- Why are units required when reporting the results of a measured quantity?
- When a quantity with a large unit (such as km) is changed into a quantity with a smaller unit (such as cm), will the numerical value of the quantity increase or decrease?

When traveling in another country, you may be faced with a unit problem. For example, if you are driving, you may encounter a sign saying that the next town is 30 km away. If your car's odometer measures distances in miles, how far will you need to go to get to that town? In this lesson, you will learn to solve this and other unit-conversion problems with a technique called dimensional analysis.

Conversion Factors

Many quantities can be expressed in several different ways. For example, the English system measurement of 4 cups is also equal to 2 pints, 1 quart, and 1/4 of a gallon.

$$4 \text{ cups} = 2 \text{ pints} = 1 \text{ quart} = 0.25 \text{ gallon}$$

Notice that the numerical component of each quantity is different, while the actual amount of material that it represents is the same. That is because the units are different. We can establish the same set of equalities for the metric system:

$$1 \text{ meter} = 10 \text{ decimeters} = 100 \text{ centimeters} = 1000 \text{ millimeters}$$

The metric system's use of powers of 10 for all conversions makes this quite simple.

Whenever two quantities are equal, a ratio can be written that is numerically equal to 1. Using the metric examples above:

$$\frac{1 \text{ m}}{100 \text{ cm}} = \frac{100 \text{ cm}}{100 \text{ cm}} = \frac{1 \text{ m}}{1 \text{ m}} = 1$$

The fraction 1 m/100 cm is called a conversion factor. A **conversion factor** *is a ratio of equivalent measurements*. Because both 1 m and 100 cm represent the exact same length, the value of the conversion factor is 1. The conversion factor is read as "1 meter per 100 centimeters." Other conversion factors from the cup measurement example can be:

$$\frac{4 \text{ cups}}{2 \text{ pints}} = \frac{2 \text{ pints}}{1 \text{ quart}} = \frac{1 \text{ quart}}{0.25 \text{ gallon}} = 1$$

Since the numerator and denominator represent equal quantities in each case, all are valid conversion factors.

Dimensional Analysis

Conversion factors are used in solving problems in which a certain measurement must be expressed with different units. When a given measurement is multiplied by an appropriate conversion factor, the numerical value changes, but the actual size of the quantity measured remains the same. **Dimensional analysis** *is a technique that uses the units (dimensions) of the measurement in order to correctly solve problems*. Dimensional analysis is best illustrated with an example.

Sample Problem 3.1: Dimensional Analysis

How many seconds are in a day?

Step 1: List the known quantities and plan the problem.

Known

- 1 day = 24 hours
- 1 hour = 60 minutes
- 1 minute = 60 seconds

Unknown

- 1 day = ? seconds

The known quantities above represent the conversion factors that we will use. The first conversion factor will have day in the denominator so that the "day" unit will cancel. The second conversion factor will then have hours in the denominator, while the third conversion factor will have minutes in the denominator. As a result, the unit of the last numerator will be seconds and that will be the units for the answer.

Step 2: Calculate

$$1 \text{ d} \times \frac{24 \text{ h}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} = 86,400 \text{ s}$$

Applying the first conversion factor, the “d” unit cancels and 1 × 24 = 24. Applying the second conversion factor, the “h” unit cancels and 24 × 60 = 1440. Applying the third conversion factor, the “min” unit cancels and 1440 × 60 = 86400. The unit that remains is “s” for seconds.

Step 3: Think about your result.

A second is a much smaller unit of time than a day, so it makes sense that there are a very large number of seconds in one day.

Practice Problems

1. How many minutes are in a year?
2. How many days are equal to one million seconds?

Dimensional Analysis and the Metric System

The metric system’s many prefixes allow quantities to be expressed in many different units. Dimensional analysis is useful to convert from one metric system unit to another.

Sample Problem 3.2: Metric Unit Conversions

A particular experiment requires 120 mL of a solution. The teacher knows that he will need to make enough solution for 40 experiments to be performed throughout the day. How many liters of solution should he prepare?

Step 1: List the known quantities and plan the problem.

Known

- 1 experiment requires 120 mL of solution
- 1 L = 1000 mL

Unknown

- 40 experiments require ? L of solution

Since each experiment requires 120 mL of solution and the teacher needs to prepare enough for 40 experiments, multiply 120 by 40 to get 4800 mL of solution needed. Now you must convert mL to L by using a conversion factor.

Step 2: Calculate

$$4800 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 4.8 \text{ L}$$

Note that conversion factor is arranged so that the mL unit is in the denominator and thus cancels out, leaving L as the remaining unit in the answer.

Step 3: Think about your result.

A liter is much larger than a milliliter, so it makes sense that the number of liters required is less than the number of milliliters.

Practice Problems

3. Perform the following conversions.
   (a) 0.074 km to m
   (b) 24,600 μg to g
   (c) 1300 ms to s
   (d) 3.8×10^{-5} L to μL

Some metric conversion problems are most easily solved by breaking them down into more than one step. When both the given unit and the desired unit have prefixes, one can first convert to the simple (unprefixed) unit, followed by a conversion to the desired unit. An example will illustrate this method.

Sample Problem 3.3: Two-Step Metric Conversion

Convert 4.3 cm to μm.

Step 1: List the known quantities and plan the problem.

Known

- 1 m = 100 cm
- 1 m = 10^6 μm

Unknown

- 4.3 cm = ? μm

You may need to consult the table in the "The International System of Units" lesson for the multiplication factor represented by each metric prefix. First convert cm to m, then convert m to μm.

Step 2: Calculate

$$4.3 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{10^6\ \mu\text{m}}{1 \text{ m}} = 43,000\ \mu\text{m}$$

Each conversion factor is written so that unit of the denominator cancels with the unit of the numerator of the previous factor.

Step 3: Think about your result.

A micrometer is a smaller unit of length than a centimeter, so the answer in micrometers is larger than the number of centimeters given.

Practice Problems

4. Perform the following conversions.
   (a) 4.9×10^7 μg to kg
   (b) 84 dm to mm
   (c) 355 nm to cm
   (d) 70.5 ML to mL

Dimensional Analysis and Derived Units

Some units are combinations of SI base units. A **derived unit** *is a unit that results from a mathematical combination of SI base units*. We have already discussed volume and energy as two examples of derived units. Some others are listed in **Table** 2.3

TABLE 2.3: Derived SI Units

Quantity	Symbol	Unit	Unit Abbreviation	Derivation
Area	A	square meter	m^2	length × width
Volume	V	cubic meter	m^3	length × width × height
Density	D	kilograms per cubic meter	kg/m^3	mass / volume
Concentration	c	moles per liter	mol/L	amount / volume
Speed (velocity)	v	meters per second	m/s	length / time
Acceleration	a	meters per second per second	m/s^2	speed / time
Force	F	newton	N	mass × acceleration
Energy	E	joule	J	force × length

Using dimensional analysis with derived units requires special care. When units are squared or cubed as with area or volume, the conversion factors themselves must also be squared. Shown below is the conversion factor for cubic centimeters and cubic meters.

$$\left(\frac{1\text{ m}}{100\text{ cm}}\right)^3 = \frac{1\text{ m}^3}{10^6\text{ cm}^3} = 1$$

Because a cube has 3 sides, each side is subject to the conversion of 1 m to 100 cm. Since 100 cubed is equal to 1 million (10^6), there are 10^6 cm^3 in 1 m^3. Two convenient volume units are the liter, which is equal to a cubic decimeter, and the milliliter, which is equal to a cubic centimeter. The conversion factor would be:

$$\left(\frac{1\text{ dm}}{10\text{ cm}}\right)^3 = \frac{1\text{ dm}^3}{1000\text{ cm}^3} = 1$$

There are thus 1000 cm^3 in 1 dm^3, which is the same thing as saying there are 1000 mL in 1 L (**Figure** 2.6).

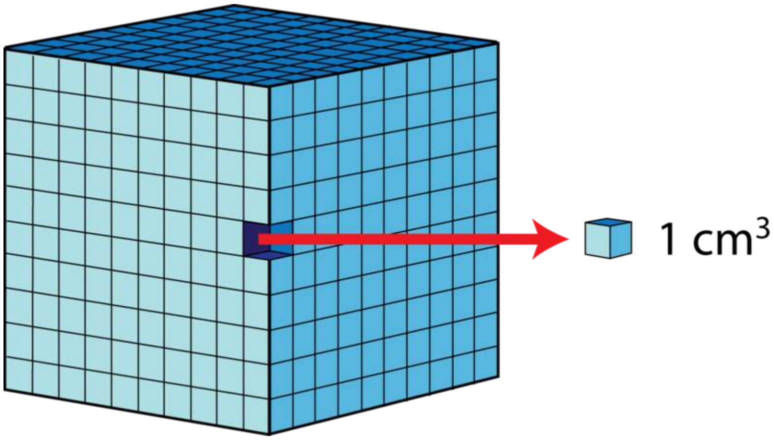


FIGURE 2.6

There are 1000 cm^3 in 1 dm^3. Since 1 cm^3 is equal to 1 mL and 1 dm^3 is equal to 1 L, we can say that there are 1000 mL in 1 L.

You can participate in an interactive version of this cube at www.dlt.ncssm.edu/core/Chapter1-Introduction/Chapter1-Animations/M3_DM3_CM3.html .

Sample Problem 3.4: Derived Unit Conversion

Convert 3.6×10^8 mm^3 to mL.

Step 1: List the known quantities and plan the problem.

Known

- 1 m = 1000 mm
- 1 mL = 1 cm^3
- 1 m = 100 cm

Unknown

- 3.6 mm^3 = ? mL

This problem requires multiple steps and the technique for converting with derived units. Simply proceed one step at a time: mm^3 to m^3 to cm^3 = mL.

Step 2: Calculate

$$3.6\ \text{mm}^3 \times \left(\frac{1\ \text{m}}{1000\ \text{mm}}\right)^3 \times \left(\frac{100\ \text{cm}}{1\ \text{m}}\right)^3 \times \frac{1\ \text{mL}}{1\ \text{cm}^3} = 0.0036\ \text{mL}$$

Numerically, the steps are to divide 3.6 by 10^9, followed by multiplying by 10^6. Alternatively, you can shorten the calculation by one step if you first determine the conversion factor between mm and cm. Using the fact that there are 10 mm in 1 cm, you can avoid the intermediate step of converting to meters.

$$3.6\ \text{mm}^3 \times \left(\frac{1\ \text{cm}}{10\ \text{mm}}\right)^3 \times \frac{1\ \text{mL}}{1\ \text{cm}^3} = 0.0036\ \text{mL}$$

Both conversion methods give the same result of 0.0036 mL.

Step 3: Think about your result.

Cubic millimeters are much smaller than cubic centimeters, so the final answer is much less than the original number of mm^3.

Practice Problems

5. Perform the following conversions.
   (a) 0.00722 km^2 to m^2
   (b) 129 cm^3 to L
   (c) 4.9×10^5 μm^3 to mm^3

You can find more help with dimensional analysis by watching this video at www.chemcollective.org/stoich/dimensionalanalysis.php .

You can download an instructional packet that explains dimensional analysis step by step and provides practice problems at https://docs.google.com/open?id=0B_ZuEGrhVEfMS09WeUtOMFNTaFk .

Density

A golf ball and a table tennis ball are about the same size. However, the golf ball is much heavier than the table tennis ball. Now imagine a similar size ball made out of lead. That would be very heavy indeed! What are we comparing?

By comparing the mass of an object relative to its size, we are studying a property called density. **Density** *is the ratio of the mass of an object to its volume.*

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad \text{D} = \frac{\text{m}}{\text{V}}$$

Density is an intensive property, meaning that it does not depend on the amount of material present in the sample. For example, water has a density of 1.0 g/mL. That density is the same whether you have a small glass of water or a swimming pool full of water. Density is a property that is constant for a specific type of matter at a given temperature.

The SI units of density are kilograms per cubic meter (kg/m^3), because the kg and the m are the SI units for mass and length, respectively. Unfortunately, this unit is awkwardly large for everyday usage in the laboratory. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter (g/cm^3). Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L. Densities of some common substances at 20°C are listed in **Table** 2.4. Since most materials expand as temperature increases, the density of a substance is temperature-dependent and usually decreases as temperature increases.

TABLE 2.4: **Densities of Some Common Substances**

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084
Ice (0°C)	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water (4°C)	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.92		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

You know from experience that ice floats in water. This is consistent with the values in **Table** 2.4, which show that ice is less dense than liquid water. Alternatively, corn syrup would sink if placed into water, because it has a higher density. **Figure** 2.7 shows a number of substances arranged into a density column.

With gases, a balloon filled with helium floats because helium is less dense than air. However, a balloon filled with carbon dioxide sinks because carbon dioxide is denser than air.

Sample Problem 3.5: Density Calculations

An 18.2 g sample of zinc metal has a volume of 2.55 cm^3. Calculate the density of zinc.

Step 1: List the known quantities and plan the problem.

Known

- mass = 18.2 g
- volume = 2.55 cm^3

Unknown

- density = ? g/cm^3

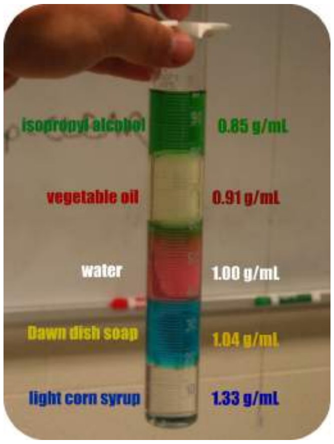


FIGURE 2.7

A density column made of some familiar materials.

Use the equation for density, D = m/V, to solve the problem.

Step 2: Calculate

$$D = \frac{m}{V} = \frac{18.2\text{ g}}{2.55\text{ cm}^3} = 7.14\text{ g/cm}^3$$

Step 3: Think about your result.

If 1 cm^3 of zinc has a mass of about 7 grams, then a sample that is approximately 2.5 cm^3 will have a mass between 2 and 3 times its density, which is consistent with the values given in this problem. Additionally, metals are expected to have a density that is greater than that of water, and zinc's density falls within the range of the other metals listed in **Table** 2.4.

Since density values are known for many substances, density can be used to determine an unknown mass or an unknown volume. Dimensional analysis will be used to ensure that units cancel appropriately.

Sample Problem 3.6: Using Density to Determine Mass and Volume

1. What is the mass of 2.49 cm^3 of aluminum?
2. What is the volume of 50.0 g of aluminum?

Step 1: List the known quantities and plan the problem.

Known

- density = 2.70 g/cm^3
- 1. volume = 2.49 cm^3
- 2. mass = 50.0 g

Unknown

- 1. mass = ? g
- 2. volume = ? cm^3

Use the equation for density, D = m/V, and dimensional analysis to solve each problem.

Step 2: Calculate

1. $2.49\ \text{cm}^3 \times \dfrac{2.70\ \text{g}}{1\ \text{cm}^3} = 6.72\ \text{g}$
2. $50.0\ \text{g} \times \dfrac{1\ \text{cm}^3}{2.70\ \text{g}} = 18.5\ \text{cm}^3$

In problem 1, the mass is equal to the density multiplied by the volume. In problem 2, the volume is equal to the mass divided by the density.

Step 3: Think about your results.

Because a 1 cm^3 sample of aluminum has a mass of 2.70 g, the mass of a 2.49 cm^3 sample should be greater than 2.70 g. The mass of a 50-g block of aluminum is substantially more than the value of its density in g/cm^3, so that amount should occupy a volume that is significantly larger than 1 cm^3.

Practice Problems

6. A student finds the mass of a "gold" ring to be 41.7 g and its volume to be 3.29 cm^3. Calculate the density of the ring. Is it pure gold? (See **Table** 2.4.)
7. What is the mass of 125 L of oxygen gas?
8. The density of silver is 10.5 g/cm^3. What is the volume of a 13.4 g silver coin?

Finally, conversion problems involving density or other derived units like speed or concentration may involve a separate conversion of each unit. To convert a density in g/cm^3 to kg/m^3, two steps must be used. One step converts g to kg, while the other converts cm^3 to m^3. They may be performed in either order.

Sample Problem 3.7: Density Conversions

The average density of the planet Jupiter is 1.33 g/cm^3. What is Jupiter's density in kg/m^3?

Step 1: List the known quantities and plan the problem.

Known

- density = 1.33 g/cm^3
- 1 kg = 1000 g
- 1 m = 100 cm

Unknown

- density = ? kg/m^3

Use separate conversion factors to convert the mass from g to kg and the volume from cm^3 to m^3.

Step 2: Calculate

$$\frac{1.33\ \text{g}}{\text{cm}^3} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \left(\frac{100\ \text{cm}}{1\ \text{m}}\right)^3 = 1330\ \text{kg/m}^3$$

Step 3: Think about your result.

Since a cubic meter is so much larger (1 million times) than a cubic centimeter, the density of Jupiter is larger in kg/m^3 than in g/cm^3.

You can perform a density experiment to identify a mystery object online. Find this simulation at http://phet.colorado.edu/en/simulation/density .

Lesson Summary

- Conversion factors are ratios of equivalent quantities expressed in different units. When multiplying by a conversion factor, the numerical value and the unit changes while the actual size of the quantity remains the same.
- Dimensional analysis employs conversion factors to solve problems in which the units are changing. Dimensional analysis can be used to solve metric system conversion problems.
- Density is a derived unit of mass per unit volume and is a physical property of a substance. Density problems can be solved using dimensional analysis.

Lesson Review Questions

Reviewing Concepts

1. What must be true for a ratio of two measurements to be a conversion factor?
2. Which of the following ratios qualify as conversion factors? For the ones that do not, explain why.
   a. $\frac{\text{10 pennies}}{\text{1 dime}}$
   b. $\frac{\text{3 dogs}}{\text{several}}$
   c. $\frac{\text{1 hour}}{\text{60 seconds}}$
   d. $\frac{\text{1 dozen}}{\text{12 donuts}}$
3. How do you decide which unit should go in the denominator of a conversion factor?
4. What is a derived unit?
5. Explain what is wrong with this statement: "The density of a heavy bar of pure gold is greater than the density of a small ingot of pure gold."

Problems

6. Make the following conversions.
   a. 128 mL to L
   b. 2.5×10^5 μg to g
   c. 0.481 km to m
   d. 1890 cm to km
   e. 6.2×10^{-5} ms to ns
   f. 75,000 pg to cg
7. Make the following conversions.
   a. 2800 cm^3 to m^3
   b. 5.8 g/cm^3 to g/L
   c. A speed of 60.0 miles per hour to m/s (1 mile = 1608 m)
   d. A flow rate of 125 mL/min to liters per hour
8. The speed of light is 3.0×10^8 m/s. If the distance from Earth to the Sun is 1.5×10^8 km, how many minutes does it take for light from the Sun to reach the Earth?
9. A regular solid has dimensions of 3.20 cm by 4.90 cm by 5.40 cm. The mass of the solid is 235 g. What is its density in g/cm^3?

10. What is the mass of a cube of copper that is 1.80 cm on each side? The density of copper is 8.92 g/cm^3.
11. A balloon is filled with 2300 mL of an unknown gas. The mass of the gas is 3.24 g. Will the balloon float or sink in air?
12. A cube of lead (density = 11.35 g/cm^3) has a mass of 145.7 g. What is the length of each side of the cube in cm?

Further Reading / Supplemental Links

- Barber, Jacqueline; Buegler, Marion; Lowell, Laura, *Discovering Density*. GEMS –Regents of the Univ of CA, 1998.
- Huggins, Shawna, *How to Make a Lava Lamp: Exploring Density*. Blue and Green Wonders Publishing, 2011.
- Fun With Dimensional Analysis: (http://www.alysion.org/dimensional/fun.htm
- Can you make a golf ball float on water? Find out at www.youtube.com/watch?v=cXLTSLa3yYs .
- Watch the density of water change when heated at www.youtube.com/watch?v=GThtJTncPf0 .
- Density differences in liquids and gases can cause convection currents. You can see this at www.youtube.com/watch?v=ovSMAujQbz4 .
- Test your skills with a unit conversion hangman game at http://education.jlab.org/vocabhangman/measurement_04/1.html .

Points to Consider

Measurements all must have a certain amount of uncertainty in them, since no measuring tool is 100% accurate. The uncertainty in a measurement must be considered both when reporting measured values and when doing calculations.

- How is the uncertainty in a given measurement indicated in the reported value?
- When a quantity such as density is calculated from two measurements (mass and volume), is it important to measure both accurately or is just one sufficient?
- What are the meanings of the terms precision and accuracy when dealing with measurements?

2.3 Uncertainty in Measurements

Lesson Objectives

- Distinguish between accuracy and precision in measurements.
- Calculate the percent error of a measured quantity.
- Report measured values to the correct number of significant figures based on the measuring tool.
- Perform calculations with measured quantities, rounding the answers to the correct number of significant figures.

Lesson Vocabulary

- accepted value
- accuracy
- error
- experimental value
- percent error
- precision
- significant figures

Check Your Understanding

Recalling Prior Knowledge

- Suppose that a baseball pitcher throws very accurately. What does that mean? Is it different from throwing very precisely?
- When you make a measurement with a specific measuring tool, how well can you read that measurement? Will individual measurements have an effect on quantities that are calculated from those measurements?

When making a measurement, there is always going to be some uncertainty. Some of that uncertainty is related to the reliability of the measuring tool, while some of it is related to the skill of the measurer. When you are performing measurements, you should always strive for the greatest accuracy and precision that you possibly can.

Accuracy and Precision

In everyday speech, the terms accuracy and precision are frequently used interchangeably. However, their scientific meanings are quite different. **Accuracy** *is a measure of how close a measurement is to the correct or accepted value* *of the quantity being measured.* **Precision** *is a measure of how close a series of measurements are to one another*. Precise measurements are highly reproducible, even if the measurements are not near the correct value.

Darts thrown at a dartboard are helpful in illustrating accuracy and precision (**Figure** 2.8).

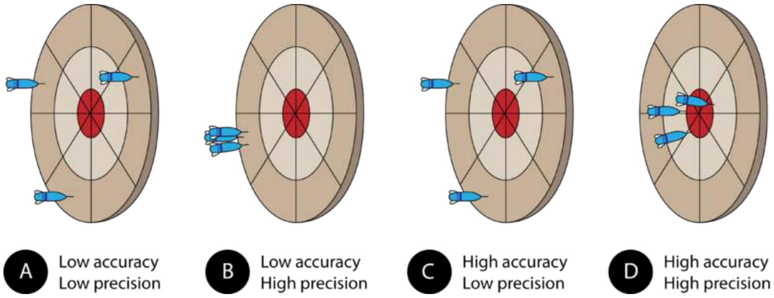


FIGURE 2.8
The distribution of darts on a dartboard shows the difference between accuracy and precision.

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. **Figure** 2.8 demonstrates four possible outcomes.

(A) The darts have landed far from each other and far from the bulls-eye. This grouping demonstrates measurements that are neither accurate, nor precise.

(B) The darts are close to one another, but far from the bulls-eye. This grouping demonstrates measurements that are precise, but not accurate. In a laboratory situation, high precision with low accuracy often results from a systematic error. Either the measurer makes the same mistake repeatedly or the measuring tool is somehow flawed. A poorly calibrated balance may give the same mass reading every time, but it will be far from the true mass of the object.

(C) The darts are not grouped very near to each other, but they are generally centered around the bulls-eye. This demonstrates poor precision but fairly high accuracy. This situation is not desirable in a lab situation because the "high" accuracy may simply be due to random chance and is not a true indicator of good measuring skill.

(D) The darts are grouped together and have hit the bulls-eye. This demonstrates high precision and high accuracy. Scientists always strive to maximize both in their measurements.

FIGURE 2.9
Students in a chemistry lab are making careful measurements with a series of volumetric flasks. Accuracy and precision are critical in every experiment.

Percent Error

An individual measurement may be accurate or inaccurate, depending on how close it is to the true value. Suppose that you are doing an experiment to determine the density of a sample of aluminum metal. The **accepted value** *of a measurement is the true or correct value based on general agreement with a reliable reference*. For aluminum, the accepted density is 2.70 g/cm^3. The **experimental value** *of a measurement is the value that is measured during the experiment*. Suppose that in your experiment you determine an experimental value of 2.42 g/cm^3 for the density of aluminum. The **error** *of an experiment is the difference between the experimental and accepted values*.

$$\text{Error} = \text{experimental value} - \text{accepted value}$$

If the experimental value is less than the accepted value, the error is negative. If the experimental value is larger than the accepted value, the error is positive. Often, error is reported as the absolute value of the difference in order to avoid the confusion of a negative error. The **percent error** *is the absolute value of the error divided by the accepted value and multiplied by 100%*.

$$\text{Percent Error} = \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

To calculate the percent error for the aluminum density measurement, we can substitute the given values of 2.45 g/cm^3 for the experimental value and 2.70 g/cm^3 for the accepted value.

$$\text{Percent Error} = \frac{|2.45 \text{ g/cm}^3 - 2.70 \text{ g/cm}^3|}{2.70 \text{ g/cm}^3} \times 100\% = 9.26\%$$

If the experimental value is equal to the accepted value, the percent error is equal to 0. As the accuracy of a measurement decreases, the percent error of that measurement rises.

Significant Figures in Measurements

Uncertainty

Some error or uncertainty always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. **Figure** 2.10 shows an object (indicated by the blue arrow) whose length is being measured by two different rulers.

With either ruler, it is clear that the length of the object is between 2 and 3 cm. The bottom ruler contains no millimeter markings, so the tenths digit can only be estimated, and the length may be reported by one observer as 2.5 cm. However, another person may judge that the measurement is 2.4 cm or perhaps 2.6 cm. While the 2 is known for certain, the value of the tenths digit is uncertain.

The top ruler contains marks for tenths of a centimeter (millimeters). Now, the same object may be measured as 2.55 cm. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be 2.54 cm or 2.56 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

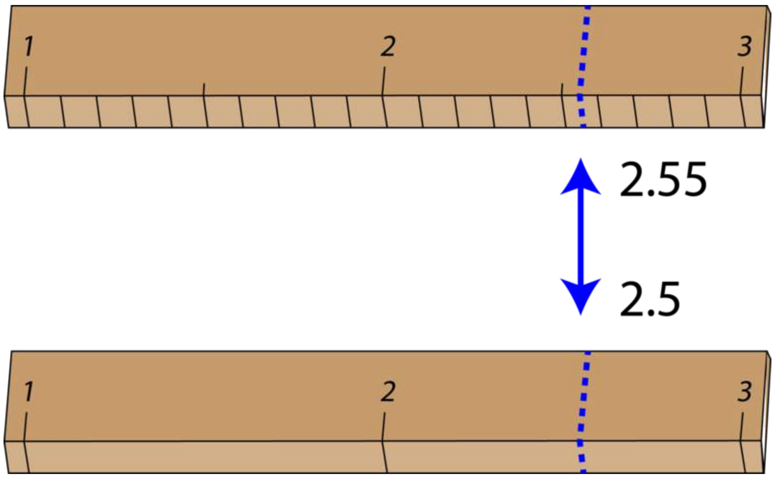


FIGURE 2.10

Determining Significant Figures

The **significant figures** *in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit*. In the ruler example, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 3 significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. It would not be correct to report the length as 2.553 cm with either ruler, because there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. **Table** 2.5 details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

TABLE 2.5: **Significant Figure Rules**

Rule	Examples
1. All nonzero digits in a measurement are significant	A. 237 has three significant figures. B. 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits are always significant.	A. 39,004 has five significant figures. B. 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called left-end zeros. Left-end zeros are never significant.	A. 0.008 has one significant figure. B. 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called right-end zeros. Right-end zeros in a number that lacks a decimal point are not significant.	A. 140 has two significant figures. B. 75,210 has four significant figures.
5. Right-end zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	A. 620.0 has four significant figures. B. 19.000 has five significant figures

It needs to be emphasized that just because a certain digit is not significant does not mean that it is not important or that it can be left out. Though the zero in a measurement of 140 may not be significant, the value cannot simply be reported as 14. An insignificant zero functions as a placeholder for the decimal point. When numbers are written in scientific notation, this becomes more apparent. The measurement 140 can be written as 1.4×10^2, with two significant figures in the coefficient. A number with left-end zeros, such as 0.000416, can be written as 4.16×10^{-4}, which has 3 significant figures. In some cases, scientific notation is the only way to correctly indicate the correct number of significant figures. In order to report a value of 15,000,000 with four significant figures, it would need to be written as 1.500×10^7. The right-end zeros after the 5 are significant. The original number of 15,000,000 only has two significant figures.

How many significant figures are there in each of the following measurements?

1. 19.5 m
2. 0.0051 L
3. 204.80 g
4. 1.90 × 10^5 s
5. 14 beakers
6. 700 kg

Step 1: Plan the problem.

Follow the rules for counting the number of significant figures in a measurement, paying special attention to the location of zeros in each. Note each rule that applies according to **Table** 2.5.

Step 2: Solve

1. three (rule 1)
2. two (rule 3)
3. five (rules 2 5)
4. three (rule 5)
5. infinite
6. one (rule 4)

The 14 beakers is a counted set of items and not a measurement, so it has an infinite number of significant figures.

Practice Problems

1. Count the number of significant figures in each measurement.
    a. 0.00090 L
    b. 255 baseballs
    c. 435,210 m
    d. 40.1 kg
    e. 9.026 × 10^{-6} mm
    f. 12.40°C

Significant Figures in Calculations

Many reported quantities in science are the result of calculations involving two or more measurements. Density involves mass and volume, both of which are measured quantities. As an example, say that you have a precise balance that gives the mass of a certain object as 21.513 g. However, the volume is measured very roughly by water displacement, using a graduated cylinder that can only be read to the nearest tenth of a milliliter. The volume of the object is determined to be 8.2 mL. On a calculator, the density (mass divided by volume) would come out as 2.623536585 g/mL. Hopefully, it should be apparent that the calculator is giving us far more digits than we actually can be certain of knowing. In fact, the density should be reported as 2.6 g/mL. This is because the result of a calculated answer can be no more precise than the least precise measurement from which it was calculated. Since the volume was known only to two significant figures, the resultant density needs to be rounded to two significant figures.

Rounding

Before we get to the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To round a number, first decide how many significant figures the number should have. Once you know that, round to the correct number of digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped, and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement 207.518 m. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process in **Table** 2.6.

TABLE 2.6: **Significant Figure Rounding**

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Significant Figures in Addition and Subtraction

Consider two separate mass measurements: 16.7 g and 5.24 g. The first mass measurement (16.7 g) is known only to the tenths place, which is one digit after the decimal point. There is no information about its hundredths place, so that digit cannot be assumed to be zero. The second measurement (5.24 g) is known to the hundredths place, which is two digits after the decimal point.

When these masses are added together, the result on a calculator is 16.7 + 5.24 = 21.94 g. Reporting the answer as 21.94 g suggests that the sum is known all the way to the hundredths place. However that cannot be true because the hundredths place of the first mass was completely unknown. The calculated answer needs to be rounded in such a way as to reflect the certainty of each of the measured values that contributed to it. *For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the lowest number of decimal places*. The sum of the above masses would be properly rounded to a result of 21.9 g.

When working with whole numbers, pay attention to the last significant digit that is to the left of the decimal point and round your answer to that same point. For example, consider the subtraction problem 78,500 m –362 m. The calculated result is 78,138 m. However, the first measurement is known only to the hundreds place, as the 5 is the last significant digit. Rounding the result to that same point means that the final calculated value should be reported as 78,100 m.

Significant Figures in Multiplication and Division

The density of a certain object is calculated by dividing the mass by the volume. Suppose that a mass of 37.46 g is divided by a volume of 12.7 cm^3. The result on a calculator would be:

$$\mathrm{D} = \frac{\mathrm{m}}{\mathrm{V}} = \frac{37.46\ \mathrm{g}}{12.7\ \mathrm{cm}^3} = 2.94960299\ \mathrm{g/cm}^3$$

The value of the mass measurement has four significant figures, while the value of the volume measurement has only three significant figures. *For multiplication and division problems, the answer should be rounded to the same* *number of significant figures as the measurement with the lowest number of significant figures.* Applying this rule results in a density of 2.95 g/cm^3, which has three significant figures –the same as the volume measurement.

Note that the rule for multiplication and division problems is different than the rule for addition and subtraction problems. For multiplication and division, it is the number of significant figures that must be considered. For addition and subtraction, it is the position of the decimal place that determines the correct rounding. Review the sample problem below, paying special attention to this distinction.

Sample Problem 3.9: Significant Figures in Calculations

Perform the following calculations, rounding the answers to the appropriate number of significant figures.

1. 0.048 m × 32.97 m
2. 21.9 g –19.417 g
3. 14,570 kg ÷ 5.81 L
4. 71.2 cm + 90 cm

Step 1: Plan the problem.

Decide which calculation rule applies. Analyze each of the measured values to determine how many significant figures should be in the result. Perform the calculation and round appropriately. Apply the correct units to the answer. When adding or subtracting, the units in each measurement must be identical and then remain the same in the result. When multiplying or dividing, the units are also multiplied or divided.

Step 2: Calculate.

1. 0.048 m × 32.97 m = 1.6 m^2 - Round to two significant figures because 0.048 has two.
2. 21.9 g - 19.417 g = 2.5 g - Answer ends at the tenths place because of 21.9.
3. 14,570 kg ÷ 5.81 L = 2510 kg/L - Round to three significant figures because 5.81 has three.
4. 71.2 cm + 90 cm = 160 cm - Answer ends at the tens place because of 90.

Practice Problems

2. Solve each problem, rounding each answer to the correct number of significant figures.
    (a) 132.3 g ÷ 29.600 mL
    (b) 3.27 g/cm^3 × 0.086 cm^3
    (c) 125 m + 61.3 m + 310 m
    (d) 3.0×10^4 L − 1244 L
3. A rectangular prism has dimensions of 3.7 cm by 4.81 cm by 1.90 cm. The mass of the prism is 49.72 g. Calculate its density. (Hint: Do the entire calculation without rounding until the final answer in order to reduce "rounding error.")

Lesson Summary

- Accuracy refers to how close a measured value is to the accepted value, whereas precision indicates how close individual measurements within a set are to each other.
- Percent error is the difference between the experimental and accepted values divided by the accepted value and multiplied by 100.

- The measuring tool dictates how many significant figures can be reported in a measurement. Significant figures include all of the certain digits plus one uncertain digit. A set of rules is followed for determining the number of significant figures in numbers that contain zeros. Counted quantities have infinite significant figures.
- For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the lowest number of decimal places. For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the lowest number of significant figures.

Lesson Review Concepts

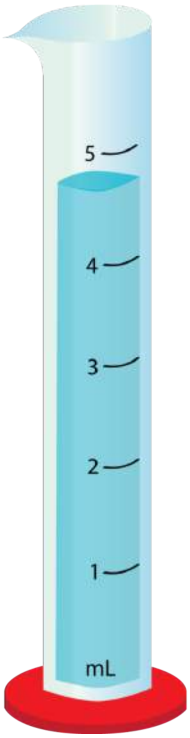
Reviewing Concepts

1. The density of a sample of copper metal was determined by three different students (see **Table** 2.7). Each performed the measurement three times. Describe the accuracy and precision of each student's measurements. The accepted value for the density of copper is 8.92 g/cm^3.

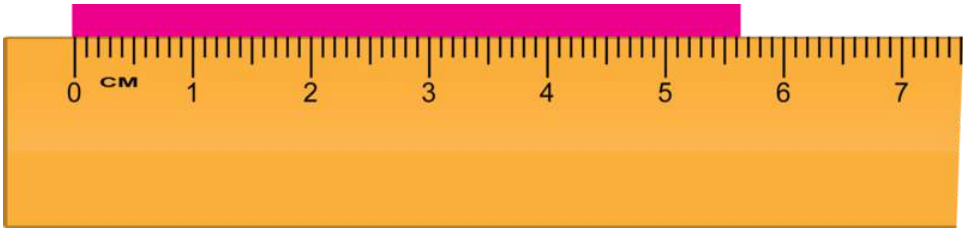
TABLE 2.7: **Density of copper (g/cm**

Student	Trial 1	Trial 2	Trial 3
Jane	8.94	8.89	8.91
Justin	8.32	8.31	8.34
Julia	8.64	9.71	9.13

2. What is wrong with the following statement? "My measurement of 8.45 m for the width of the room is very precise."
3. Consider the following 5 mL graduated cylinders, which contain identical quantities of liquid. Which cylinder yields a measurement with a greater number of significant figures? How many significant figures can be reported for each cylinder?



4. Report the length measurement of the pink bar to the correct number of significant figures. Which digits in your measurement are certain? Which are uncertain?



5. How many significant figures are in each of the following measurements?
   a. 9 potatoes
   b. 4.05 cm
   c. 0.0061 kg
   d. 50 mL
   e. 8.00×10^9 μg
   f. 720.00 s
6. Round the measured quantity of 31.0753 g to each of the following amounts of significant figures.
   a. five
   b. four
   c. three
   d. two
   e. one

Problems

7. Kyle measures the mass of a solid sample to be 8.09 g. The accepted value for the mass is 8.42 g. Calculate Kyle's percent error.

8. Jamelle performs three separate determinations of the density of a mineral sample. She gets values of 4.58 g/cm^3, 4.79 g/cm^3, and 4.55 g/cm^3.
   a. Calculate the average value of the density of the mineral.
   b. The deviation of a measured value is defined as the absolute value of the difference between the measured value and the average value: Deviation = |measured value − average value|. Calculate the deviation for each of the three measurements. According to the deviations, which measurement appears to be poorest compared to the others?
   c. The average deviation is the sum of all the deviations divided by the total number of measurements. Calculate the average deviation of the three measurements.
   d. When average deviation is high, does that indicate good precision or poor precision in the measurements? Explain.
9. Convert a measurement of 2.75 hours to seconds. Are the required conversion factors measured or exact quantities?
10. Convert 466.84 cm to inches, given that 1 inch = 2.54 cm. The conversion factor between centimeters and inches is a measured quantity.
11. Perform the following calculations and round to the correct number of significant figures.
    a. 78.2 g ÷ 32 cm^3
    b. 3.0 m/s × 9.21 s
    c. 59 g + 4 g + 0.79 g
    d. 34,000 km − 2430 km
    e. (9.59 g + 1.098 g) ÷ 2.313 mL

Further Reading / Supplemental Links

- Tutorial on the Use of Significant Figures: http://www.chem.sc.edu/faculty/morgan/resources/sigfigs/index.html
- Another tutorial video can be found at www.youtube.com/watch?v=ctj07mSIJ0w .
- Significant Figures Calculator: http://calculator.sig-figs.com/
- Winter, Paul K., *Significant Figures*. International Textbook Company, 1965.

Points to Consider

Measurements will be a constant consideration throughout your study of chemistry. Next you will begin a study of the atom, its component parts, and the evolution of the atomic model.

- Atoms are extremely small and extremely light. How do you think that the mass and size of an atom can be measured? Do you think the accuracy of these measurements has improved over time?

2.4 References

1. Courtesy of NASA/SDO. http://en.wikipedia.org/wiki/File:The_Sun_by_the_Atmospheric_Imaging_Assembly_of_NASA%27s_Solar_Dynamics_Observatory_-_20100819.jpg . Public Domain
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CHAPTER 3 Matter and Change

Chapter Outline

Freshly baked bread and cherry pie are two delicious parts of any meal. What happens to the ingredients that go into the bread and the pie as they are heated in the oven? If you have ever done any baking, you certainly know that the bread dough that is put into the oven changes dramatically as it bakes. Many chemical reactions occur, and one of those reactions releases a gas that causes the bread to rise. In this chapter, you will learn about some of the basic ways in which matter is capable of changing and how we can make sense of those changes. You will also learn how to classify the various types of matter that are part of the world around us.

Opening image courtesy of jcroninone. CC−BY 2.0.

3.1 Properties of Matter

Lesson Objectives

- Describe a substance according to its physical properties.
- Distinguish between extensive and intensive properties.
- Describe the three states of matter.
- Identify physical changes to matter.

Lesson Vocabulary

- extensive property
- gas
- intensive property
- liquid
- mass
- physical change
- physical property
- solid
- substance
- vapor
- volume

Check Your Understanding

Recalling Prior Knowledge

- What are some physical properties that can be used to describe matter?
- What are the three states of matter? Describe their similarities and differences.

The science of chemistry is largely concerned with the understanding and classification of matter. This understanding naturally begins with careful observations. What does a particular sample of matter look like? Is it a solid, liquid, or gas? Is it hard or soft? Is it shiny or dull? These types of questions are where we will begin our study of matter and its properties.

If you would like to refresh your memory about states of matter, go to "The Phantom's Portrait Parlor: Phases of Matter" at www.miamisci.org/af/sln/phases/watersolid.html and view solids, liquids, and gases at the molecular level.

Description of Matter

Recall that matter is considered to be anything that has mass and occupies space. Matter can be best understood in terms of its properties. Some properties of matter can be determined without altering the identity of the matter, and it is there that we will begin.

Physical Properties

Figure 3.1 shows two very different objects. Each has a different chemical composition, meaning that each is made up of a different type of matter.

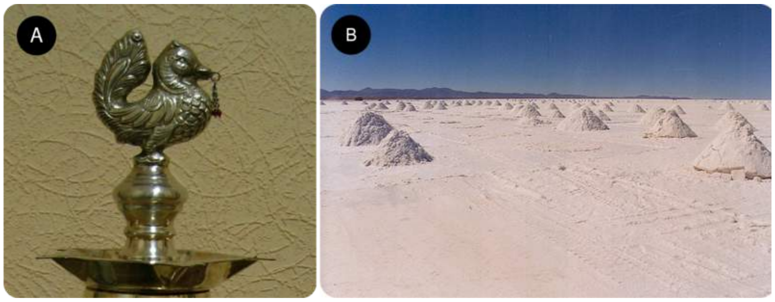

FIGURE 3.1
(A) This lamp from the country of Sri Lanka is made of silver. (B) These large salt mounds from Bolivia were produced during a salt mining operation.

Silver and ordinary table salt are two examples of chemical substances. A **substance** *is matter that has a uniform and definite composition.* All samples of substances, sometimes called pure substances, have identical properties.

A **physical property** *is a characteristic of a substance that can be observed or measured without changing the identity of the substance.* For example, silver is a shiny metal that conducts electricity very well. It can also be molded into thin sheets, a property called malleability. These are all physical properties of silver. In contrast, salt is dull and brittle, but it can also conduct electricity after being dissolved in water, which it does quite readily. In addition to color, hardness, malleability, solubility, and electrical conductivity, some other physical properties of matter include density, melting point, and boiling point.

There are many examples of physical properties that can be further investigated on the internet.

- Magnetism is a physical property. Are all metals magnetic? Find out at www.youtube.com/watch?v=iK-5-BRLn78 .
- You can view a video of the melting points of various substances at www.youtube.com/watch?v=NYb2iTNFCsI .
- Do you think metals like lead melt easily? Check out this alloy containing lead melting in boiling water! www.youtube.com/watch?v=JI9qKbQCukg .
- The physical property density can be observed by floating bubbles on carbon dioxide. Watch the video at www.youtube.com/watch?v=h3EWYrxrsns .
- The melting and freezing of *tert*-butanol can be viewed at these two links:
  - Melting: www.dlt.ncssm.edu/core/Chapter2-Phys_and_Chem_Prop/melt_tbutanol-lg.htm
  - Freezing: www.youtube.com/watch?v=lQnrRIDpPlU

Extensive and Intensive Properties

Some properties of matter depend on the size of the sample, while some do not. An **extensive property** *is a property that depends on the amount of matter in a sample*. The **mass** of an object *is a measure of the amount of matter that an object contains*. A small sample of a certain type of matter will have a small mass, while a larger sample will have a greater mass. Another extensive property is volume. The **volume** of an object *is a measure of the space that is occupied by that object*.

An **intensive property** *is a property of matter that depends only on the type of matter in a sample and not on the amount*. For example, the electrical conductivity of a substance is a property that depends only on the type of substance. Silver, gold, and copper are excellent conductors of electricity, while glass and plastic are poor conductors. A larger or smaller piece of glass will not change this property. Other intensive properties include color, temperature, density, and solubility.

For more information on extensive and intensive properties, visit www.chem.tamu.edu/class/majors/tutorialnotefiles/intext.htm .

States of Matter

Water is a very common substance that we frequently encounter in all three states of matter, as seen in **Figure** 3.2. When water is in the solid state, we call it ice, while water in the gaseous state is referred to as steam or water vapor. The physical state of matter is a physical property because the identity of a substance does not change when it is melted, frozen, or boiled.

FIGURE 3.2
Water is the same substance in any of its three states. (A) A frozen waterfall in Hungary. (B) The Nile River in Egypt. (C) A steam powered train in Wales.

Solid

A **solid** *is a form of matter that has a definite shape and volume*. The shape of a solid does not change if it is transferred from one container to another. The particles of a solid are packed tightly together in fixed positions, usually in an orderly arrangement (**Figure** 3.3). Solids are almost completely incompressible, meaning that solids cannot be squeezed into a smaller volume. When a solid is heated, it expands only slightly.

Liquid

A **liquid** *is a form of matter that has a definite volume, but an indefinite shape*. As water is poured from one container into another, it adopts the shape of its new container. However, the volume of the sample does not change, because the water molecules are still relatively close to one another in the liquid state. Its ability to flow means that water is considered a fluid. Unlike a solid, the arrangement of particles in a liquid is not rigid and orderly. Liquids are also incompressible, and the volume of water expands slightly when heated.

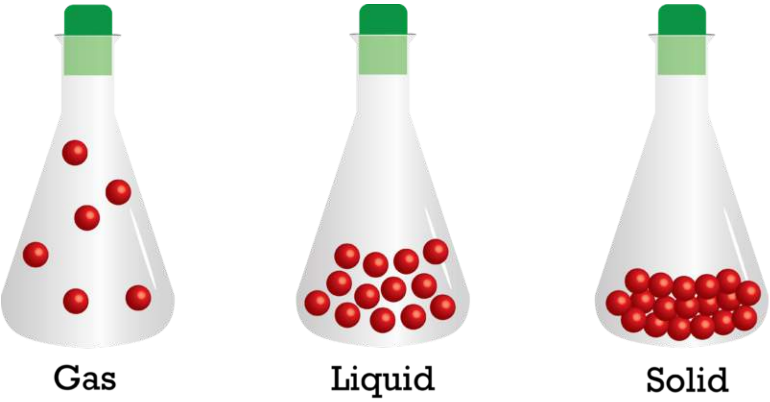

FIGURE 3.3
The particles of a gas are very far apart compared to the particles of a liquid or a solid.

Gas

A **gas** *is a form of matter that has neither a definite shape nor a definite volume*. A gas takes up the shape and volume of its container. This is because the particles of a gas are very, very far apart from one another compared to the particles that make up solids and liquids. Gases, like liquids, are fluids. Gases are easily compressed because of the large spaces in between gas particles. The term gas is most often used to refer to a substance that is found in the gas state at room temperature, such as hydrogen and nitrogen. However, substances that are solids or liquids at room temperature can also evaporate into the gaseous state. A **vapor** *is the gaseous state of a substance that is a solid or liquid at room temperature*. Examples include water vapor or sodium vapor (**Figure** 3.4).

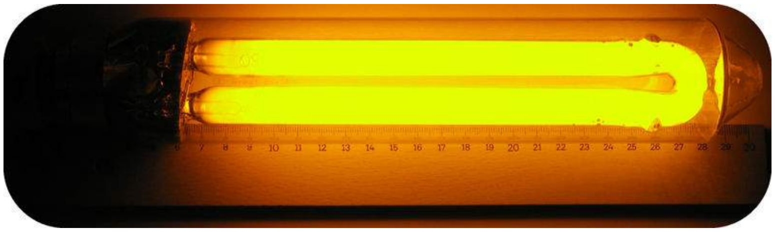

FIGURE 3.4
Sodium vapor lamps glow with a distinctive yellow color.

For a summary of the phases of matter, visit www.chem.purdue.edu/gchelp/atoms/states.html .

A fun, interactive simulation about phases of matter can be downloaded at www.concord.org/activities/states-matter .

Another interactive simulation that will help you understand solids, liquids, and gases can be found at http://phet.colorado.edu/en/simulation/states-of-matter . A variety of activity guides can be found by scrolling down.

NASA provides diagrams and explanations of states of matter at http://www.grc.nasa.gov/WWW/k-12/airplane/state.html

NASA explains properties of gases in an informative website at http://www.grc.nasa.gov/WWW/k-12/airplane/gasprop.html .

Physical Changes

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. Melting is an example of a physical change. A **physical change** *is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not*. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include vaporization (liquid to gas), freezing (liquid to solid), and condensation (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved in water, the salt is said to have entered the aqueous state. Boiling off the water will cause the salt to reform in its solid state.

An example of an irreversible physical change would be grinding a piece of wood into sawdust. Such a change is irreversible because the sawdust cannot be reconstituted into the same piece of wood that it was before. Cutting the grass or pulverizing a rock would be other irreversible physical changes.

Some other examples of physical changes can be seen on the web:

- Iodine undergoes an interesting physical change called sublimation. Click www.youtube.com/watch?v=4E0sy-FN2M8 to watch!
- Metals expand when heated. This physical change can be seen using a ball and ring at www.youtube.com/watch?v=QNoE5IoRheQ .

Lesson Summary

- A substance is a sample of matter with a fixed composition and identifiable physical properties that are the same for every sample of a particular substance.
- Extensive properties depend on the size of a sample, while intensive properties do not.
- Solid, liquid, and gas are the three common states of matter.
- Physical changes do not alter the identity of a substance.

Lesson Review Questions

Reviewing Concepts

1. Explain why water is considered a substance, while milk is not.
2. List at least three physical properties of silver.
3. Give an example of an extensive property and of an intensive property of an iron nail. Explain.
4. Compare the arrangements of particles in the three states of matter.
5. The odor of gasoline is very strong and distinctive. Should gasoline in the gaseous state be called a gas or a vapor? Explain.
6. Describe how the sharpening of a pencil is a different type of physical change than dissolving some sugar in water.

Problems

Use the **Table** 3.1 to answer the following questions.

TABLE 3.1: Properties of Substances

Substance	Color	Melting Point (°C)	Boiling Point (°C)
Bromine	red-brown	-7	59
Chlorine	yellow-green	-101	-34
Gold	yellow	1064	2856
Iron	silver	1538	2862
Mercury	silvery-white	-117	78
Nitrogen	colorless	-210	-196
Oxygen	colorless	-219	-183
Sodium chloride	white	801	1413
Water	colorless	0	100

7. What is the physical state of each of the following substances at room temperature?
   a. oxygen
   b. iron
   c. mercury
8. What is the name of the physical change that occurs in each of the following instances?
   a. Oxygen is cooled from -150°C to -200°C.
   b. Iron is cooled from 1600°C to 1500°C.
   c. Mercury is heated from -130°C to -100°C.
9. Which colorless substance is a solid at -50°C?
10. Which substance is a liquid at 1000°C?
11. Which substance condenses from a gas to a liquid as the temperature is cooled from 20°C to -50°C?

Further Reading / Supplemental Links

- Take a quiz on properties of matter at www.chem.tamu.edu/class/majors/tutorialnotefiles/intext.htm .

Points to Consider

A mixture is different from a pure substance because it does not have a uniform and definite composition. Pure substances include both elements and compounds.

- Can mixtures be further classified into more than one type?
- How are compounds different than elements?

3.2 Classification of Matter

Lesson Objectives

- Define a mixture and understand why mixtures are not substances.
- Classify mixtures as homogeneous or heterogeneous.
- Describe several ways to separate mixtures.
- Distinguish between elements and compounds.

Lesson Vocabulary

- chemical change
- chemical formula
- chemical symbol
- compound
- distillation
- element
- filtration
- heterogeneous mixture
- homogeneous mixture
- mixture
- phase

Check Your Understanding

Recalling Prior Knowledge

- What is a pure substance?
- What are some examples of physical changes, and what happens to the identity of a substance when it undergoes a physical change?

In the previous lesson, "Properties of Matter," we learned about the physical properties of pure substances and about the physical changes that matter can undergo. However, there are many examples of matter that cannot be classified as substances because they have a composition that is variable rather than fixed.

Mixtures

Ordinary table salt is called sodium chloride. It is considered a substance because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given sample of water. A **mixture** *is a physical blend of two or more components, each of which retains its own identity and properties*. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

An activity that makes it easy to understand the differences between elements, compounds and mixtures can be found at www.bbc.co.uk/schools/ks3bitesize/science/chemical_material_behaviour/compounds_mixtures/activity.shtml .

Homogeneous Mixtures

A **homogeneous mixture** *is a mixture in which the composition is uniform throughout the mixture*. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. It is often easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same. Air is a mixture because the different gases that make up air may be present in varying amounts. A troublesome aspect of modern society is that the burning of fossil fuels in large amounts has increased the amount of carbon dioxide in our atmosphere compared to the levels that were present hundreds of years ago before the industrial revolution. Carbon dioxide acts as a greenhouse gas, trapping heat in our atmosphere. The increase in carbon dioxide levels has led to a corresponding increase in average global temperatures (**Figure** 3.5).

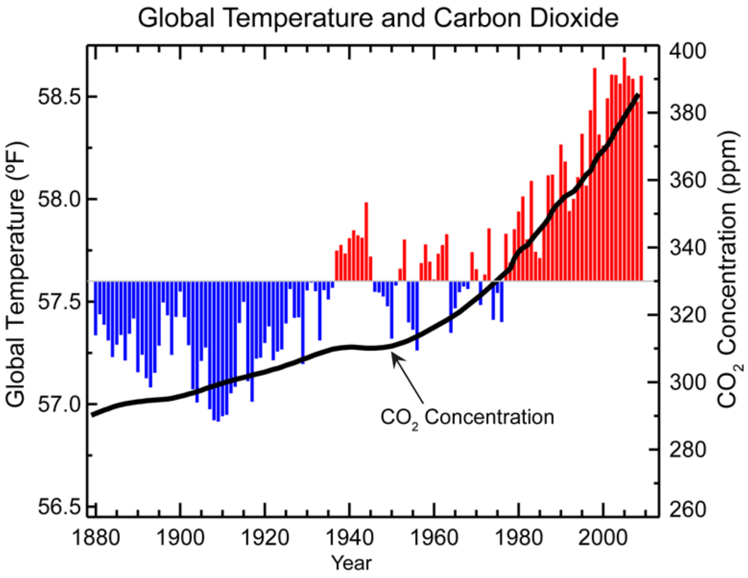

FIGURE 3.5

Carbon dioxide levels in Earth's atmosphere (measured in parts per million) have increased dramatically since 1900. The average global temperature has also seen an increase, as seen on the left-hand axis.

Heterogeneous Mixtures

A **heterogeneous mixture** *is a mixture in which the composition is not uniform throughout the mixture*. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup.

A **phase** *is any part of a sample that has a uniform composition and properties*. By definition, a pure substance or a homogeneous mixture consists of a single phase. A heterogeneous mixture consists of two or more phases. When oil and water are combined, they do not mix evenly (**Figure** 3.6). Instead, they form two separate layers, each of which is considered a phase.

FIGURE 3.6

Oil and water do not mix, instead forming two distinct layers called phases. The oil phase is less dense than the water phase, so the oil floats on top of the water.

Separation of Mixtures

A pizza is a heterogeneous mixture. If you have a pizza slice with green peppers on it and you don't like green peppers, you could easily pick them off before eating the slice. Other mixtures may not be as easy to separate. The oil and water mixture can be separated by carefully pouring off the oil layer in a process called decanting. The components of all mixtures can be separated by some physical means.

Filtration

Sand does not dissolve in water. In order to separate sand from water, the mixture can be poured through a filter. The tiny holes of a filter, called pores, allow the water to pass through while trapping the sand grains behind. **Filtration** *is a technique that separates solids from liquids in a heterogeneous mixture* (**Figure** 3.7).

Distillation

Distillation *is a method of separating a mixture based on differences in boiling points*. A distillation apparatus is shown in **Figure** 3.8.

For example, the impure liquid could be a mixture of alcohol and water. As the mixture is heated, the alcohol begins to boil before the water does because of its lower boiling point. The alcohol vapors rise up in the flask and into the long cylindrical tube, which is called a condenser. Cold water is continually pumped into the outside of the condenser, which causes the alcohol vapors to condense back into the liquid state. The purified alcohol then drips into the flask labeled "distilled liquid," while the water is left behind in the original flask.

Distillation is commonly used to purify water by separating it from the various dissolved minerals that are found in

FIGURE 3.7

This filtering apparatus uses suction to pull the water through a fine glass filter, leaving behind a yellow-colored solid.

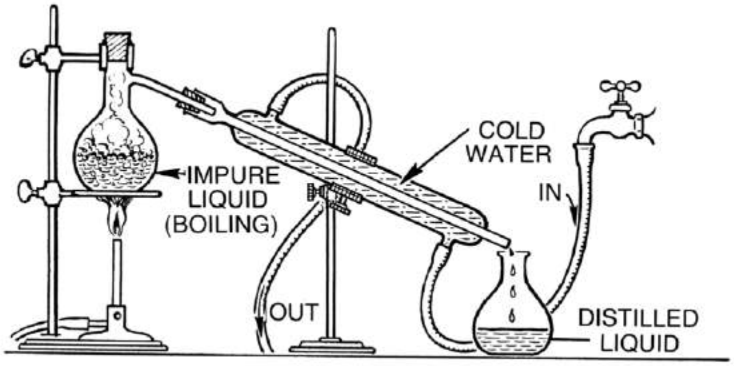

FIGURE 3.8

A small-scale laboratory distillation apparatus is used to separate substances based on differences in their boiling points.

natural water sources.

You can view the distillation of methanol to separate it from a dye at www.youtube.com/watch?v=74rOJ1GFbC0 (1:16).

MEDIA

Click image to the left for more content.

Pure Substances

As described in the lesson "Properties of Matter," pure substances have a definite and fixed composition. Pure substances can further be classified as either elements or compounds.

Elements

An **element** *is the simplest form of matter that has a unique set of properties*. Examples of well-known elements include oxygen, iron, and gold (**Figure** 3.9). Elements cannot be broken down into simpler substances. Additionally, one element cannot be chemically converted into a different element.

FIGURE 3.9

Chemical elements are the simplest of substances. (A) An oxygen tank like this is used by people with a need for breathing assistance. (B) A simple skillet can be made from cast iron. (C) Gold bars are formed and used for monetary purposes.

Compounds

A **compound** *is a substance containing two or more elements that have been chemically combined in a fixed proportion*. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen particles as carbon particles. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances (the elements carbon and hydrogen). The process of breaking down a compound into its elements is sometimes called decomposition.

Recall that the components of a mixture can be separated from one another by physical means. This is not true for a compound. Table salt is a compound consisting of equal parts of the elements sodium and chlorine. Salt cannot be separated into its two elements by filtering, distillation, or any other physical process. Salt and other compounds can only be decomposed into their elements by a chemical process. A **chemical change** *is a change that produces matter with a different composition*. Many compounds can be decomposed into their elements by heating. When sugar is heated, it decomposes to carbon and water. Water is also a compound, but it cannot be broken down into hydrogen and oxygen by heating. However, the passage of an electrical current through water will produce hydrogen and oxygen gases.

The properties of compounds are generally very different than the properties of the elements from which the compound is formed. Sodium is an extremely reactive soft metal that cannot be exposed to air or water. Chlorine is a deadly gas. The compound sodium chloride is a mostly unreactive white solid that is essential for all living things (**Figure** 3.10).

Distinguishing Substances and Mixtures

The flowchart in **Figure** 3.11 summarizes the types of matter discussed in this chapter.

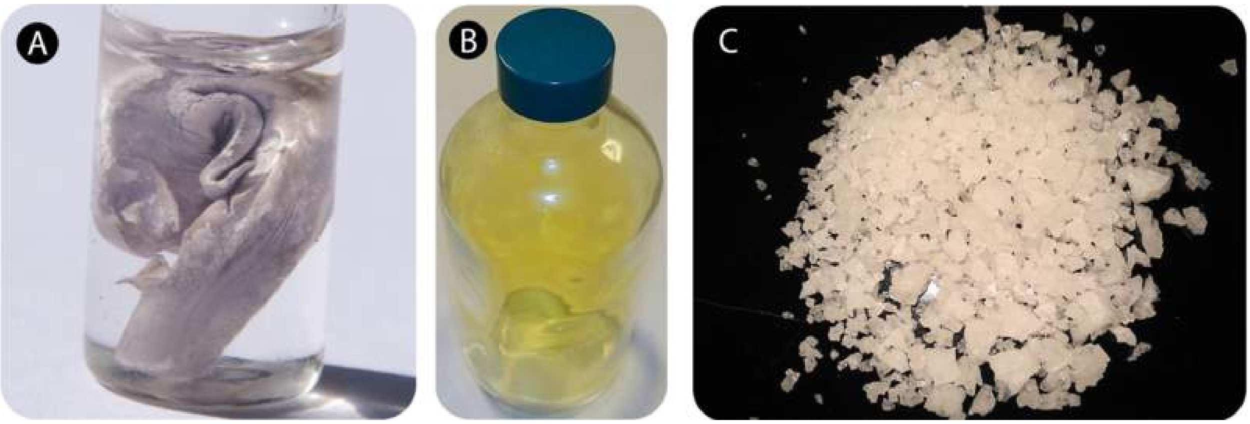

FIGURE 3.10

(A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals are a compound of sodium and chlorine.

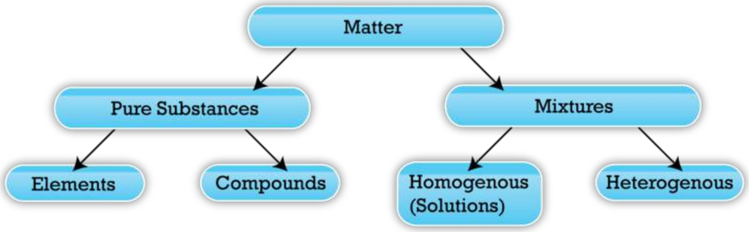

FIGURE 3.11

The flowchart describes the classification of matter.

Matter can be divided into two categories: substances and mixtures. Pure substances can be either elements or compounds. Mixtures can either be homogeneous or heterogeneous. Another term for a homogeneous mixture is a solution. Sometimes it can be difficult to tell the difference between a compound and a homogeneous mixture based solely on appearance. Consider gasoline. It is certainly homogeneous and seems a likely candidate for a compound, meaning that it would have a fixed composition. However, as you may know, gasoline can be purchased at a station in a variety of grades, such as regular or premium grade. Premium grades have a higher octane rating and cost more than regular grades. Gasoline is thus a blend of different compounds, so it is a mixture.

Chemical Symbols and Formulas

Chemists use symbols and formulas as an abbreviated way of writing chemical reactions and the elements and compounds involved in them. A **chemical symbol** *is a unique one- or two-letter designation of an element*. Some examples of chemical symbols are O for oxygen, Zn for zinc, and Fe for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names (**Table** 3.2).

TABLE 3.2: Symbols and Latin Names for Various Elements

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

As stated earlier, compounds are combinations of two or more elements. A **chemical formula** *is an expression that shows the elements in a compound and the relative proportions of those elements*. For example, water is composed of hydrogen and oxygen in a two to one ratio. The chemical formula for water is H_2O. Sulfuric acid is one of the most widely produced chemicals in the Unites States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is H_2SO_4.

Lesson Summary

- Mixtures are physical blends of two or more substances and can either be homogeneous or heterogeneous. Mixtures can be separated into their components by physical means such as filtration or distillation.
- Pure substances may be either elements or compounds. Elements are the simplest types of matter, while a compound is a chemical combination of two or more elements. A compound may only be decomposed into its elements by a chemical process.
- Chemical symbols and formulas are used to represent elements and compounds.

Lesson Review Questions

Reviewing Concepts

1. Why can't a mixture be classified as a substance?
2. Classify each of the following as a homogeneous or heterogeneous mixture.
    a. salad
    b. iced tea
    c. soil
    d. cooking oil
3. When is filtration used to separate a mixture? When is distillation used?
4. What is wrong with the following explanation of an experiment? "When it was heated, the sample of carbon decomposed."
5. Identify the following as either a chemical symbol or a chemical formula.
    a. CO
    b. Co
    c. W
    d. Be

   e. HI

Problems

6. Describe a way to separate a mixture of fine sand and table salt without having to directly sort tiny crystals using tweezers.
7. Name the elements found in each of the following compounds.
   a. ammonia (NH_3)
   b. saltpeter (KNO_3)
   c. glucose ($C_6H_{12}O_6$)
   d. millerite (NiS)
8. Identify each of the following items as an element, compound or mixture. If it is a mixture, further classify it as homogeneous or heterogeneous.
   a. gasoline
   b. calcium
   c. carbon dioxide
   d. blood

Further Reading / Supplemental Links

- Another method for separating a mixture is chromatography. You can see how bomb specialists use chromatography to identify the chemical components that were used in the making of the bomb. There is also a chromatography activity that you can do at home! Go to www.pbs.org/wgbh/nova/bombingactivity/chromatog.html .

Points to Consider

In a chemical reaction, one or more substances are changed into different substances.

- How are chemical changes different from physical changes?
- What are the visual clues that a chemical reaction is taking place?

3.3 Changes in Matter

Lesson Objectives

- Identify the chemical properties of a substance.
- Describe chemical changes and differentiate them from physical changes.
- Use various visual clues to identify whether a chemical reaction is taking place.

Lesson Vocabulary

- chemical property
- chemical reaction
- precipitate
- product
- reactant

Check Your Understanding

Recalling Prior Knowledge

- What is a chemical change?
- What are the three common states of matter?

As you have seen, physical changes are those that do not alter the identities of the substances involved. In contrast, chemical changes produce substances with different compositions. In this lesson, you will learn how to recognize chemical changes.

Chemical Changes

When exposed to air, an object made of iron will eventually begin to rust (**Figure** 3.12).

As the rust forms on the surface of the iron, it flakes off to expose more iron, which will continue to rust. Rust is clearly a substance that is different from iron. Rusting is an example of a chemical change.

A **chemical property** *describes the ability of a substance to undergo a specific chemical change*. A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust. A more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are *burn, rot, explode, decompose, and ferment*. Chemical properties are very useful as a way of identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

FIGURE 3.12
Rust (iron oxide) forms on an unprotected iron surface.

A chemical change is also called a chemical reaction. A **chemical reaction** *is a process that occurs when one or more substances are changed into one or more new substances*. Zinc (Zn) is a silver-gray element that can be ground into a powder. If zinc is mixed at room temperature with powdered sulfur (S), a bright yellow element, the result will simply be a mixture of zinc and sulfur. No chemical reaction occurs. However, if energy is provided to the mixture in the form of heat, the zinc will chemically react with the sulfur to form the compound zinc sulfide (ZnS). **Figure** 3.13 shows the substances involved in this reaction.

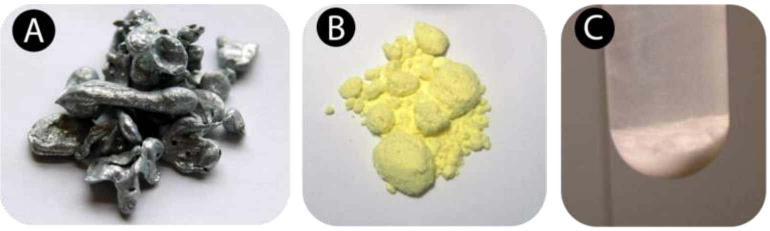

FIGURE 3.13
Zinc (A) and sulfur (B) are two elements that undergo a chemical reaction when heated to from the compound zinc sulfide (C).

The reaction between zinc and sulfur can be depicted in something called a chemical equation. In words, we could write the reaction as:

zinc + sulfur → zinc sulfide

A more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved:

$$\mathrm{Zn} + \mathrm{S} \rightarrow \mathrm{ZnS}$$

The substance(s) to the left of the arrow in a chemical equation are called reactants. A **reactant** *is a substance that is present at the start of a chemical reaction*. The substance(s) to the right of the arrow are called products. A **product** *is a substance that is present at the end of a chemical reaction*. In the equation above, zinc and sulfur are the reactants that chemically combine to form zinc sulfide as a product.

Recognizing Chemical Reactions

How can you tell if a chemical reaction is taking place? Certain visual clues indicate that a chemical reaction is likely (but not necessarily) occurring, including the following examples:

1. A change of color occurs during the reaction.
2. A gas is produced during the reaction.
3. A solid product called a precipitate is produced in the reaction.
4. A visible transfer of energy occurs in the form of light as a result of the reaction.

Mercury(II) oxide is a red solid. When it is heated to a temperature above 500°C, it easily decomposes into mercury and oxygen gas. The red color of the reactant, mercury oxide, is gradually replaced by the silver color of the product, mercury. The color change is one sign that this reaction is occurring. Watch this decomposition take place at http://www.youtube.com/watch?v=_Y1alDuXm6A (1:12).

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When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced (**Figure** 3.14). The production of a gas is also an indication that a chemical reaction may be occurring.

FIGURE 3.14

Zinc reacts with hydrochloric acid to produce bubbles of hydrogen gas.

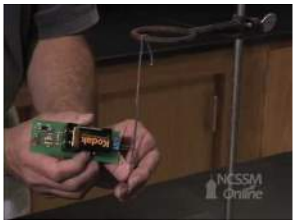
When a colorless solution of lead(II) nitrate is added to a colorless solution of potassium iodide, a yellow solid called a precipitate is instantly produced (**Figure** 3.15). A **precipitate** *is a solid product that forms from a reaction* *and settles out of a liquid mixture*. The formation of a precipitate may also indicate the occurrence of a chemical reaction.

FIGURE 3.15

A yellow precipitate of solid lead(II) iodide forms immediately when solutions of lead(II) nitrate and potassium iodide are mixed.

All chemical changes involve a transfer of energy. When zinc reacts with hydrochloric acid, the test tube becomes very warm as energy is released during the reaction. Some other reactions absorb energy. While energy changes are a potential sign of a chemical reaction, care must be taken to ensure that a chemical reaction is indeed taking place. Physical changes may also involve a transfer of energy. A solid absorbs energy when it melts, and the condensation of a gas releases energy. The only way to be certain that a chemical reaction has occurred is to test the composition of the substances after the change has taken place to see if they are different from the starting substances.

Combustion is a chemical change. Watch as magnesium reacts with oxygen to form magnesium oxide. View this lab-scale fireworks display at www.youtube.com/watch?v=EZ3JT2nWfMA (1:31).



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Lesson Summary

- The composition of matter changes during a chemical reaction. The ability of a substance to react with other substances is reflected in its chemical properties.
- In a chemical reaction, the reacting substances are called reactants, and the substances that are produced by the reaction are called products. A chemical equation shows reactants becoming products.
- Four clues to a possible chemical reaction include a color change, the production of a gas, the formation of a precipitate, and an observable transfer of energy.

Lesson Review Questions

Reviewing Concepts

1. What is the main difference between a physical change and a chemical change?
2. How must the chemical properties of a given substance be observed?
3. Give an example of a situation where a color change occurs, but not because of a chemical reaction.

Problems

4. Classify the following as physical properties or chemical properties of the element nickel.
    a. It is gray in color.
    b. It melts at 1455°C.
    c. It reacts with sulfuric acid.
    d. It is a conductor of electricity.
    e. It forms a compound with oxygen.
5. Combustion, also called burning, is a reaction with oxygen gas. When propane is combusted, carbon dioxide and water are formed. Identify the reactants and the products of this combustion reaction.
6. Classify each of the following as a chemical change or a physical change.
    a. Sugar dissolves in water.
    b. A peach rots.
    c. Icicles melt in the warm sunlight.
    d. A baking cake rises in the oven.
    e. A leaf changes its color in the fall.
    f. Food coloring is added to a glass of water.
7. Describe clues that you might observe during the following situations that would indicate that a chemical reaction is likely occurring.
    a. A log burns in the fireplace.
    b. A mixture of baking soda and vinegar fizzes.
    c. A silver fork tarnishes.
    d. A plumbing pipe becomes clogged with lime.

Points to Consider

A chemist must do more than simply make observations about chemical changes that are occurring in his or her experiments. He or she must also take measurements in order to fully understand what is taking place during chemical reactions.

- What types of measurements are routinely made by chemists?
- How does a chemist indicate the degree of confidence in a measurement?
- How are different numerical measurements manipulated when performing calculations?

3.4 References

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CHAPTER 4

The Structure of Atoms, Isotopes and Ions

Chapter Outline

NC Standards addressed in this chapter

Chm.1.1.1 Analyze the structure of atoms, isotopes, and ions.

- Characterize protons, neutrons, electrons by location, relative charge, relative mass (p=1, n=1, e=1/2000).
- Use symbols: A= mass number, Z=atomic number • Use notation for writing isotope symbols:235U or U-235 92
- Identify isotope using mass number and atomic number and relate to number of protons, neutrons and electrons.
- Differentiate average atomic mass of an element from the actual isotopic mass and mass number of specific isotopes. (Use example calculations to determine average atomic mass of atoms from relative abundance and actual isotopic mass to develop understanding).

4.1 Atoms

Lesson Objectives

- Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Describe John Dalton's atomic theory.

Lesson Vocabulary

- atom
- law of conservation of mass
- law of definite proportions
- law of multiple proportions

Early Atomic Models

Democritus was a Greek philosopher who lived from 460 B.C. until 370 B.C. He was among the first known individuals to suggest the idea of the atom as the basic unit of matter. The word atom comes from the Greek word *atomos*, meaning "indivisible." Democritus (see **Figure** 4.1) was indeed ahead of his time, but his ideas were not useful at that time in describing chemical behavior because there was no experimental evidence to support it. His approach was a philosophical one rather than a truly scientific one. Many centuries would pass before the notion of atoms was merged with modern scientific experimentation and thought.

FIGURE 4.1

Democritus believed that all matter consisted of tiny, indestructible particles called atoms.

Beginnings of Atomic Theory

By the late 1700s, chemists had accepted the definition of an element as a substance that cannot be broken down into simpler substances by ordinary chemical means. It was also clear that elements combine with one another to form more complex substances called compounds. The chemical and physical properties of these compounds are different than the properties of the elements from which they were formed. There was some disagreement, however, about whether elements always combined in exactly the same ratio when forming a particular compound.

In the 1790s, a greater emphasis began to be placed on the quantitative analysis of chemical reactions. Accurate and reproducible measurements of the masses of reacting elements and the compounds that they form led to the formulation of several basic laws. One of these is called the **law of conservation of mass**, *which states that during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants*. In other words, mass cannot be created or destroyed during a chemical reaction, but must always be conserved.

The law of conservation of mass is demonstrated in this video: http://www.youtube.com/watch?v=J5hM1DxaPLw (1:05).

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Click image to the left for more content.

The discovery that mass is always conserved in chemical reactions was soon followed by the **law of definite proportions**, *which states that a given chemical compound always contains the same elements in the exact same proportions by mass*. As an example, any sample of pure water contains 11.19% hydrogen and 88.81% oxygen by mass. It does not matter where the sample of water came from or how it was prepared. Its composition, like that of every other compound, is fixed (**Figure** 4.2).

FIGURE 4.2

Regardless of its source or its form (solid, liquid, or gas), water always has the exact same elemental composition: 11.19% of its mass is hydrogen and 88.81% of its mass is oxygen.

Experiments also began to demonstrate that the same pairs of certain elements could combine to form more than one compound. Consider the elements carbon and oxygen. Combined in one way, they form the familiar compound called carbon dioxide. In every sample of carbon dioxide, there is 32.0 g of oxygen present for every 12.0 g of carbon. By dividing 32.0 by 12.0, this simplifies to an oxygen/carbon mass ratio of 2.66 to 1. Another compound that forms from the combination of carbon and oxygen is called carbon monoxide. Every sample of carbon monoxide contains 16.0 g of oxygen for every 12.0 g of carbon, which simplifies to an oxygen/carbon mass ratio of 1.33 to 1. In other words, a given mass of carbon needs to combine with exactly twice as much oxygen to make carbon dioxide as it would to produce carbon monoxide. **Figure** 4.3 illustrates the **law of multiple proportions**. *Whenever the same two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.*

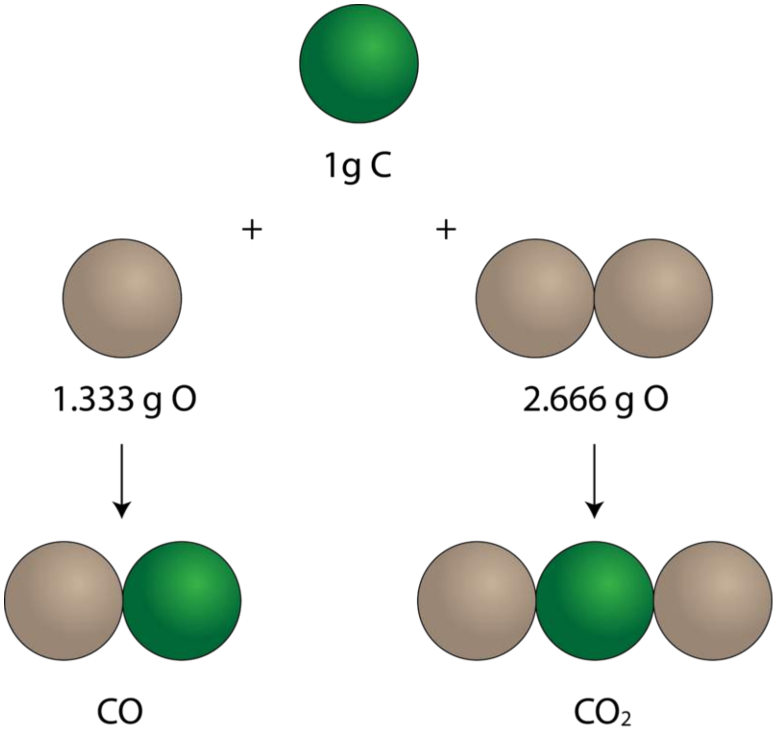


FIGURE 4.3

Carbon monoxide, on the left, contains 1.333 g of oxygen for every 1 g of carbon. Carbon dioxide, on the right, contains 2.666 g of oxygen for every gram of carbon. The ratio of oxygen in these two compounds is 1:2, which is a ratio of small whole numbers.

Sample Problem 4.1: Calculating Mass Ratios

Copper reacts with chlorine to form two compounds. Compound A contains 4.08 g of copper for every 2.28 g of chlorine. Compound B contains 7.53 g of copper for every 8.40 g of chlorine. What is the lowest whole number mass ratio of copper that combines with a given mass of chlorine?

Step 1: List the known quantities and plan the problem.

Known

- Compound A = 4.08 g Cu and 2.28 g Cl
- Compound B = 7.53 g Cu and 8.40 g Cl

Apply the law of multiple proportions to the two compounds. For each compound, find the grams of copper that combine with 1.00 g of chlorine by dividing the mass of copper by the mass of chlorine. Then, find the ratio of the masses of copper in the two compounds by dividing the larger value by the smaller value.

Step 2: Calculate.

$$\text{Compound A} \qquad \frac{4.08 \text{ g Cu}}{2.28 \text{ g Cl}} = \frac{1.79 \text{ g Cu}}{1.00 \text{ g Cl}}$$

$$\text{Compound B} \qquad \frac{7.53 \text{ g Cu}}{8.40 \text{ g Cl}} = \frac{0.896 \text{ g Cu}}{1.00 \text{ g Cl}}$$

Compare the masses of copper per gram of chlorine in the two samples.

$$\frac{1.79 \text{ g Cu (in compound A)}}{0.896 \text{ g Cu (in compound B)}} = \frac{2.00}{1} = 2:1$$

The mass ratio of copper per gram of chlorine in the two compounds is 2:1.

Step 3: Think about your result.

The ratio is a small whole-number ratio. For a given mass of chlorine, compound A contains twice the mass of copper as does compound B.

Dalton's Atomic Theory

In 1808, an English chemist and schoolteacher named John Dalton (1766-1844) formulated an atomic theory based on the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. His theory can be summarized in the following statements, illustrated in **Figure** 4.4.

1. All matter is composed of extremely small particles called atoms.
2. Atoms of the same element are identical in terms of size, mass, and other properties. Atoms of one element are different from the atoms of any other element.
3. Atoms of different elements can chemically combine with one another in simple whole-number ratios to form chemical compounds.
4. Chemical reactions can be described as the separation, combination, or rearrangement of atoms. Atoms of one element cannot be changed into atoms of a different element as a result of a chemical reaction.

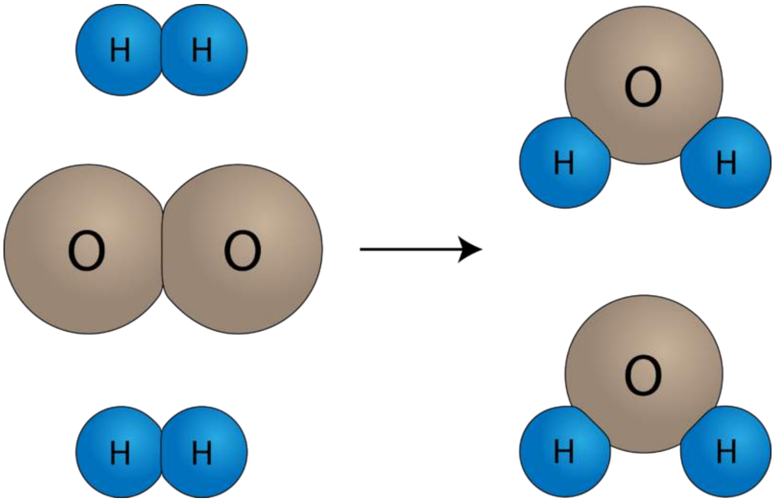


FIGURE 4.4

Dalton's atomic theory states that chemical reactions are due to rearrangements of atoms. On the left, hydrogen molecules are shown as two H atoms, while oxygen molecules are shown as two O atoms. The chemical reaction forms water as the atoms rearrange. There are identical numbers of hydrogen and oxygen atoms before and after the reaction.

Sizes of Atoms

The graphite in your pencil is composed of the element carbon. Imagine taking a small piece of carbon and grinding it until it is a fine dust. Each speck of carbon would still have all of the physical and chemical properties of carbon. Now imagine that you could somehow keep dividing the speck of carbon into smaller and smaller pieces. Eventually, you would reach a point where your carbon sample is as small as it could possibly be. This final particle is called an **atom**, *which is defined as the smallest particle of an element that retains the properties of that element*.

Atoms, as you probably know, are extremely small. In fact, the graphite in an ordinary pencil contains about 5×10^{20} atoms of carbon. This is an almost incomprehensibly large number. The population of the entire Earth is about 7×10^9 people, meaning that there are about 7×10^{10} times as many carbon atoms in your pencil as there are people on the Earth! For this to be true, atoms must be extremely small. Can we see atoms? It's not easy, but a modern instrument called a scanning tunneling microscope allows scientists to visualize the atom, as shown in **Figure** 4.5.

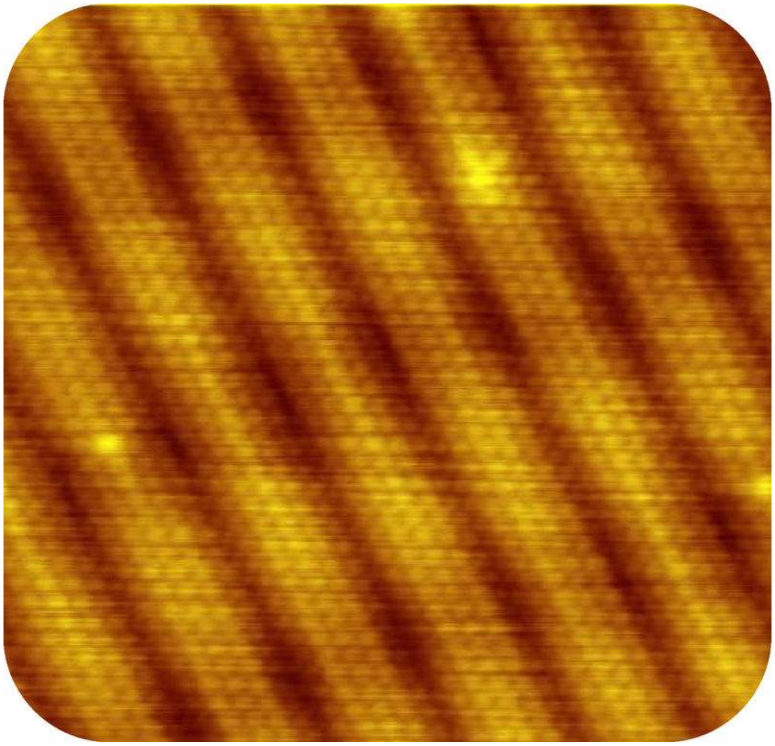


FIGURE 4.5

Images of individual gold atoms can be seen on the surface of a smooth sheet of gold metal using scanning tunneling microscopy.

Lesson Summary

- The Greek philosopher Democritus believed that matter is composed of indivisible and indestructible building blocks, which he called atoms.
- John Dalton transformed the ideas of Democritus into a scientific atomic theory, which began to explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Matter cannot be created or destroyed in chemical reactions.
- Elements react to form compounds in fixed proportions by mass.
- Elements combine with one another in simple whole-number ratios.
- An atom is the smallest unit of an element that retains the properties of that element.

Lesson Review Questions

Recall

1. Describe the contribution of Democritus to modern atomic theory.
2. What happens to the total mass of a system during a chemical reaction?
3. What does "fixed composition" mean with regard to chemical compounds?
4. What law explains the relationships between different chemical compounds that contain only the elements carbon and oxygen?
5. What are the four essential points of Dalton's atomic theory?

Apply Concepts

6. Alchemy is an ancient tradition that predated the modern science of chemistry. One goal of the alchemist was to change base metals such as iron or lead into valuable metals such as gold or silver. Explain, using Dalton's theory, why alchemists were unable to do this.

Think Critically

7. If an atom of element A has a mass of 2 units, while an atom of element B has a mass of 5 units, what would be the ratio of element A to element B in a compound that had a mass of 11 units?
8. Sulfur reacts with oxygen to form two compounds. Compound A consists of 1.89 g of sulfur for every 1.89 g of oxygen. Compound B consists of 3.72 g of sulfur for every 5.57 g of oxygen. What is the lowest whole number mass ratio of sulfur that combines with a given mass of oxygen?
9. Hydrocarbons are a class of organic compounds that contain only carbon and hydrogen. Methane, an important hydrocarbon fuel, has a carbon/hydrogen mass ratio of 3:1. An 80-g sample of a hydrocarbon compound is analyzed and found to contain 64 g of carbon. Is the hydrocarbon sample methane? Explain.
10. The mass of 6.02×10^{23} atoms of iron is 55.85 g. What is the mass of one atom of iron?

Further Reading / Supplemental Links

- For more information about Dalton's contributions, go to the video Early Ideas About Atoms at http://www.teachertube.com/viewVideo.php?title=John_Dalton_s_Atomic_Theory&video_id=228303 .

4.2 Democritus' Idea of the Atom

- Review how and when Democritus arrived at his idea of the atom.
- Describe the atom as Democritus understood it.
- State how Democritus explained the diversity of matter.
- Explain why Democritus' idea was ignored until about 1800.

The man shown above has been called the "laughing philosopher" because of his cheerful disposition. He certainly looks cheerful in this picture. Why is a philosopher featured in a science text? He made an amazing contribution to science, although it was ridiculed by others and then ignored for more than 2000 years. His name was Democritus, and he introduced the idea of the atom as the basic building block of all matter.

Cutting the Cheese

Democritus lived in Greece from about 460 to 370 B.C.E. Like many other ancient Greek philosophers, he spent a lot of time wondering about the natural world. Democritus wondered, for example, what would happen if you cut a chunk of matter—such as a piece of cheese like the one in the **Figure** 4.6—into smaller and smaller pieces. He thought that a point would be reached at which the cheese could not be cut into still smaller pieces. He called these pieces *atomos*, which means "uncuttable" in Greek. This is where the modern term *atom* comes from.

FIGURE 4.6

How many times could you cut this piece of cheese in half? How small would the smallest pieces be?

Just a Guess

Democritus' idea of the atom has been called "the best guess in antiquity." That's because it was correct in many ways, yet it was based on pure speculation. It really was just a guess. Here's what Democritus thought about the atom:

- All matter consists of atoms, which cannot be further subdivided into smaller particles.
- Atoms are extremely small—too small to see.
- Atoms are solid particles that are indestructible.
- Atoms are separated from one another by emptiness, or "void."

Q: How are Democritus' ideas about atoms similar to modern ideas about atoms?

A: Modern ideas agree that all matter is made up of extremely small building blocks called atoms.

Q: How are Democritus' ideas different from modern ideas?

A: Although atoms are extremely small, it is now possible to see them with very powerful microscopes. Atoms also aren't the solid, uncuttable particles Democritus thought. Instead, they consist of several kinds of smaller, simpler particles as well as a lot of empty space. In addition, atoms aren't really indestructible because they can be changed to other forms of matter or energy.

Why Matter Varies

Democritus thought that different kinds of matter vary because of the size, shape, and arrangement of their atoms. For example, he suggested that sweet substances are made of smooth atoms and bitter substances are made of sharp atoms. He speculated that atoms of liquids are slippery, which allows them to slide over each other and liquids to flow. Atoms of solids, in contrast, stick together, so they cannot move apart. Differences in the weight of matter, he argued, could be explained by the closeness of atoms. Atoms of lighter matter, he thought, were more spread out and separated by more empty space.

Q: Democritus thought that different kinds of atoms make up different types of matter. How is this similar to modern ideas about atoms?

A: The modern view is that atoms of different elements differ in their numbers of protons and electrons and this gives them different physical and chemical properties.

That's Ridiculous!

Democritus was an important philosopher, but he was less influential than another Greek philosopher named Aristotle, who lived about 100 years after Democritus. Aristotle rejected Democritus' idea of the atom. In fact, Aristotle thought the idea was ridiculous. Unfortunately, Aristotle's opinion was accepted for more than 2000 years, and Democritus' idea was more or less forgotten. However, the idea of the atom was revived around 1800 by the English scientist John Dalton. Dalton developed an entire theory about the atom, much of which is still accepted today. He based his theory on experimental evidence, not on lucky guesses.

Summary

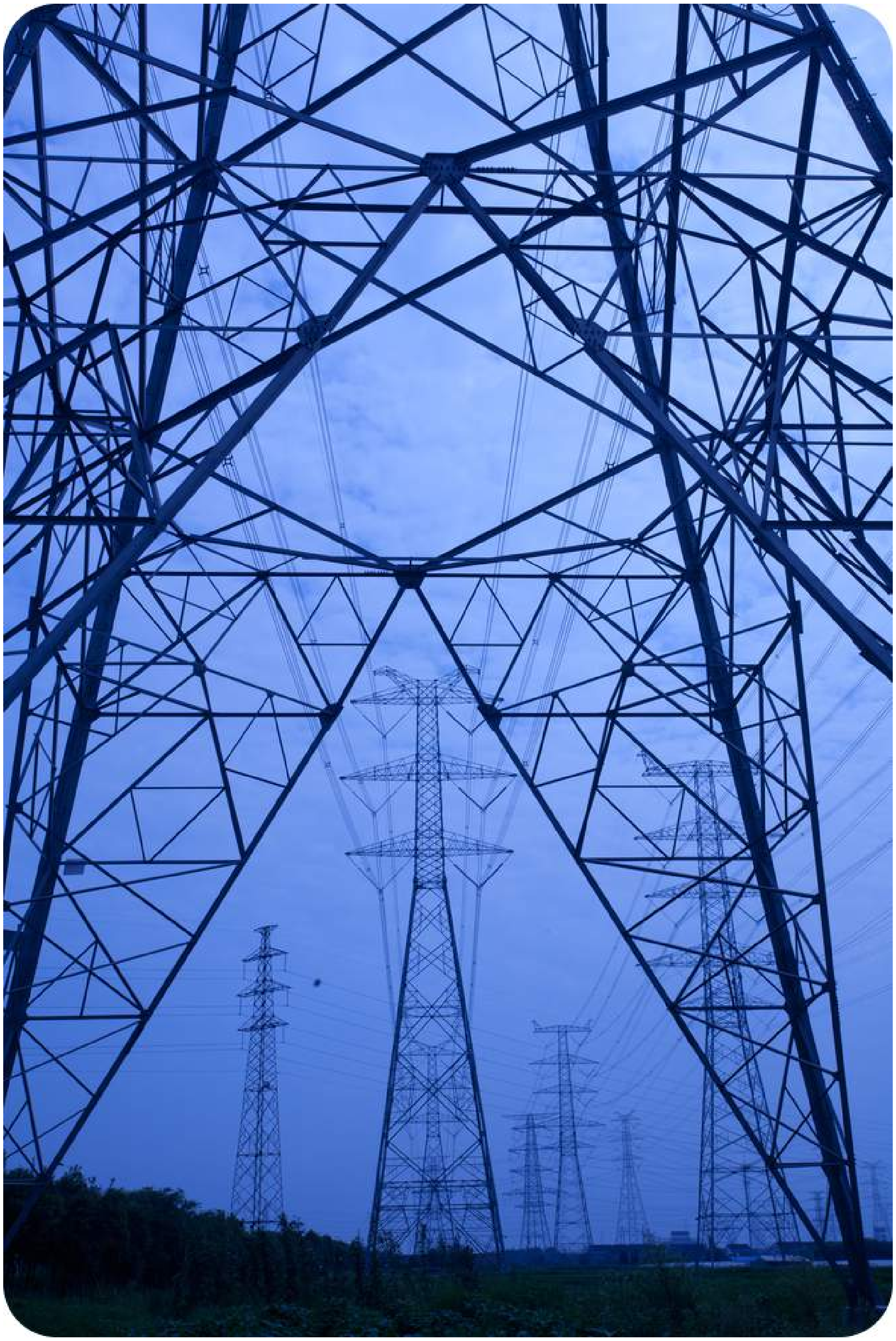
- Around 400 B.C.E., the Greek philosopher Democritus introduced the idea of the atom as the basic building block matter.
- Democritus thought that atoms are tiny, uncuttable, solid particles that are surrounded by empty space and constantly moving at random.
- Democritus surmised that different kinds of matter consist of different types or arrangements of atoms.

Review

1. Who was Democritus?
2. How did Democritus reason for the existence of atoms?
3. How did Democritus describe atoms?
4. Discuss how well Democritus' ideas about atoms have withstood the test of time.
5. Iron and lead are both metals, but iron is much harder than lead. How do you think Democritus might have explained this difference?

4.3 Thomson's Atomic Model

- Explain how J. J. Thomson discovered the electron.
- Describe Thomson's plum pudding model of the atom.



You probably know that the wires strung between these high towers carry electricity. But do you know what electricity is? It actually consists of a constant stream of tiny particles called electrons. Electrons are negatively charged fundamental particles inside atoms. Atoms were discovered around 1800, but almost 100 years went by before electrons were discovered.

Thomson Discovers Electrons

John Dalton discovered atoms in 1804. He thought they were the smallest particles of matter, which could not be broken down into smaller particles. He envisioned them as solid, hard spheres. It wasn't until 1897 that a scientist named Joseph John (J. J.) Thomson discovered that there are smaller particles within the atom. Thomson was born in England and studied at Cambridge University, where he later became a professor. In 1906, he won the Nobel Prize in physics for his research on how gases conduct electricity. This research also led to his discovery of the electron. You can see a picture of Thomson 4.7. You can take a museum tour of his discovery at this online exhibit: http://www.aip.org/history/electron/ .

FIGURE 4.7

Thomson's Experiments

In his research, Thomson passed current through a cathode ray tube, similar to the one seen in the **Figure** 4.8. A cathode ray tube is a glass tube from which virtually all of the air has been removed. It contains a piece of metal called an electrode at each end. One electrode is negatively charged and known as a cathode. The other electrode is positively charged and known as an anode. When high-voltage electric current is applied to the end plates, a cathode ray travels from the cathode to the anode.

What is a cathode ray? That's what Thomson wanted to know. Is it just a ray of energy that travels in waves like a ray of light? That was one popular hypothesis at the time. Or was a cathode ray a stream of moving particles? That was the other popular hypothesis. Thomson tested these ideas by placing negative and positive plates along the sides of the cathode ray tube to see how the cathode ray would be affected. The cathode ray appeared to be repelled by the negative plate and attracted by the positive plate. This meant that the ray was negative in charge and that is must consist of particles that have mass. He called the particles "corpuscles," but they were later renamed electrons. For a video demonstration of Thomson's experiment, go to this URL: http://ericsaltchemistry.blogspot.com/2010/10/jj-thomsons-experiments-with-cathode.html .

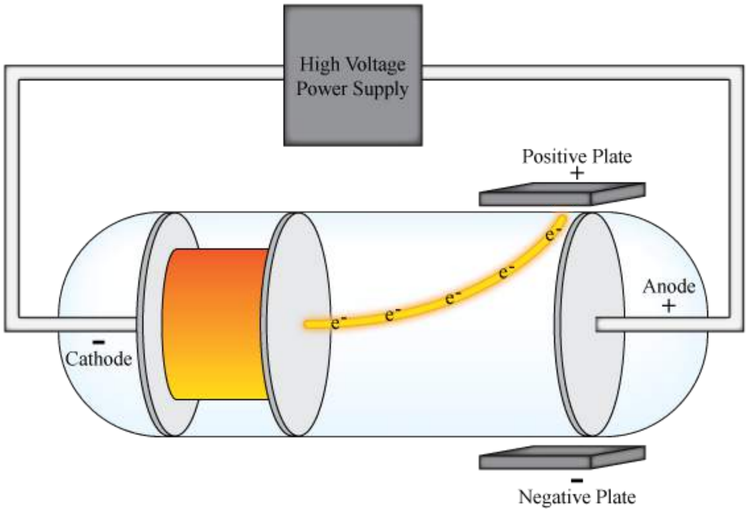


FIGURE 4.8

Thomson also measured the mass of the particles he had identified. He did this by determining how much the cathode rays were bent when he varied the voltage. He found that the mass of the particles was 2000 times smaller than the mass of the smallest atom, the hydrogen atom. In short, Thomson had discovered the existence of particles smaller than atoms. This disproved Dalton's claim that atoms are the smallest particles of matter. From his discovery, Thomson also inferred that electrons are fundamental particles within atoms.

Q: Atoms are neutral in electric charge. How can they be neutral if they contain negatively charged electrons?

A: Atoms also contain positively charged particles that cancel out the negative charge of the electrons. However, these positive particles weren't discovered until a couple of decades after Thomson discovered electrons.

The Plum Pudding Model

Thomson also knew that atoms are neutral in electric charge, so he asked the same question: How can atoms contain negative particles and still be neutral? He hypothesized that the rest of the atom must be positively charged in order to cancel out the negative charge of the electrons. He envisioned the atom as being similar to a plum pudding, like the one pictured in the **Figure** 4.9—mostly positive in charge (the pudding) with negative electrons (the plums) scattered through it.

FIGURE 4.9

Q: How is our modern understanding of atomic structure different from Thomson's plum pudding model?

A: Today we know that all of the positive charge in an atom is concentrated in a tiny central area called the nucleus, with the electrons swirling through empty space around it, as in the **Figure** below. The nucleus was discovered just a few years after Thomson discovered the electron, so the plum pudding model was soon rejected.

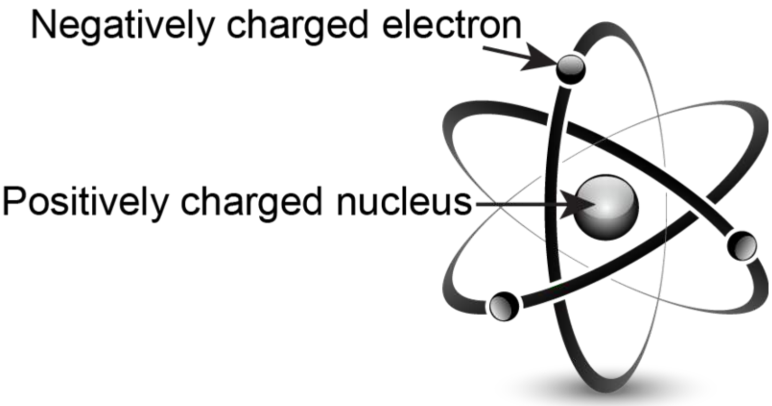


FIGURE 4.10

Summary

- In 1897, J. J. Thomson discovered the first subatomic particle, the electron, while researching cathode rays.
- To explain the neutrality of atoms, Thomson proposed a model of the atom in which negative electrons are scattered throughout a sphere of positive charge. He called his atom the plum pudding model.

Practice

Watch this detailed presentation about J. J. Thomson's discovery of the electron, and then answer the question below.

http://www.youtube.com/watch?v=SZSEnCydeb0 (12:30)

MEDIA

Click image to the left for more content.

1. Thomson not only discovered that a cathode ray consists of flowing negatively charged particles that are smaller than atoms. He also made the logical leap that these particles help make up atoms. What reasoning did Thomson use to make this inference?

Review

1. Who was J. J. Thomson?
2. Explain how Thomson discovered negatively charged particles smaller than atoms.
3. Thomson compared his idea of atomic structure to a plum pudding. Invent an original analogy for Thomson's plum pudding model of the atom.
4. Why was Thomson's model soon rejected?

4.4 Rutherford's Atomic Model

- Explain how Ernest Rutherford discovered the nucleus of the atom.
- Describe Rutherford's planetary model of the atom.

Thom is shooting baskets. He's trying to hit the backboard so the ball will bounce off it and into the basket. If only the backboard was bigger! It would be a lot easier to hit. If the ball misses the backboard, it will just keep going and Thom will have to run after it. Believe it or not, the research that led to the discovery of the nucleus of the atom was a little like shooting baskets.

Narrowing Down the Nucleus

In 1804, almost a century before the nucleus was discovered, the English scientist John Dalton provided evidence for the existence of the atom. Dalton thought that atoms were the smallest particles of matter, which couldn't be divided into smaller particles. He modeled atoms with solid wooden balls. In 1897, another English scientist, named J. J. Thomson, discovered the electron. It was first subatomic particle to be identified. Because atoms are neutral in electric charge, Thomson assumed that atoms must also contain areas of positive charge to cancel out the negatively charged electrons. He thought that an atom was like a plum pudding, consisting mostly of positively charged matter with negative electrons scattered through it.

The nucleus of the atom was discovered next. It was discovered in 1911 by a scientist from New Zealand named Ernest Rutherford, who is pictured in **Figure** 4.11. Through his clever research, Rutherford showed that the positive charge of an atom is confined to a tiny massive region at the center of the atom, rather than being spread evenly throughout the "pudding" of the atom as Thomson had suggested. You can watch a video about Rutherford and his discovery at this URL:

http://www.youtube.com/watch?v=wzALbzTdnc8 (3:28)

MEDIA

Click image to the left for more content.

Go for the Gold!

The way Rutherford discovered the atomic nucleus is a good example of the role of creativity in science. His quest actually began in 1899 when he discovered that some elements give off positively charged particles that can penetrate just about anything. He called these particles alpha (α) particles (we now know they were helium nuclei). Like all good scientists, Rutherford was curious. He wondered how he could use alpha particles to learn about the structure of the atom. He decided to aim a beam of alpha particles at a sheet of very thin gold foil. He chose gold because it can be pounded into sheets that are only 0.00004 cm thick. Surrounding the sheet of gold foil, he placed a screen that glowed when alpha particles struck it. It would be used to detect the alpha particles after they passed through the foil. A small slit in the screen allowed the beam of alpha particles to reach the foil from the particle emitter. You can see the setup for Rutherford's experiment in the **Figure** 4.12.

Q: What would you expect to happen when the alpha particles strike the gold foil?

A: The alpha particles would penetrate the gold foil. Alpha particles are positive, so they might be repelled by any areas of positive charge inside the gold atoms.

Assuming a plum pudding model of the atom, Rutherford predicted that the areas of positive charge in the gold atoms would deflect, or bend, the path of all the alpha particles as they passed through. You can see what really happened in the diagram above. Most of the alpha particles passed straight through the gold foil as though it wasn't there. The particles seemed to be passing through empty space. Only a few of the alpha particles were deflected from their straight path, as Rutherford had predicted. Surprisingly, a tiny percentage of the particles bounced back from the foil like a basketball bouncing off a backboard! You can see an animation of this experiment at the following URL: http://www.clickandlearn.org/gr9_sci/atoms/modelsoftheatom.html .

Q: What can you infer from these observations?

A: You can infer that most of the alpha particles were not repelled by any positive charge, whereas a few were repelled by a strong positive charge.

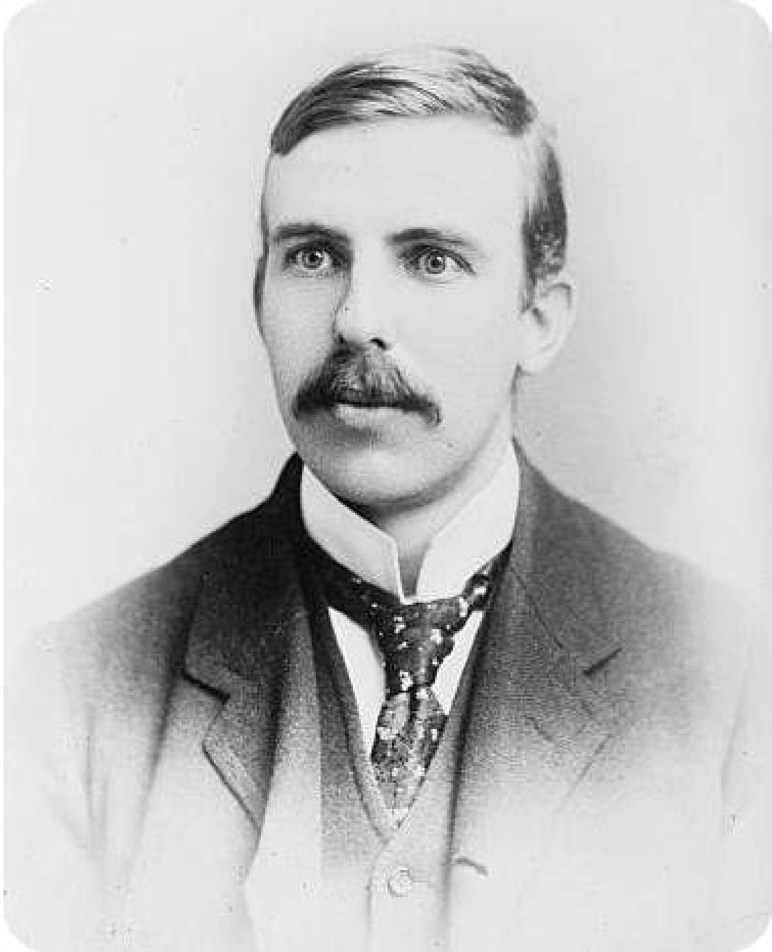


FIGURE 4.11

The Nucleus Takes Center Stage

Rutherford made the same inferences. He concluded that all of the positive charge and virtually all of the mass of an atom are concentrated in one tiny area and the rest of the atom is mostly empty space. Rutherford called the area of concentrated positive charge the nucleus. He predicted—and soon discovered—that the nucleus contains positively charged particles, which he named protons. Rutherford also predicted the existence of neutral nuclear particles called neutrons, but he failed to find them. However, his student James Chadwick discovered them several years later. You can learn how at this URL: http://www.light-science.com/chadwick.html .

The Planetary Model

Rutherford's discoveries meant that Thomson's plum pudding model was incorrect. Positive charge is not spread evenly throughout an atom. Instead, it is all concentrated in the tiny nucleus. The rest of the atom is empty space except for the electrons scattered through it. In Rutherford's model of the atom, which is shown in the **Figure** 4.13, the electrons move around the massive nucleus like planets orbiting the sun. That's why his model is called the planetary model. Rutherford didn't know exactly where or how electrons orbit the nucleus. That research would be undertaken by later scientists, beginning with Niels Bohr in 1913. New and improved atomic models would also be developed. Nonetheless, Rutherford's model is still often used to represent the atom. You can see an animated version of the model at this URL: http://www.clickandlearn.org/gr9_sci/atoms/modelsoftheatom.html .

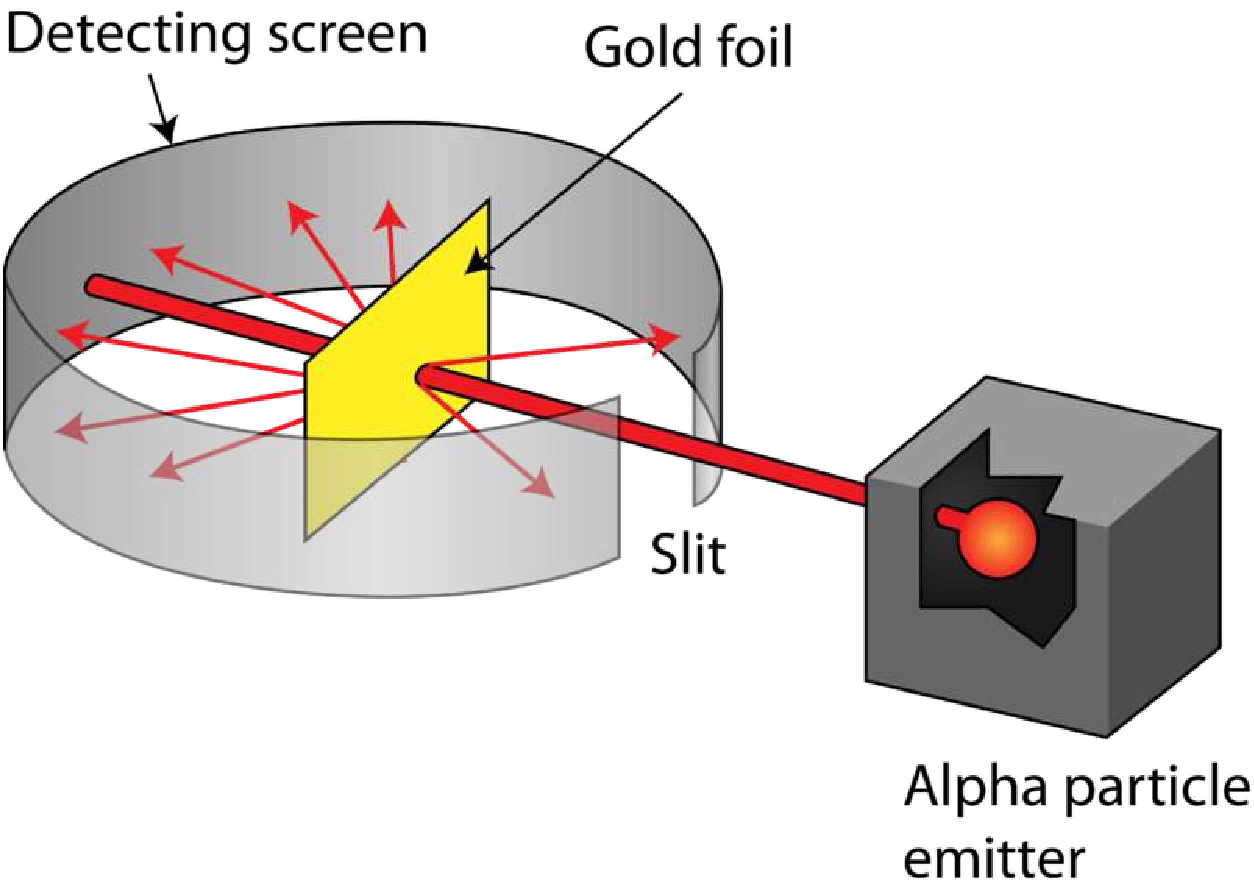


FIGURE 4.12

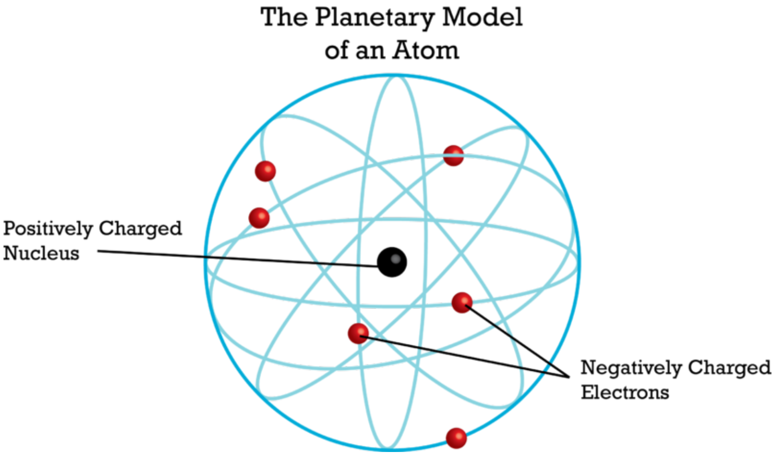


FIGURE 4.13

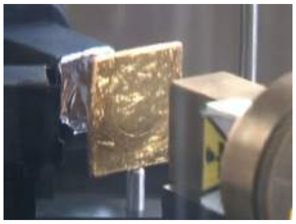
Summary

- Ernest Rutherford discovered the nucleus of the atom in 1910. He sent a beam of alpha particles toward gold foil and observed the way the particles were deflected by the gold atoms. From his results, he concluded that all of the positive charge and virtually all of the mass of an atom are concentrated in one tiny area, called the nucleus, and the rest of the atom is mostly empty space.
- In Rutherford's planetary model of the atom, the electrons move through empty space around the tiny positive nucleus like planets orbiting the sun.

Practice

Watch this video about Rutherford's gold foil experiment, and then answer the questions below.

http://www.youtube.com/watch?v=XBqHkraf8iE (4:06)



MEDIA

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1. How did Rutherford observe alpha particles in his experiment? In the modern version of Rutherford's experiment, which is shown in the video, how are alpha particles observed? Which way do you think is more accurate?
2. Based on the animation in the video, draw a sketch showing what happens to alpha particles as they pass through gold atoms.
3. How has Rutherford's gold foil experiment been adopted by modern researchers?

Review

1. How did Ernest Rutherford discover the nucleus of the atom?
2. Place Rutherford's discovery in the broader history of the atom. (*Hint:* See the timeline at the following URL for more information: http://www.clickandlearn.org/gr9_sci/atoms/modelsoftheatom.html .)
3. Describe how you could make a three-dimensional version of Rutherford's planetary model of the atom.

4.5 The Bohr Model of the Atom

Lesson Objectives

The student will:

- describe an electron cloud containing Bohr's energy levels.
- describe how the Bohr model of the atom explains the existence of atomic spectra.
- explain the limitations of the Bohr model and why it had to be replaced.

Vocabulary

- energy level

Introduction

By 1913, our concept of the atom had evolved from Dalton's idea of indivisible spheres to Thomson's plum-pudding model and then to Rutherford's nuclear atom theory.

Rutherford, in addition to carrying out the experiment that demonstrated the presence of the atomic nucleus, proposed that the electrons circled the nucleus in a planetary-like motion. The planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system. Unfortunately, there was a serious flaw in the planetary model. At that time, it was already known that when a charged particle moves in a curved path, the particle emits some form of light or radio waves and loses energy in doing so. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus (because of the loss of potential energy) and would eventually crash into the nucleus. Scientists, however, saw no evidence that electrons were constantly emitting energy or crashing into the nucleus. These difficulties cast a shadow on the planetary model and indicated that it would eventually be replaced.

The replacement model came in 1913 when the Danish physicist Niels Bohr (pictured in **Figure** 4.14) proposed an electron cloud model where the electrons orbit the nucleus but did not have to lose energy.

Bohr's Energy Levels

The key idea in the Bohr model of the atom is that electrons occupy definite orbits that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in one of the allowable orbits and have the precise energy required for that orbit. Orbits closer to the nucleus would require the electrons to have a smaller amount of energy, and orbits farther from the nucleus would require the electrons to have a greater amount of energy. The possible orbits are known as **energy levels**.

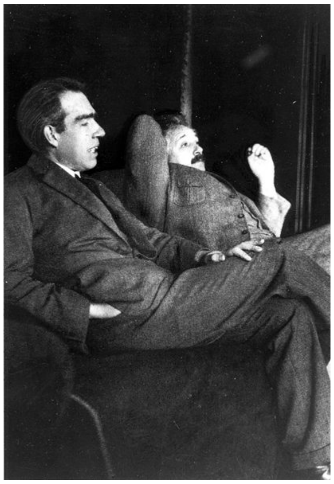
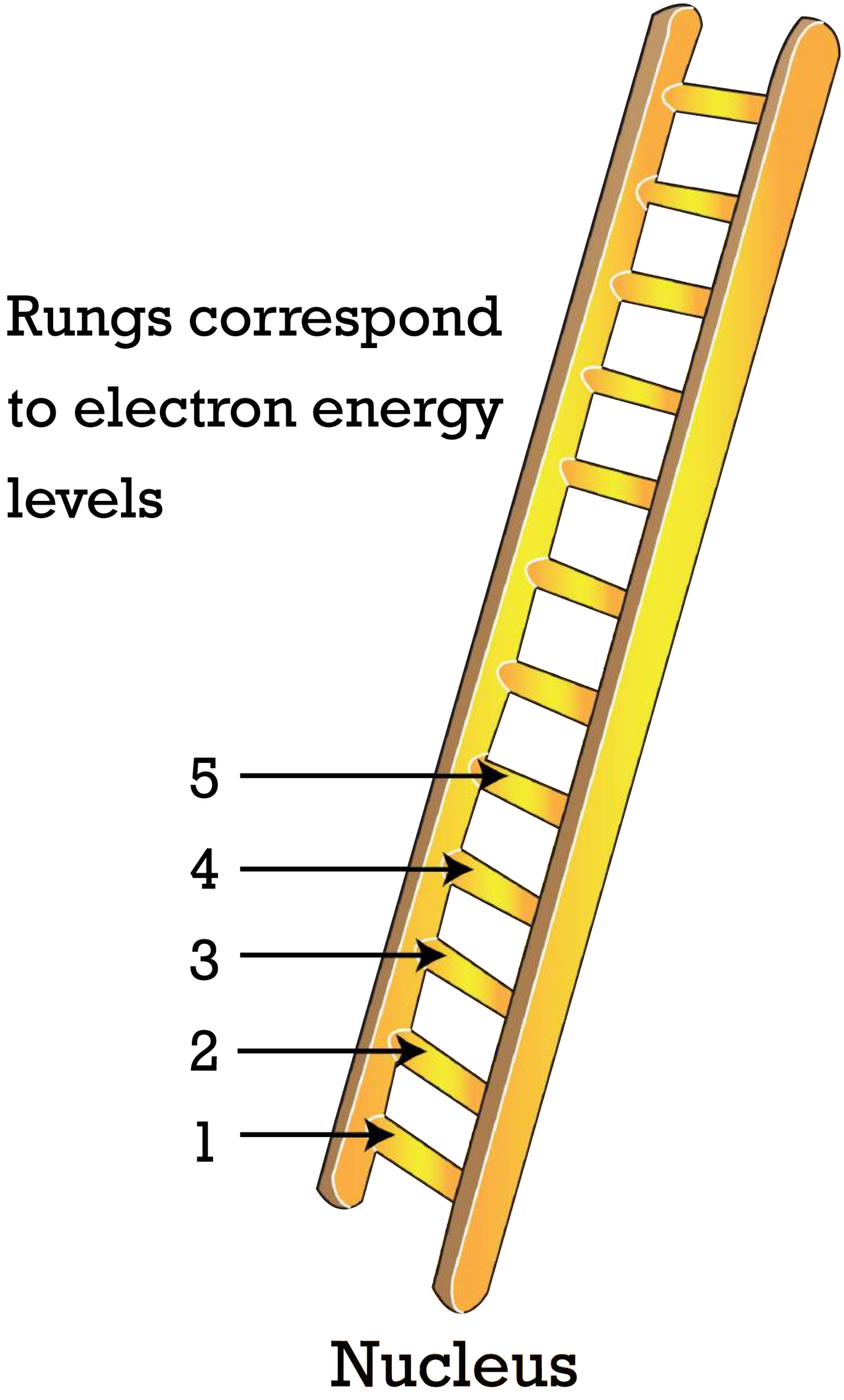


FIGURE 4.14

Niels Bohr and Albert Einstein in 1925.

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. It would be like a ladder that had rungs at certain heights (see image below). The only way you can be on that ladder is to be on one of the rungs, and the only way you could move up or down is to move to one of the other rungs. Other rules for the ladder are that only one person can be on a given rung and that the ladder occupants must be on the lowest rung available. Suppose we had such a ladder with 10 rungs. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all the lower rungs are full. Bohr worked out the rules for the maximum number of electrons that could be in each energy level in his model. In its normal state (ground state), this would require the atom to have all of its electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, the Bohr model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.



Bohr Model and Atomic Spectra

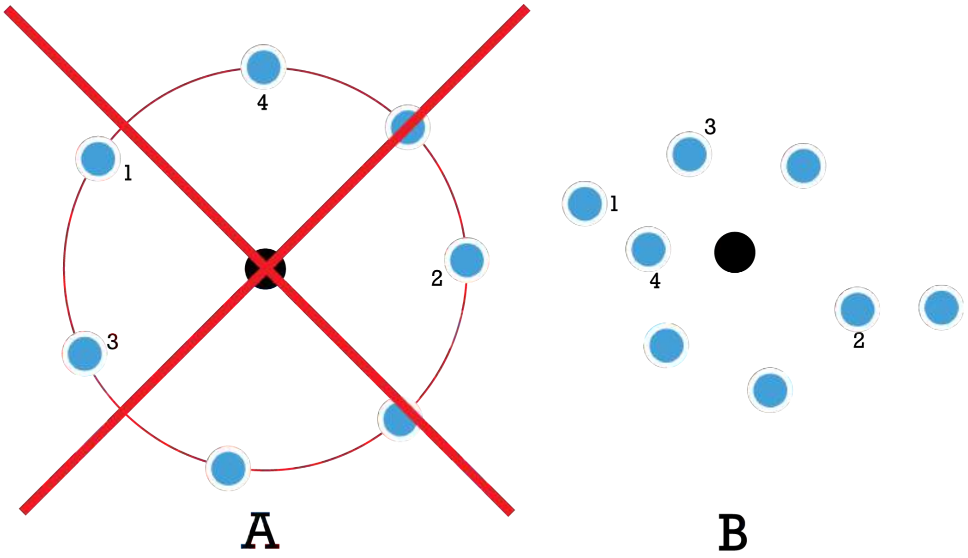
The evidence used to support the Bohr model came from the atomic spectra. Bohr suggested that an atomic spectrum is created when the electrons in an atom move between energy levels. The electrons typically have the lowest energy possible, but upon absorbing energy, the electrons would jump to a higher energy level, producing an excited and unstable state. The electrons would then immediately fall back to a lower energy level and re-emit the absorbed energy. The energy emitted during these electron "step downs" would be emitted as light and would correspond with a specific line in the atomic emission spectrum. Bohr was able to mathematically produce a set of energy levels for the hydrogen atom. In his calculations, the differences between the energy levels were the exact same energies of the frequencies of light emitted in the hydrogen spectrum. One of the most convincing aspects of the Bohr model was that it predicted that the hydrogen atom would emit some electromagnetic radiation outside the visible range. When scientists looked for these emissions in the infrared region, they were able to find them at the exact frequencies predicted by the Bohr model. Bohr's theory was rapidly accepted and he received the Nobel Prize for physics in 1922.

Shortcomings of the Bohr Model

The development of the Bohr model is a good example of applying the scientific method. It shows how the observations of the atomic spectra led to the creation of a hypothesis about the nature of electron clouds. The hypothesis also made predictions about emissions that had not yet been observed (the infrared light emissions). Predicted observations such as these provide an opportunity to test the hypothesis through experimentation. When these predictions were found to be correct, they provided evidence in support of the theory. Of course, further observations can also provide insupportable evidence that will cause the theory to be rejected or modified. In the case of the Bohr model of the atom, it was determined that the energy levels in atoms with more than one electron could not be successfully calculated. Bohr's system was only successful for atoms that have a single electron, which meant that the Bohr model did not accurately reflect the behaviors of most atoms.

Another problem with Bohr's theory was that the Bohr model did not explain why certain energy levels existed. As mentioned earlier in this lesson, at the time it was already known that charged particles emit some form of light or radio waves when moving in a curved path. Scientists have used this principle to create radio signals since 1895. This was the serious flaw in Rutherford's planetary model of the atom, which Bohr attempted to deal with by suggesting his electron cloud model. Although his calculated energy levels for the hydrogen were supported by hydrogen's emission spectrum, Bohr did not, however, explain why only the exact energy levels he calculated were present.

Yet another problem with the Bohr model was the predicted positions of the electrons in the electron cloud. If Bohr's model were correct, the electron in the hydrogen atom in ground state would always be the same distance from the nucleus. Although the actual path that the electron followed could not be determined, scientists were able to determine the positions of the electron at various times. If the electron circled the nucleus as suggested by the Bohr model, the electron positions would always be the same distance from the nucleus. In reality, the electron is found at many different distances from the nucleus. In the figure below, the left side of the image (labeled as A) shows the random positions an electron would occupy as predicted by the Bohr model, while the right side (labeled as B) shows some actual positions of an electron.

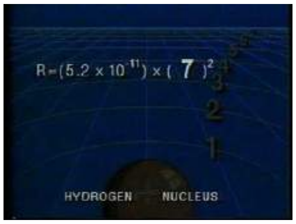


The Bohr model was not, however, a complete failure. It provided valuable insights that triggered the next step in the development of the modern concept of the atom.

Lesson Summary

- The Bohr model suggests each atom has a set of unchangeable energy levels, and electrons in the electron cloud of that atom must be in one of those energy levels.
- The Bohr model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for the Bohr model of the atom.
- The Bohr model was only successful in calculating energy levels for the hydrogen atom.

This video provides a summary of the Bohr atomic model and how the Bohr model improved upon Rutheford's model **(1i; 1g IE, 1k IE)**: http://www.youtube.com/watch?v=bDUxygs7Za8 (9:08).



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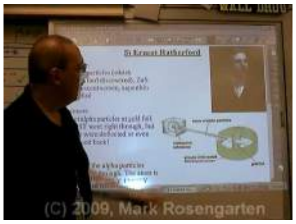
This video describes the important contributions of many scientists to the modern model of the atom. It also explains Rutherford's gold foil experiment **(1g IE)**: http://www.youtube.com/watch?v=6773jO6fMnM (9:08).



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Further Reading / Supplemental Links

These various videos examine the components of the Bohr model of the atom.

- http://www.youtube.com/watch?v=QI50GBUJ48s&feature=related
- http://www.youtube.com/watch?v=hpKhjKrBn9s
- http://www.youtube.com/watch?v=-YYBCNQnYNM&feature=related

Review Questions

1. What is the key concept in the Bohr model of the atom?
2. What is the general relationship between the amount of energy of an electron energy level and its distance from the nucleus?
3. According to Bohr's theory, how can an electron gain or lose energy?
4. What happens when an electron in an excited atom returns to its ground level?
5. What concept in Bohr's theory makes it impossible for an electron in the ground state to give up energy?
6. Use the Bohr model to explain how an atom emits a specific set of frequencies of light when it is heated or has electric current passed through it.
7. How do scientists know that the sun contains helium atoms when no one has even taken a sample of material from the sun?

- *Bohr and Einstein* . Public domain.
- Images of neon (http://en.wikipedia.org/wiki/File:NeTube.jpg) and argon (http://en.wikipedia.org/wiki/File:ArTube.jpg) signs created by Pslawinski, created into a composite by Richard Parsons. *Neon and argon gas signs*. CC-BY-SA 2.5.
- *Iron emission spectrum* . Public domain.
- *Hydrogen Emission Spectrum* . Public domain.

4.6 Isotopes and Atomic Mass

Lesson Objectives

- Define atomic number.
- Define mass number.
- Understand how isotopes differ from one another and be able to designate them by various methods.
- Be able to calculate the average atomic mass of an element.

Lesson Vocabulary

- atomic mass
- atomic mass unit
- atomic number
- isotope
- mass number
- nuclide

Introduction

Atoms are the fundamental building blocks of all matter and are composed of protons, neutrons, and electrons. Because atoms are electrically neutral, the number of positively charged protons must be equal to the number of negatively charged electrons. One of Dalton's points in his atomic theory was that all atoms of a given element are identical. In this section, we will see how this is not strictly true, thanks to variability in the number of neutrons that an atom may contain.

Atomic Number

The **atomic number (Z)** *of an element is the number of protons in the nucleus of each atom of that element*. An atom can be classified as a particular element based solely on its atomic number. For example, any atom with an atomic number of 8 (its nucleus contains 8 protons) is an oxygen atom, and any atom with a different number of protons would be a different element. The periodic table (**Figure** 4.15) displays all of the known elements and is arranged in order of increasing atomic number. In this table, an element's atomic number is indicated above the elemental symbol. Hydrogen, at the upper left of the table, has an atomic number of 1. Every hydrogen atom has one proton in its nucleus. Next on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, beryllium atoms have four, and so on.

Since atoms are neutral, the number of electrons in an atom is equal to the number of protons. Hydrogen atoms all have one electron occupying the space outside of the nucleus.

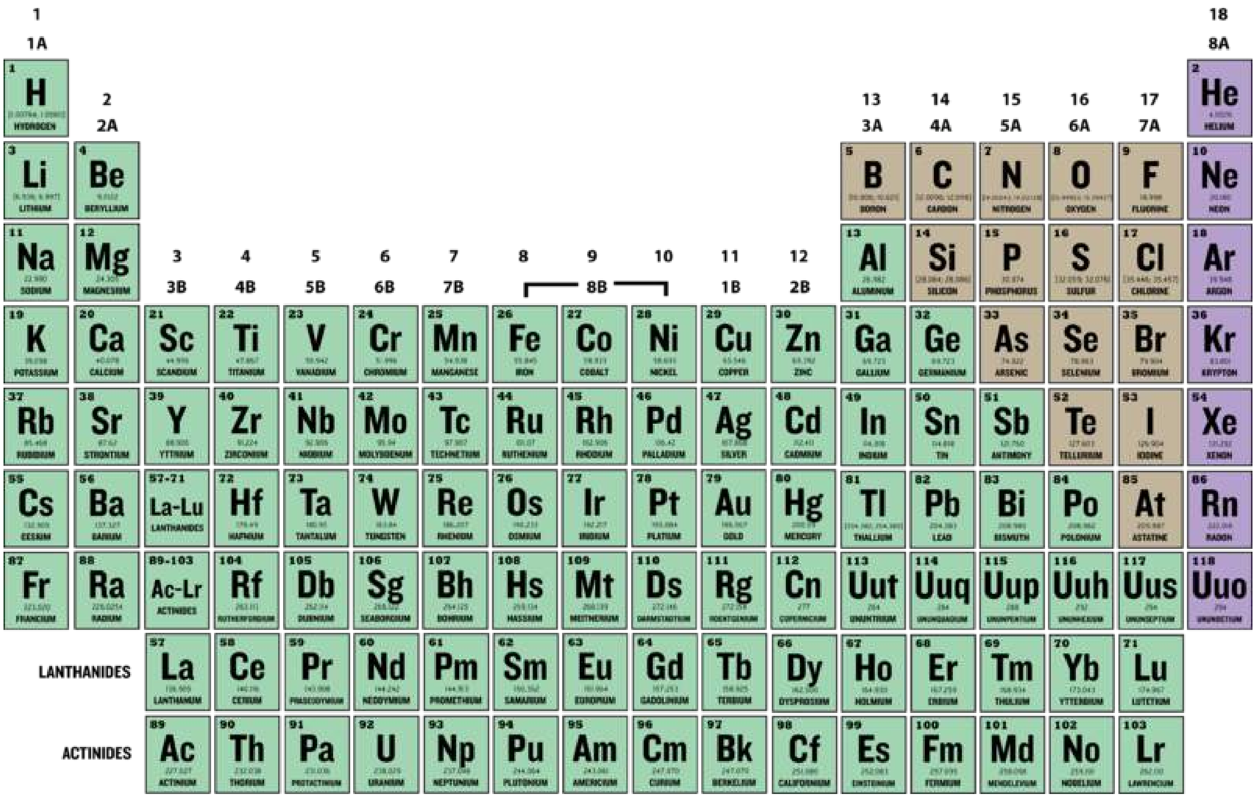


FIGURE 4.15

The periodic table of the elements.

Mass Number

Rutherford's experiment showed that the vast majority of the mass of an atom is concentrated in its nucleus, which is composed of protons and neutrons. The **mass number** *is defined as the total number of protons and neutrons in an atom.* Consider **Table** 4.1, which shows data from the first six elements of the periodic table.

TABLE 4.1: **Atoms of the First Six Elements**

Name	Symbol	Atomic Number	Protons	Neutrons	Electrons	Mass Number
Hydrogen	H	1	1	0	1	1
Helium	He	2	2	2	2	4
Lithium	Li	3	3	4	3	7
Beryllium	Be	4	4	5	4	9
Boron	B	5	5	6	5	11
Carbon	C	6	6	6	6	12

Consider the element helium. Its atomic number is 2, so it has two protons in its nucleus. Its nucleus also contains two neutrons. Since 2 + 2 = 4, we know that the mass number of the helium atom is 4. Finally, the helium atom also contains two electrons, since the number of electrons must equal the number of protons. This example may lead you to believe that atoms have the same number of protons and neutrons, but a further examination of **Table** 4.1 will show that this is not the case. Lithium, for example has three protons and four neutrons, giving it a mass number of 7.

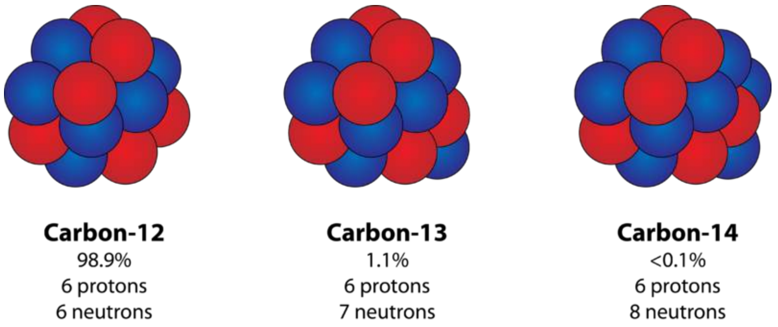


FIGURE 4.16

Nuclei of the three isotopes of carbon: Almost 99% of naturally occurring carbon is carbon-12, whose nucleus consists of six protons and six neutrons. Carbon-13 and carbon-14, with seven or eight neutrons, respectively, have a much lower natural abundance.

Each carbon atom has the same number of protons (6), which is equal to its atomic number. Each carbon atom also contains six electrons, allowing the atom to remain electrically neutral. However the number of neutrons varies from six to eight. **Isotopes** *are atoms that have the same atomic number but different mass numbers due to a change in the number of neutrons*. The three isotopes of carbon can be referred to as carbon-12 ($^{12}_{6}$C), carbon-13 ($^{13}_{6}$C), and carbon-14 ($^{14}_{6}$C). Naturally occurring samples of most elements are mixtures of isotopes. Carbon has only three natural isotopes, but some heavier elements have many more. Tin has ten stable isotopes, which is the most of any element. The term **nuclide** *refers to the nucleus of a given isotope of an element*. The nucleus of a given carbon atom will be one of the three possible nuclides discussed above.

While the presence of isotopes affects the mass of an atom, it does not affect its chemical reactivity. Chemical behavior is governed by the number of electrons and the number of protons. Carbon-13 behaves chemically in exactly the same way as the more plentiful carbon-12.

Sample Problem 4.2: Composition of an Atom

How many protons, neutrons, and electrons are present in each of the nuclides below?

1. Iron (Fe): atomic number = 26, mass number = 56
2. Iodine-127 (atomic number = 53)
3. $^{31}_{15}$P

Step 1: List the known and unknown quantities and plan the problem.

Known

- Atomic number and mass number for each

Unknown

- Number of protons, electrons, and neutrons

Each shows a different way to specify an isotope of an atom. Use the definitions of atomic number and mass number to calculate the numbers of protons, neutrons, and electrons.

Step 2: Calculate

Number of protons = atomic number

1. 26
2. 53
3. 15

Number of electrons = number of protons

1. 26
2. 53
3. 15

Number of neutrons = mass number - atomic number

1. 56 - 26 = 30
2. 127 - 53 = 74
3. 31 - 15 = 16

Step 3: Think about your result.

For each atom, the results are consistent with the definitions of atomic number and mass number.

Do the practice problems below. If necessary, refer to the periodic table in **Figure** 4.15 for the atomic number or symbol of the given element.

Practice Problems

1. How many protons, neutrons, and electrons are there in the atom $^{19}_{9}F$?
2. How many protons, neutrons, and electrons are there in an atom of lead-207?
3. A certain atom has an atomic number of 36 and a mass number of 84. Write out the designation for this isotope in both nuclide symbol form and in hyphenated form.
4. An atom has a mass number of 59 and contains 32 neutrons in its nucleus. What element is it?

Atomic Mass

The masses of individual atoms are very, very small. However, using a modern device called a mass spectrometer, it is possible to measure such minuscule masses. An atom of oxygen-16, for example, has a mass of 2.66×10^{-23} g. While comparisons of masses measured in grams would have some usefulness, it is far more practical to have a system that will allow us to more easily compare relative atomic masses. Scientists decided on using the carbon-12 nuclide as the reference standard by which all other masses would be compared. By definition, one atom of carbon-12 is assigned a mass of exactly 12 atomic mass units (amu). An **atomic mass unit** *is defined as a mass equal to one twelfth the mass of an atom of carbon-12*. The mass of any isotope of any element is expressed in relation to the carbon-12 standard. For example, one atom of helium-4 has a mass of 4.0026 amu. An atom of sulfur-32 has a mass of 31.972 amu.

The carbon-12 atom has six protons and six neutrons in its nucleus for a mass number of 12. Since the nucleus accounts for nearly all of the mass of the atom, a single proton or single neutron has a mass of approximately 1 amu. However, as seen by the helium and sulfur examples, the masses of individual atoms are not quite whole numbers. This is because an atom's mass is affected very slightly by the interactions of the various particles within the nucleus and also includes the small mass added by each electron.

As stated in the section on isotopes, most elements occur naturally as a mixture of two or more isotopes. **Table** 4.2 shows the naturally occurring isotopes of several elements along with the percent natural abundance of each.

TABLE 4.2: Atomic Masses and Percent Abundances of Some Natural Isotopes

Element	Isotope (symbol)	Percent natural abundance	Atomic mass (amu)	Average atomic mass (amu)
Hydrogen	${}^{1}_{1}H$	99.985	1.0078	1.0079
	${}^{2}_{1}H$	0.015	2.0141	
	${}^{3}_{1}H$	negligible	3.0160	
Carbon	${}^{12}_{6}C$	98.89	12.000	12.011
	${}^{13}_{6}C$	1.11	13.003	
	${}^{14}_{6}C$	trace	14.003	
Oxygen	${}^{16}_{8}O$	99.759	15.995	15.999
	${}^{17}_{8}O$	0.037	16.995	
	${}^{18}_{8}O$	0.204	17.999	
Chlorine	${}^{35}_{17}Cl$	75.77	34.969	35.453
	${}^{37}_{17}Cl$	24.23	36.966	
Copper	${}^{63}_{29}Cu$	69.17	62.930	63.546
	${}^{65}_{29}Cu$	30.83	64.928	

For some elements, one particular isotope is much more abundant than any other isotopes. For example, naturally occurring hydrogen is nearly all hydrogen-1, and naturally occurring oxygen is nearly all oxygen-16. For many other elements, however, more than one isotope may exist in substantial quantities. Chlorine (atomic number 17) is a yellowish-green toxic gas. About three quarters of all chlorine atoms have 18 neutrons, giving those atoms a mass number of 35. About one quarter of all chlorine atoms have 20 neutrons, giving those atoms a mass number of 37. Were you to simply calculate the arithmetic average of the precise atomic masses, you would get approximately 36.

$$(34.969 + 36.966)/2 = 35.968 \text{ amu}$$

As you can see, the average atomic mass given in the last column of **Table** 4.2 is significantly lower. Why? The reason is that we need to take into account the natural abundance percentages of each isotope in order to calculate what is called the weighted average. The **atomic mass** *of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element.* The sample problem below demonstrates how to calculate the atomic mass of chlorine.

Sample Problem 4.3: Calculating Atomic Mass

Use the atomic masses of each of the two isotopes of chlorine along with their percent abundances to calculate the average atomic mass of chlorine.

Step 1: List the known and unknown quantities and plan the problem.

Known

- chlorine-35: atomic mass = 34.969 amu and % abundance = 75.77%
- chlorine-37: atomic mass = 36.966 amu and % abundance = 24.23%

Unknown

- Average atomic mass of chlorine

Change each percent abundance into decimal form by dividing by 100. Multiply this value by the atomic mass of that isotope. Add together the results for each isotope to get the average atomic mass.

Step 2: Calculate

chlorine-35	$0.7577 \times 34.969 = 26.50$ amu
chlorine-37	$0.2423 \times 36.966 = 8.957$ amu
average atomic mass	$26.50 + 8.957 = 35.45$ amu

Note: Applying significant figure rules results in the 35.45 amu result without excessive rounding error. In one step:

$$(0.7577 \times 34.969) + (0.2423 \times 36.966) = 35.45 \text{ amu}$$

Step 3: Think about your result.

The calculated average atomic mass is closer to 35 than to 37 because a greater percentage of naturally occurring chlorine atoms have a mass number of 35. It agrees with the value from **Table** 4.2.

Practice Problem

5. The element bromine consists of two naturally occurring isotopes. The isotope with a mass of 78.92 amu has a percent abundance of 50.69%, while the isotope with a mass of 80.92 amu has a percent abundance of 49.31%. Calculate the average atomic mass of bromine.

The atomic masses for each element on the periodic table are average atomic masses. For later calculations involving atomic mass, we will use these values and round each one to four significant figures.

Lesson Summary

- The atomic number of an element is equal to the number of protons in its nucleus.
- The mass number of an element is equal to the sum of the protons and the neutrons in its nucleus.
- Isotopes are atoms of the same element that have a different mass number because of a variation in the number of neutrons.
- The average atomic mass of an element can be calculated from the atomic masses and percent natural abundances of each naturally occurring isotope.

Lesson Review Questions

Recall

1. Why are all atoms electrically neutral?
2. How many protons are in the nucleus of each of the following atoms?
   a. neon
   b. gold
   c. strontium
   d. uranium
3. What part of Dalton's atomic theory is disproved by the existence of isotopes?
4. Which isotope is used as the reference standard for the atomic mass unit?

Apply Concepts

5. The average atomic mass of all naturally occurring lithium atoms is 6.941 amu. The two isotopes of lithium are lithium-6 and lithium-7. Are these isotopes equally common? If not, which is more plentiful in nature, and how do you know?

Think Critically

6. A certain atom contains 28 protons, 28 electrons, and 31 neutrons. Provide the following:
   a. atomic number
   b. mass number
   c. name of element
7. How many protons, neutrons, and electrons are in an atom of cesium-133?
8. Complete **Table** 4.3:

TABLE 4.3: Table for Problem 8

Isotope	Nuclide Symbol	Atomic Number	Mass Number
sodium-23			
	$^{75}_{33}$As		
silver-108			

9. Which of the following is an isotope of $^{40}_{18}$Ar? Explain.
   a. $^{40}_{20}$Ca
   b. $^{39}_{18}$Ar
   c. $^{40}_{18}$Ar
10. Fill in **Table** 4.4:

TABLE 4.4: Table for Problem 10

Isotope	Number of Protons	Number of Electrons	Number of Neutrons	Nuclide Symbol
hydrogen-1				
hydrogen-2				
beryllium-9				
aluminum-27				

11. Fill in **Table** 4.5:

TABLE 4.5: Table for Problem 11

Element	Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Nuclide Symbol
Nitrogen			14				
	B		11				
		30				35	

TABLE 4.5: (continued)

Element	Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Nuclide Symbol
					77	116	
							$^{56}_{26}\text{Fe}$

12. The element tungsten (W) is known best as a metal that is used as filaments for light bulbs. Naturally occurring tungsten consists of the five isotopes shown below. Calculate the atomic mass of tungsten.

tungsten-180 atomic mass = 179.947 amu percent abundance = 0.12%
tungsten-182 atomic mass = 181.948 amu percent abundance = 26.50%
tungsten-183 atomic mass = 182.950 amu percent abundance = 14.31%
tungsten-184 atomic mass = 183.951 amu percent abundance = 30.64%
tungsten-186 atomic mass = 185.954 amu percent abundance = 28.43%

Further Reading / Supplemental Links

There are lots of websites to help you understand the atom and its history.

Lectures:

- Elements and Atoms: http://www.youtube.com/watch?v=IFKnq9QM6_A
- Introduction to the Atom: http://www.youtube.com/watch?v=1xSQlwWGT8M

Informative websites:

- Atomic History - A Brief Discovery: http://www.pbs.org/wgbh/nova/diamond/insidehistory.html
- All About Atoms: http://education.jlab.org/atomtour/index.html

Simulations:

- Build an Atom: http://phet.colorado.edu/en/simulation/build-an-atom
- Molecular Workbench - Atomic Structure: http://workbench.concord.org/database/activities/47.html
- See Inside a Diamond: http://www.pbs.org/wgbh/nova/diamond/insidewave.html
- Isotopes and Atomic Mass: http://phet.colorado.edu/en/simulation/isotopes-and-atomic-mass
- Atomic Structure: http://freezeray.com/flashFiles/atomicStructure.htm
- Atom Builder: http://freezeray.com/flashFiles/atomBuilder.htm
- Tennis Ball Isotopes: http://www.youtube.com/watch?v=oLnuXpf4hsA

Games:

- Element Math Game: http://education.jlab.org/elementmath/index.html
- Looking For the Top Quark: http://education.jlab.org/topquarkgame/index.html
- Atoms and Matter Crossword Puzzle: http://education.jlab.org/sciencecrossword/atoms_01.html

4.7 The Nuclear Model of the Atom

Lesson Objectives

- Distinguish between the three main subatomic particles.
- Understand the contributions of J. J. Thomson, Robert Millikan, and Ernest Rutherford to atomic theory.
- Describe the structure of the nuclear atom.

Lesson Vocabulary

- atomic model
- cathode ray
- cathode ray tube
- electron
- neutron
- nucleus
- proton

Introduction

Dalton's atomic theory represented an improvement over the idea of Democritus because the theory was based on experimental findings and the scientific method. However, his theory did have its shortcomings. He believed that atoms were indivisible, meaning that the atom was the smallest possible component of matter. Further investigations in the late 1800s proved that atoms can indeed be broken down into smaller particles. It is the unique number and arrangement of these subatomic particles that makes atoms of one element different from those of every other element. The three fundamental particles are called the proton, the neutron, and the electron.

Discovery of the Electron

The first discovery of a subatomic particle resulted from experiments into the nature of the relationship between electricity and matter.

Cathode Rays

In 1897, English physicist J. J. Thomson (1856-1940) experimented with *a device called a* **cathode ray tube**, *in which an electric current was passed through gases at low pressure*. A cathode ray tube (**Figure** 4.17) consists of a sealed glass tube fitted at both ends with metal disks called electrodes. The electrodes are then connected to a source of electricity. One electrode, called the anode, becomes positively charged while the other electrode, called the cathode, becomes negatively charged. Once this happens, *a beam called a* **cathode ray** *travels from the cathode to the anode*.

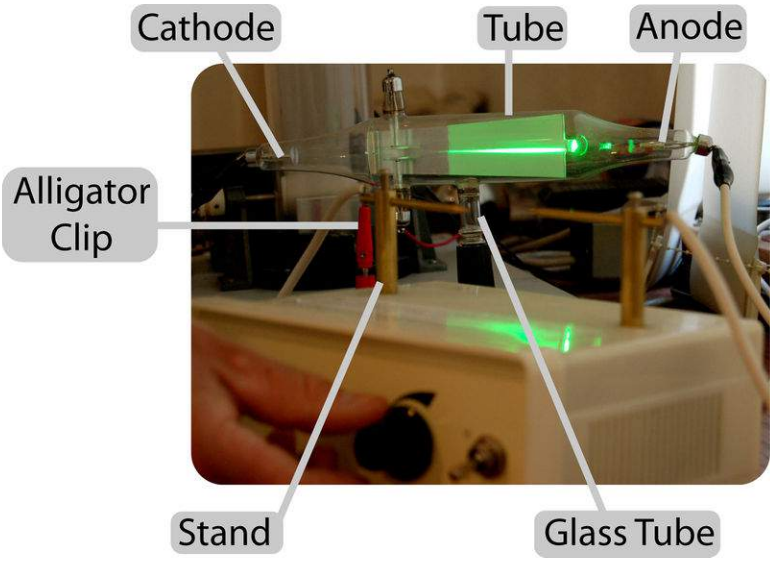


FIGURE 4.17
To produce a cathode ray, a glass tube filled with a low-pressure gas is connected to a power source. The green beam, or cathode ray, moves from the cathode to the anode.

Investigations were carried out to determine the nature of the cathode ray. The results of two further experiments supported the hypothesis that the cathode ray consisted of a stream of particles.

1. When an object was placed between the cathode and the opposite end of the tube, it cast a shadow on the glass.
2. A cathode ray tube was constructed with a small metal rail between the two electrodes. Attached to the rail was a paddle wheel capable of rotating along the rail. Upon connecting the cathode ray tube to a power source, the wheel rotated from the cathode towards the anode. This suggested that the cathode ray was made of particles that must have mass.

In order to determine if the cathode ray consisted of charged particles, Thomson used magnets and charged plates to deflect the ray (**Figure** 4.18). His findings are summarized below.

3. Cathode rays were deflected by a magnetic field in the same manner as a wire carrying an electric current, which was known to be negatively charged.
4. Cathode rays were deflected away from a negatively charged metal plate and towards a positively charged plate.

Thomson knew that opposite charges attract one another, while like charges repel one another. Together, the results of the cathode ray tube experiments showed that cathode rays are actually streams of tiny negatively charged particles moving at very high speeds. While Thomson originally called these particles corpuscles, they were later named electrons.

Thomson conducted further experiments which allowed him to calculate the charge-to-mass ratio (e/m_e) of the electron. In units of coulombs to grams, $e/m_e = 1.8 \times 10^8$ C/g. He found that this value was a constant and did not depend on the gas used in the cathode ray tube or on the metal used as the electrodes. He concluded that **electrons** *were negatively charged subatomic particles present in atoms of all elements*.

Watch a video of a cathode ray tube experiment at www.dlt.ncssm.edu/core/Chapter3-Atomic_Str_Part1/cathode-rm-lg.htm .

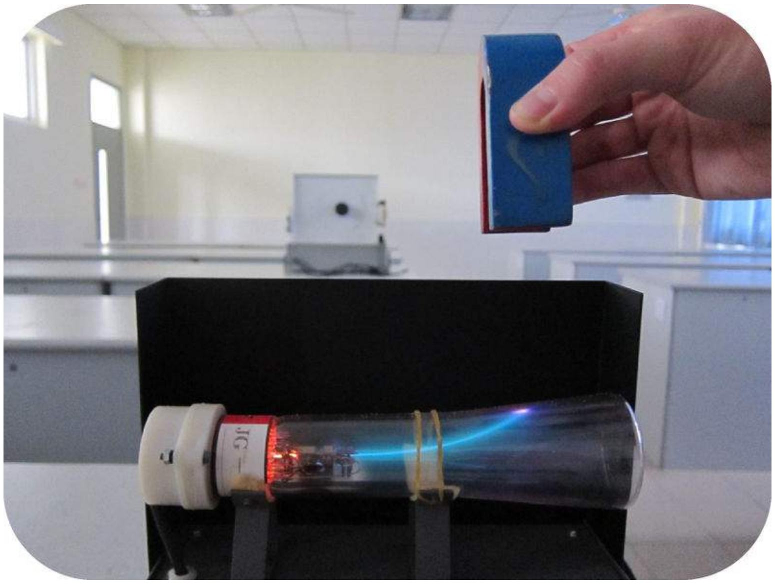


FIGURE 4.18

The deflection of a cathode ray by a magnet. From the direction and extent of the deflection, Thomson was able to determine the charge-to-mass ratio of the electron.

Charge and Mass of the Electron

American physicist Robert Millikan (1868-1953) carried out a series of experiments between 1908 and 1917 that allowed him to determine the charge of a single electron. Millikan's experiment was called the oil drop experiment (**Figure** 4.19).

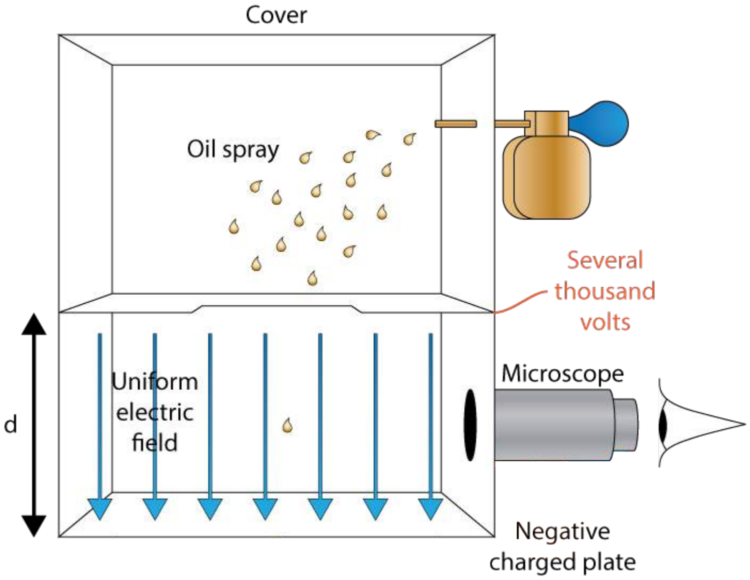


FIGURE 4.19

Millikan's oil drop experiment: Oil drops that are sprayed into the main chamber fall through a tiny hole into an electric field, after which they can be viewed through a microscope. This experiment allowed Millikan to determine the charge of the electron.

When tiny drops of oil were sprayed into a chamber, the oil drops picked up a static charge and were suspended between two charged plates. Millikan was able to observe the motion of the oil drops with a microscope and found that the drops lined up in a specific way between the plates, based on the number of electric charges that they had acquired. From the data gathered in this experiment, he was able to accurately determine the charge of an individual electron. Then, using Thomson's previous measurement of an electron's charge-to-mass ratio, he was also able to calculate the mass of a single electron.

$$\text{Charge of one electron} = -1.602 \times 10^{-19}\ \text{C}$$
$$\text{Mass of one electron} = 9.11 \times 10^{-28}\ \text{g}$$

The incredibly small mass of the electron was found to be approximately 1/1840 the mass of a hydrogen atom, so scientists realized that atoms must also contain other, far more massive particles. Additionally, at least one of these particles must carry a positive charge, because complete atoms are electrically neutral.

Participate in a simulation of Millikan's oil drop experiment at this site: http://www.dlt.ncssm.edu/core/Chapter3-Atomic_Str_Part1/Chapter3-Animations/OilDrop.htm .

Protons and Neutrons

If cathode rays are electrons that are given off by the metal atoms of the cathode, then what remains of the atoms that have lost those electrons? We know several basic things about electrical charges. They are carried by particles of matter. Millikan's experiment showed that they exist as whole-number multiples of a single basic unit. Atoms have no overall electrical charge, meaning that each and every atom contains an exactly equal number of positively and negatively charged particles. A hydrogen atom, the simplest kind of atom, contains only one electron. When that electron is removed, a positively charged particle should remain.

In 1886, Eugene Goldstein (1850-1930) discovered evidence for the existence of this positively charged particle. Using a cathode ray tube with holes in the cathode, he noticed that there were rays traveling in the opposite direction from the cathode rays. He called these canal rays and showed that they were composed of positively charged particles. The **proton** *is a positively charged subatomic particle that is present in all atoms*. The mass of the proton is about 1840 times the mass of the electron.

In 1932, English physicist James Chadwick (1891-1974) discovered a third subatomic particle. The **neutron** *is a subatomic particle with no electrical charge and a mass that is approximately the same as the mass of a proton*. **Table** 4.6 summarizes the properties of the three fundamental subatomic particles.

TABLE 4.6: **Properties of Subatomic Particles**

Particle	Symbol	Relative Electrical Charge	Relative Mass (amu)*	Actual Mass (g)
Electron	e^-	1-	1/1840	9.11×10^{-28}
Proton	p^+	1+	1	1.67×10^{-24}
Neutron	n^0	0	1	1.67×10^{-24}

- 1 amu (atomic mass unit) = 1.66×10^{-24} g

Discovery of the Atomic Nucleus

The next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This is a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible. Following the discovery of the electron, J. J. Thomson developed what became known as the "plum pudding" model (**Figure** 4.20). In this model, the electrons were suspended in a uniform lump of positive charge like blueberries in a muffin. This model of the atom soon gave way, however, to a new model developed by New Zealander Ernest Rutherford (1871-1937).

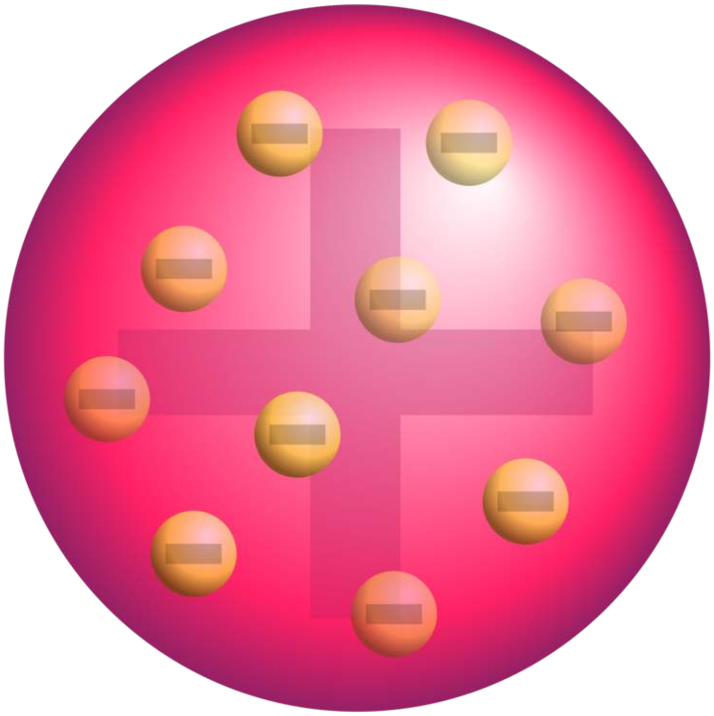


FIGURE 4.20

In Thomson's plum pudding model of the atom, electrons are embedded in a uniform sphere of positive charge. Plum pudding is an English dessert that is similar to a blueberry muffin.

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. The experimental setup is shown in **Figure** 4.21. When they bombarded very thin sheets of gold foil with fast moving alpha particles, they got some unexpected results. An alpha particle is a type of positively charged particle whose mass is about four times that of a hydrogen atom. It occurs naturally as a product of radioactive decay.

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed undeflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. Nothing had prepared them for this discovery. In a famous quotation, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue paper and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** *is the tiny, dense, central core of the atom and is composed of protons and neutrons.*

Rutherford's atomic model became known as the nuclear model. In this model, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in a nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not

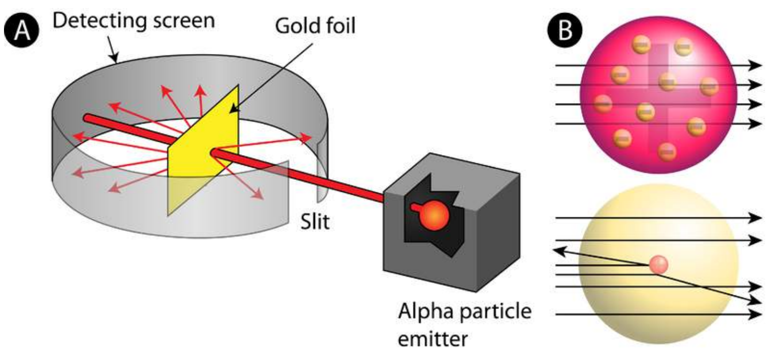


FIGURE 4.21

(A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil, which was surrounded by a screen that would allow detection of the deflected particles. (B) According to the plum pudding model (*top*) all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be better explained by an atom that contained a very small, dense, positively-charged nucleus (*bottom*).

completely address the nature of the electrons and the way in which they occupied the vast space around the nucleus. It was not until some years later that a more complete understanding of the electron was achieved. This proved to be the key to understanding the chemical properties of elements.

Watch a simulation of Rutherford's experiment through The Concord Consortium's Molecular Workbench. You will need to download the Molecular Workbench application from http://mw.concord.org/modeler/index.html . After installing Molecular Workbench, open it. You will see a browser window. In the address bar, type: "http://mw2.concord.org/public/student/motionandforce/rutherford.cml ." You will be taken to the simulation.

Lesson Summary

- The three fundamental subatomic particles are the electron, the proton, and the neutron.
- Thomson used the cathode ray tube to discover the electron and determine its negative charge.
- Millikan determined the charge and mass of the electron with the oil-drop experiment.
- Rutherford's gold foil experiment provided evidence for the atomic nucleus, a small dense core of the atom which contains the positive charge and most of the mass.
- The nuclear model of the atom is one in which the nucleus is composed of protons and neutrons, while electrons are distributed throughout the rest of the space.

Lesson Review Questions

Recall

1. What evidence did Thomson have for each statement below?
    a. Electrons are negatively charged.
    b. Electrons are identical and are present in all atoms.

2. What additional knowledge about electrons did Millikan's experiment provide?
3. Compare the mass (in amu) and relative charge of a neutron to that of a proton.
4. What evidence did Rutherford have for each statement below?
    a. Atoms are mostly empty space.
    b. The nucleus of the atom is positively charged.

Apply Concepts

5. Trace the development of the atomic model from Dalton to Thomson to Rutherford. Explain what experimental findings led each scientist to alter the previous model and how his new model fit with the new evidence.

Think Critically

6. A hydrogen nucleus, which contains only a single proton, has a diameter of 3.7×10^{-10} cm. The equation for the volume of a sphere is $\frac{4}{3}\pi r^3$, where r is the radius of the sphere.
    a. Assuming that the nucleus is spherical, calculate its volume in cm^3.
    b. Calculate the density of the hydrogen nucleus in g/cm^3.
    c. The densest element known is osmium, which has a density of 22.6 g/cm^3. Comment on the difference in density between osmium and a hydrogen nucleus.
7. How could Rutherford's experiment be modified in order to determine the relative sizes of different nuclei?
8. Why was it not possible to detect the existence of the neutron with experiments analogous to those used in the discovery of the electron?
9. All matter is composed of atoms, which are in turn composed mostly of empty space. Why is it not possible to walk through a wall or to put your hand right through your desk?

Further Reading / Supplemental Links

- A cathode ray is a stream of negatively charged particles, or electrons. An anode ray (also called a canal ray) is composed of positively charged particles. You can view a simulation of a canal ray at http://www.dlt.ncssm.edu/core/Chapter3-Atomic_Str_Part1/Chapter3-Animations/Canal_Ray-CRT.htm .
- You can find other simulations that will help you understand how the structure of the atom was determined without actually being able to see inside:
    - Models of the Hydrogen Atom: http://phet.colorado.edu/en/simulation/hydrogen-atom
    - Rutherford Scattering: http://phet.colorado.edu/en/simulation/rutherford-scattering

4.8 Ions

- State why atoms are neutral in electric charge.
- Describe ions.
- Explain how ions form.
- Identify properties of ions.

The incredible green lights in this cold northern sky consist of charged particles known as ions. Their swirling pattern is caused by the pull of Earth's magnetic north pole. Called the northern lights, this phenomenon of nature shows that ions respond to a magnetic field. Do you know what ions are? Read on to find out.

Atoms Are Neutral

The northern lights aren't caused by atoms, because atoms are not charged particles. An atom always has the same number of electrons as protons. Electrons have an electric charge of -1 and protons have an electric charge of +1. Therefore, the charges of an atom's electrons and protons "cancel out." This explains why atoms are neutral in electric charge.

Q: What would happen to an atom's charge if it were to gain extra electrons?

A: If an atom were to gain extra electrons, it would have more electrons than protons. This would give it a negative charge, so it would no longer be neutral.

Atoms to Ions

Atoms cannot only gain extra electrons. They can also lose electrons. In either case, they become **ions**. Ions are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons. If atoms lose electrons, they become positive ions, or cations. If atoms gain electrons, they become negative ions, or anions.

Consider the example of fluorine (see **Figure** 4.22). A fluorine atom has nine protons and nine electrons, so it is electrically neutral. If a fluorine atom gains an electron, it becomes a fluoride ion with an electric charge of -1.

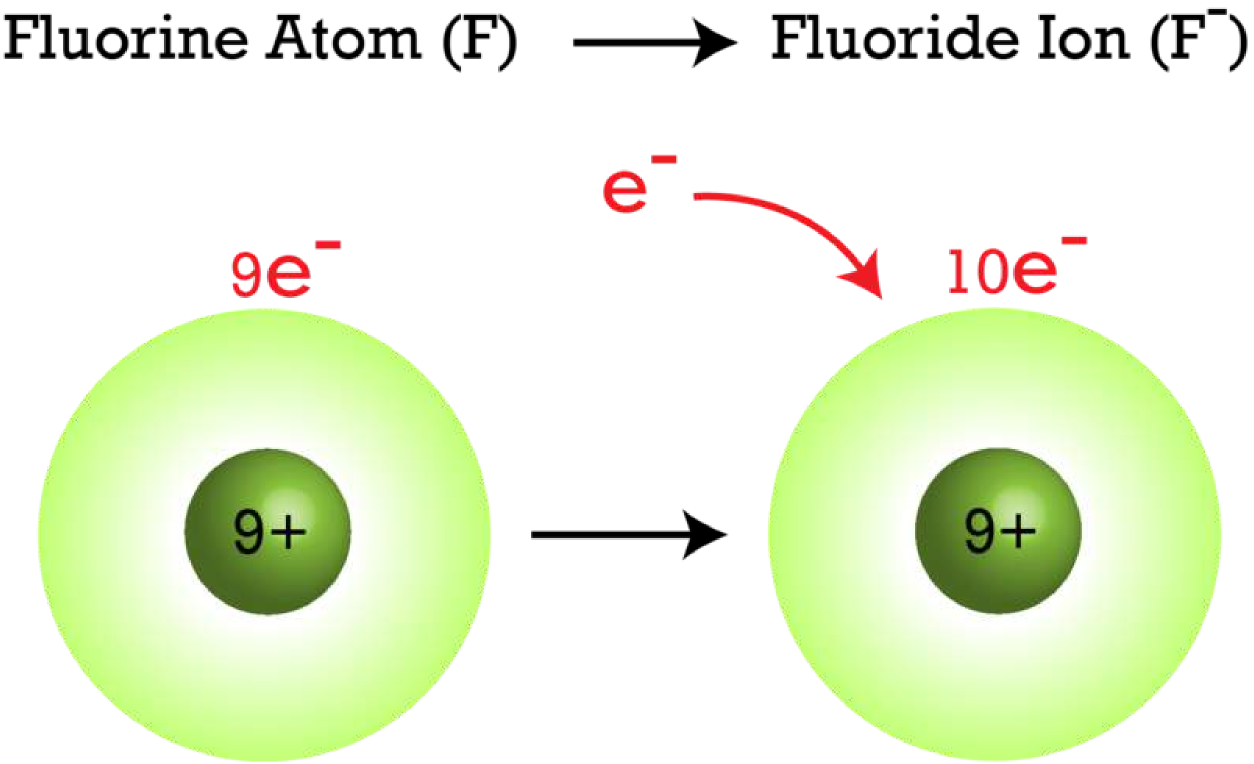


FIGURE 4.22

Names and Symbols

Like fluoride, other negative ions usually have names ending in *–ide*. Positive ions, on the other hand, are just given the element name followed by the word *ion*. For example, when a sodium atom loses an electron, it becomes a positive sodium ion. The charge of an ion is indicated by a plus (+) or minus sign (-), which is written to the right of and just above the ion's chemical symbol. For example, the fluoride ion is represented by the symbol F^-, and the sodium ion is represented by the symbol Na^+. If the charge is greater than one, a number is used to indicate it. For example, iron (Fe) may lose two electrons to form an ion with a charge of plus two. This ion would be represented by the symbol Fe^{2+}. This and some other common ions are listed with their symbols in the **Table** 4.7.

TABLE 4.7: **Some Common Ions**

Cations		Anions	
Name of Ion	**Chemical Symbol**	**Name of Ion**	**Chemical Symbol**
Calcium ion	Ca^{2+}	Chloride	Cl^-
Hydrogen ion	H^+	Fluoride	F^-
Iron(II) ion	Fe^{2+}	Bromide	Br^-

TABLE 4.7: (continued)

Cations		Anions	
Iron(III) ion	Fe^{3+}	Oxide	O^{2-}

Q: How does the iron(III) ion differ from the iron(II) ion?

A: The iron(III) ion has a charge of +3, so it has one less electron than the iron(II) ion, which has a charge of +2.

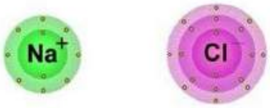
Q: What is the charge of an oxide ion? How does its number of electrons compare to its number of protons?

A: An oxide ion has a charge of -2. It has two more electrons than protons.

How Ions Form

The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation. The radiation may give their outer electrons enough energy to escape from the attraction of the positive nucleus. However, most ions form when atoms transfer electrons to or from other atoms or molecules. For example, sodium atoms may transfer electrons to chlorine atoms. This forms positive sodium ions (Na^+) and negative chloride ions (Cl^-). You can see an animation of this process at the URL below.

http://www.youtube.com/watch?v=xTx_DWboEVs (0:20)



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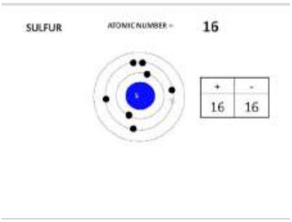
Q: Why do you think atoms lose electrons to, or gain electrons from, other atoms?

A: Atoms form ions by losing or gaining electrons because it makes them more stable and this state takes less energy to maintain. The most stable state for an atom is to have its outermost energy level filled with the maximum possible number of electrons. In the case of metals such as lithium, with just one electron in the outermost energy level, a more stable state can be achieved by losing that one outer electron. In the case of nonmetals such as fluorine, which has seven electrons in the outermost energy level, a more stable state can be achieved by gaining one electron and filling up the outer energy level. You can learn more about why ions form by watching the video at this URL:

http://www.youtube.com/watch?v=CV53wfl-oV8 (9:35)



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Properties of Ions

Ions are highly reactive, especially as gases. They usually react with ions of opposite charge to form neutral compounds. For example, positive sodium ions and negative chloride ions react to form the neutral compound sodium chloride, commonly known as table salt. This occurs because oppositely charged ions attract each other. Ions with the same charge, on the other hand, repel each other. Ions are also deflected by a magnetic field, as you saw above with the northern lights.

Summary

- Atoms have equal numbers of positive protons and negative electrons, so they are neutral in electric charge.
- Atoms can gain or lose electrons and become ions, which are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons.
- The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation or when atoms transfer electrons to or from other atoms.
- Ions are reactive, attracted or repulsed by other charged particles, and deflected by a magnetic field.

Vocabulary

- **ion**: Positively or negatively charged form of an atom that has lost or gained electron(s).

Practice

At the following URL, scroll down to the middle of the page and download "Ion Worksheet." Then fill in the missing information in the worksheet.

http://www.powayusd.com/teachers/kvalentine/worksheetspage.htm

Review

1. Why are atoms neutral in electric charge?
2. Define ion.
3. Compare and contrast cations and anions, and give an example of each.
4. Describe how ions form.
5. List properties of ions.
6. The model in the **Figure** 4.23 represents an atom of lithium (Li). If the lithium atom becomes an ion, which type of ion will it be, a cation or an anion? What will be the electric charge of this ion? What will the ion be named? What symbol will be used to represent it?

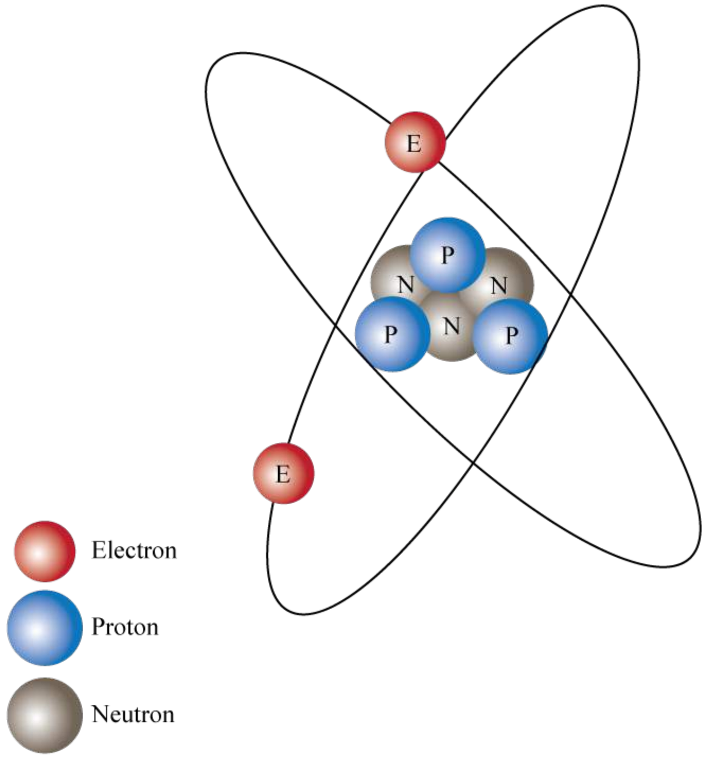


FIGURE 4.23

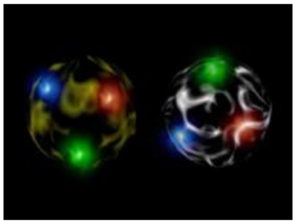
4.9 Protons

- Describe protons.
- State the relationship between protons and elements.
- Identify particles that make up protons.

What Is a Proton?

A **proton** is one of three main particles that make up the atom. The other two particles are the neutron and electron. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one (+1) and a mass of 1 atomic mass unit (amu), which is about 1.67×10^{-27} kilograms. Together with neutrons, they make up virtually all of the mass of an atom. For an excellent video on protons and other fundamental particles in atoms, go to this URL:

http://www.youtube.com/watch?v=Vi91qyjuknM (7:44)



MEDIA

Click image to the left for more content.

Q: How do you think the sun is related to protons?

A: The sun's tremendous energy is the result of proton interactions. In the sun, as well as in other stars, protons from hydrogen atoms combine, or fuse, to form nuclei of helium atoms. This fusion reaction releases a huge amount of energy and takes place in nature only at the extremely high temperatures of stars such as the sun.

Identical Protons, Different Elements

All protons are identical. For example, hydrogen protons are exactly the same as protons of helium and all other elements, or pure substances. However, atoms of different elements have different numbers of protons. In fact, atoms of any given element have a unique number of protons that is different from the numbers of protons of all other elements. For example, a hydrogen atom has just one proton, whereas a helium atom has two protons. The number of protons in an atom determines the electrical charge of the nucleus. The nucleus also contains neutrons, but they are neutral in charge. The one proton in a hydrogen nucleus, for example, gives it a charge of +1, and the two protons in a helium nucleus give it a charge of +2. To learn more about the relationship between protons and elements, go to this URL:

http://www.youtube.com/watch?v=lP57gEWcisY (1:57)

Q: There are six protons in each atom of carbon. What is the electrical charge of a carbon nucleus?

A: The electrical charge of a carbon nucleus is +6.

Summary

- A proton one of three main particles that make up the atom. It is found in the nucleus. It has an electrical charge of one +1 and a mass of 1 atomic mass unit (amu).
- Atoms of any given element have a unique number of protons that is different from the numbers of protons of all other elements.
- Protons consist of fundamental particles called quarks and gluons. Gluons carry the strong nuclear force between quarks, binding them together.

Vocabulary

- **proton**: Positively charged atomic particle inside the nucleus of an atom.

Practice

Do the activity at the URL below for a better appreciation of the size of a proton.

http://www.pbs.org/wgbh/nova/teachers/activities/pdf/3012_elegant_13.pdf

Review

1. Describe protons.
2. What is the relationship between protons and elements?
3. Atoms, which are always neutral in electric charge, contain electrons as well as protons and neutrons. An electron has an electrical charge of -1. If an atom has three electrons, infer how many protons it has.
4. Identify the fundamental particles that make up a proton.

4.10 References

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23. Zachary Wilson. Diagram of a lithium atom. CC BY-NC 3.0

CHAPTER 5 Electrons in Atoms

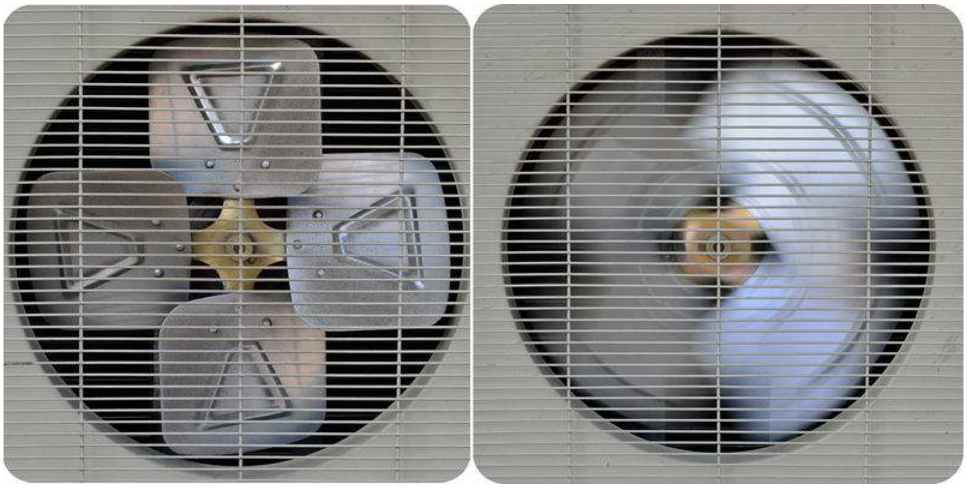
Chapter Outline

You have certainly seen the type of bright, colorful signs shown in the picture above. Have you ever looked closely at a neon sign and wondered what causes them to glow the way that they do? It's all about electrons! In the last chapter, *Atomic Structure*, you learned about some of the earliest atomic models and how they changed over time, ending with the important discovery of the nucleus of the atom by Ernest Rutherford. In this chapter we focus on the electrons. Where are they located? What are their energies? How do they move? Along the way, we will learn about the complex arrangements of electrons in the atoms of each and every element as well as about how the gases contained in a neon sign emit their bright light.

Opening image by Alex Blanck (Flickr: alexblanck). CC−BY−NC−SA 2.0.

5.1 Electron Cloud Atomic Model

- State the main limitation of Bohr's atomic model.
- Describe Schröedinger's work and atomic orbitals.
- Identify the electron cloud model of the atom.



The fan pictured here is turned off in the photo on the left and running at high speed in the photo on the right. In the right-hand photo, the blades are moving too fast for you to see the individual blades. You can't tell where any given blade is at any given moment. In some ways, rapidly moving fan blades are similar to electrons moving about the nucleus of an atom. Like fan blades, electrons move very quickly and we can never tell exactly where they are. If that's the case, how can we represent electrons in models of the atom?

Where Are the Electrons?

Up until about 1920, scientists accepted Niels Bohr's model of the atom. In this model, negative electrons circle the positive nucleus at fixed distances from the nucleus, called energy levels. You can see the model in **Figure** 5.1 for an atom of the element nitrogen. Bohr's model is useful for understanding properties of elements and their chemical interactions. However, it doesn't explain certain behaviors of electrons, except for those in the simplest atom, the hydrogen atom.

What Are the Chances?

In the mid-1920s, an Austrian scientist named Erwin Schröedinger thought that the problem with Bohr's model was restricting the electrons to specific orbits. He wondered if electrons might behave like light, which scientists already knew had properties of both particles and waves. Schrödinger speculated that electrons might also travel in waves. At the URLs below, you can see how some of these waves might look.

- http://en.wikipedia.org/wiki/File:Drum_vibration_mode13.gif
- http://en.wikipedia.org/wiki/File:Drum_vibration_mode23.gif

Q: How do you pin down the location of an electron in a wave?

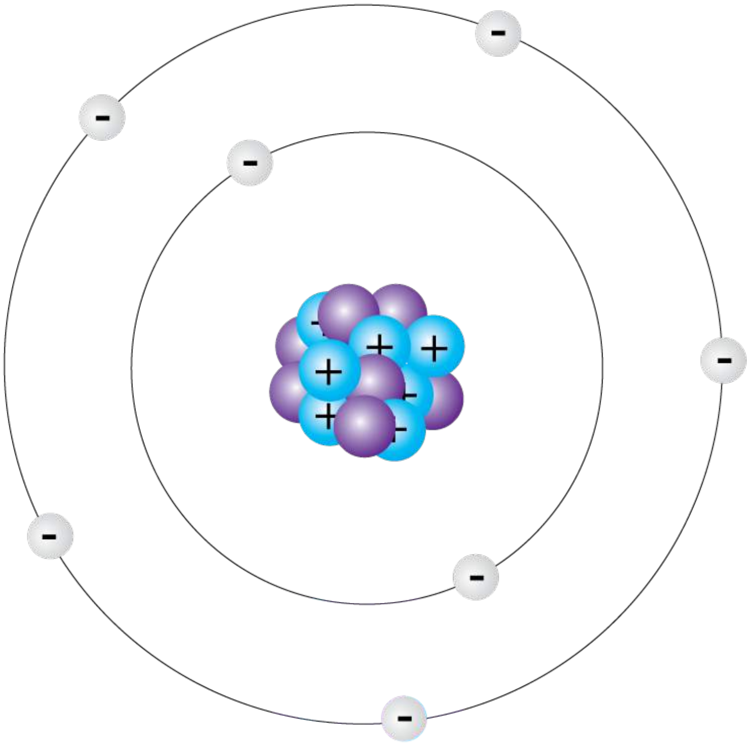


FIGURE 5.1

A: You can't specify the exact location of an electron. However, Schröedinger showed that you can at least determine where an electron is most likely to be.

Schröedinger developed an equation that could be used to calculate the chances of an electron being in any given place around the nucleus. Based on his calculations, he identified regions around the nucleus where electrons are most likely to be. He called these regions **orbitals**. As you can see in the **Figure** 5.2, orbitals may be shaped like spheres, dumbbells, or rings. In each case, the nucleus of the atom is at the center of the orbital. You can see animated models of orbitals this URL: http://www.chemistry.wustl.edu/~edudev/Orbitals/movie.html .

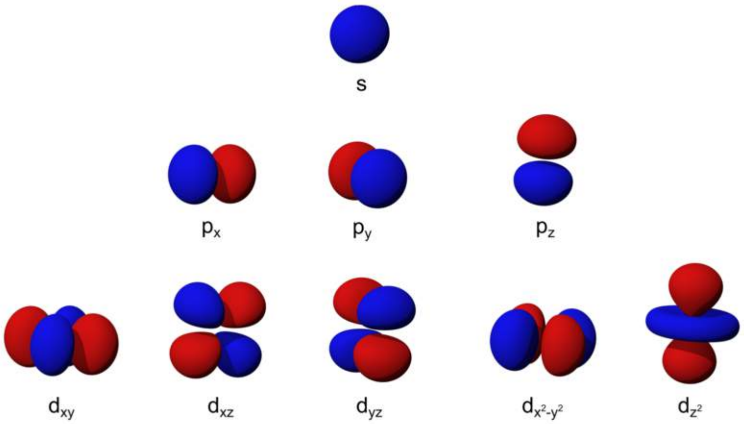


FIGURE 5.2

A Fuzzy Cloud

Schröedinger's work on orbitals is the basis of the modern model of the atom, which scientists call the quantum mechanical model. The modern model is also commonly called the **electron cloud** model. That's because each orbital around the nucleus of the atom resembles a fuzzy cloud around the nucleus, like the ones shown in the **Figure** 5.3 for a helium atom. The densest area of the cloud is where the electrons have the greatest chances of being.

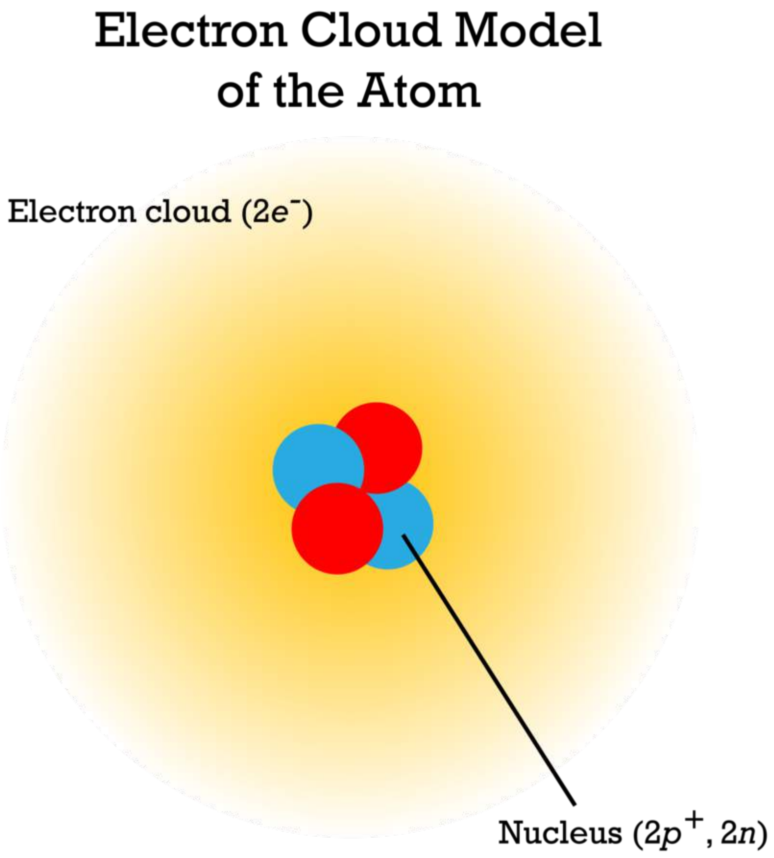


FIGURE 5.3

Q: In the model pictured above, where are the two helium electrons most likely to be?

A: The two electrons are most likely to be inside the sphere closest to the nucleus where the cloud is darkest.

Summary

- Bohr's model of the atom, in which electrons circle the nucleus at fixed energy levels, cannot explain all the behaviors of electrons.
- In the 1920s, Erwin Schröedinger proposed that electrons travel in waves, which means their exact positions cannot be determined. He developed an equation to calculate the chances of an electron being in any given place. Using his equation, he identified regions around the nucleus, called orbitals, where electrons are most likely to be.
- Orbitals are the basis of the electron cloud model of the atom. This model is still accepted by scientists today.

Vocabulary

- **electron cloud**: Area surrounding the nucleus of an atom where electrons are likely to be.
- **orbital**: Region in the electron cloud around the nucleus of an atom where electrons are most likely to be.

Practice

Watch the video about the electron cloud model at the following URL, and then answer the questions below.

http://www.youtube.com/watch?v=qNS83fOFkTM (2:45)

MEDIA

Click image to the left for more content.

1. What influences the movement of electrons in atoms?
2. What is the Heisenberg uncertainty principle?
3. How is the Heisenberg uncertainty principle related to electron orbitals?

Review

1. What is the problem with Bohr's model of the atom?
2. How did Schröedinger resolve this problem?
3. Describe orbitals.
4. Outline the electron cloud model of the atom.

5.2 Light

Lesson Objectives

- Describe the relationships between speed, wavelength, and frequency of light.
- Understand the photoelectric effect and how it is related to the wave-particle duality of light.
- Describe how changes in electron energies lead to atomic emission spectra.
- Describe the Bohr model of the atom.

Lesson Vocabulary

- atomic emission spectrum
- electromagnetic radiation
- electromagnetic spectrum
- excited state
- frequency
- ground state
- photoelectric effect
- photon
- quantum
- wavelength

Properties of Light

The nuclear atomic model proposed by Rutherford was a great improvement over previous models, but it was still not complete. It did not fully explain the location and behavior of the electrons in the vast space outside of the nucleus. For example, it was well known that oppositely charged particles attract one another. Rutherford's model did not explain why the electrons don't simply move toward and eventually collide with the nucleus. A number of experiments were conducted in the early twentieth century that focused on the ability of matter to absorb and emit light. These studies showed that certain phenomena associated with light revealed a great deal about the nature of matter, energy, and atomic structure.

Wave Nature of Light

In order to begin to understand the nature of the electron, we first need to look at the properties of light. Prior to 1900, scientists thought light behaved solely as a wave. As we will see later, this began to change as new experiments demonstrated that light also has some of the characteristics of a particle. First, we will examine the wavelike properties of light.

Visible light is one type of **electromagnetic radiation**, *which is a form of energy that exhibits wavelike behavior as it moves through space.* Other types of electromagnetic radiation include gamma rays, x-rays, ultraviolet light, infrared light, microwaves, and radio waves. **Figure** 5.4 shows the **electromagnetic spectrum**, *which includes all forms of electromagnetic radiation.* Notice that visible light makes up only a very, very small portion of the entire electromagnetic spectrum. All electromagnetic radiation moves through a vacuum at a constant speed of 2.998×10^8 m/s. While the presence of air molecules slows the speed of light by a very small amount, we will still use a value of 3.00×10^8 m/s for the speed of light in air.

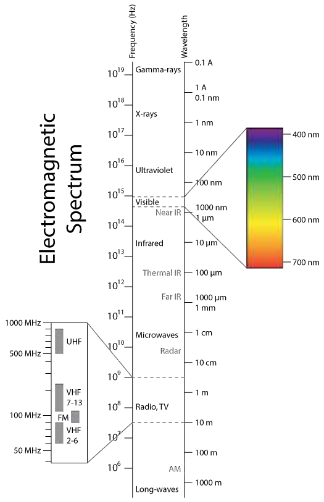


FIGURE 5.4

The electromagnetic spectrum encompasses a very wide range of wavelengths and frequencies. Visible light is only a very small portion of the spectrum, with wavelengths from 400-700 nm.

You can tour the electromagnetic spectrum at http://missionscience.nasa.gov/ems/index.html .

Waves are characterized by their repetitive motion. Imagine a toy boat riding the waves in a wave pool. As the water wave passes under the boat, it moves up and down in a regular and repeated fashion. While the wave travels horizontally, the boat only travels vertically up and down. **Figure** 5.5 shows two examples of waves.

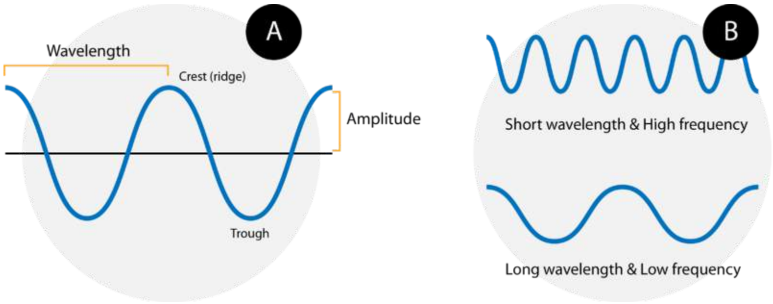


FIGURE 5.5

(A) A wave consists of alternation crests and troughs. The wavelength is defined as the distance between any two consecutive identical points on the waveform. The amplitude is the height of the wave. (B) A wave with a short wavelength (top) has a high frequency because more waves pass a given point in a certain amount of time. A wave with a longer wavelength (bottom) has a lower frequency.

A wave cycle consists of one complete wave –starting at the zero point, going up to a wave crest, going back down to a wave trough, and back to the zero point again. The **wavelength** *of a wave is the distance between any two corresponding points on adjacent waves.* It is easiest to visualize the wavelength of a wave as the distance from one wave crest to the next. In an equation, wavelength is represented by the Greek letter lambda (λ). Depending on the type of wave, wavelength can be measured in meters, centimeters, or nanometers, (1 m = 10^9 nm). The **frequency**, represented by the Greek letter nu (ν), *is the number of waves that pass a certain point in a specified amount of time.* Typically, frequency is measured in units of cycles per second or waves per second. One wave per second is also called a Hertz (Hz) and in SI units is a reciprocal second (s^{-1}).

Figure 5.5 (B) shows an important relationship between the wavelength and frequency of a wave. The top wave clearly has a shorter wavelength than the second wave. However, if you picture yourself at a stationary point watching these waves pass by, more waves of the first kind would pass by in a given amount of time. Thus the frequency of the first waves is greater than that of the second waves. Wavelength and frequency are therefore inversely related. As the wavelength of a wave increases, its frequency decreases. The equation that relates the two is:

$$c = \lambda\nu$$

The variable c is the speed of light. For the relationship to hold mathematically, if the speed of light is used in m/s, the wavelength must be in meters and the frequency in Hertz.

Returning to **Figure** 5.4, you can see how the electromagnetic spectrum displays a wide variation in wavelength and frequency. Radio waves have wavelengths of as long as hundreds of meters, while the wavelength of gamma rays are on the order of 10^{-12} m. The corresponding frequencies range from 10^6 to 10^{21} Hz. Visible light can be split into colors with the use of a prism (**Figure** 5.6), yielding the visible spectrum of light. Red light has the longest wavelength and lowest frequency, while violet light has the shortest wavelength and highest frequency. Visible light wavelength ranges from about 400 –700 nm with frequencies in the range of 10^{14} Hz.

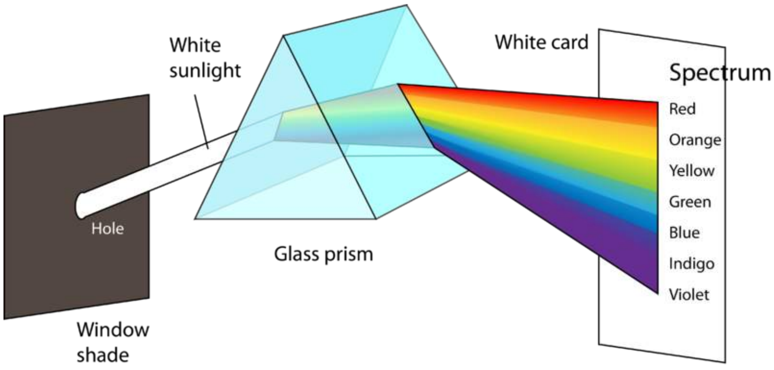


FIGURE 5.6

A small beam of white light is (refracted) bent as it passes through a glass prism. The shorter the wavelength of light, the greater is the refraction, so the light is separated into all its colors.

Sample Problem 5.1: Wavelength and Frequency

The color orange within the visible light spectrum has a wavelength of about 620 nm. What is the frequency of orange light?

Step 1: List the known quantities and plan the problem.

Known

- wavelength (λ) = 620 nm
- speed of light (c) = 3.00×10^8 m/s
- conversion factor 1 m = 10^9 nm

Unknown

- Frequency (ν)

Convert the wavelength to m, then apply the equation $c = \lambda\nu$ and solve for frequency. Dividing both sides of the equation by λ yields:

$$\nu = \frac{c}{\lambda}$$

Step 2: Calculate

$$620 \text{ nm} \times \left(\frac{1 \text{ m}}{10^9 \text{ nm}}\right) = 6.20 \times 10^{-7} \text{ m}$$

$$\nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8 \text{ m/s}}{6.20 \times 10^{-7} \text{ m}} = 4.8 \times 10^{14} \text{ Hz}$$

Step 3: Think about your result.

The value for the frequency falls within the range for visible light.

Practice Problems

1. What is the frequency of radiation of wavelength 2.7×10^{-9} m? In what region of the electromagnetic spectrum is this radiation?
2. Calculate the wavelength in nm of visible light with a frequency of 6.80×10^{14} Hz.

Quantum Physics

German physicist Max Planck (1858-1947) studied the emission of light by hot objects. You have likely seen a heated metal object glow an orange-red color (**Figure** 5.7).

Classical physics, which explains the behavior of large, everyday objects, predicted that a hot object would emit electromagnetic energy in a continuous fashion. In other words, every wavelength of light could possibly be emitted. Instead, what Planck found by analyzing the spectra was that the energy of the hot body could only be lost in small discrete units. A **quantum** *is the minimum quantity of energy that can either be lost or gained by an atom.* An analogy is that a brick wall can only undergo a change in height by units of one or more bricks and not by any possible height. Planck showed that the amount of radiant energy absorbed or emitted by an object is directly proportional to the frequency of the radiation.

$$E = h\nu$$

In the equation, E is the energy, in joules, of a quantum of radiation, ν is the frequency, and h is a fundamental constant called Planck's constant. The value of Planck's constant is $h = 6.626 \times 10^{-34}$ J•s. The energy of any system must increase or decrease in units of $h\nu$. A small energy change results in the emission or absorption of low-frequency radiation, while a large energy change results in the emission or absorption of high-frequency radiation.

The Photoelectric Effect and the Particle Nature of Light

In 1905 Albert Einstein (1879-1955) proposed that light be described as quanta of energy that behave as particles. A **photon** *is a particle of electromagnetic radiation that has zero mass and carries a quantum of energy.* The energy of

FIGURE 5.7

A heated object may glow different colors. The atoms in this piece of metal are releasing energy in discrete units called quanta.

photons of light is quantized according to the $E = h\nu$ equation. For many years light had been described using only wave concepts, and scientists trained in classical physics found this wave-particle duality of light to be a difficult idea to accept. A key experiment that was explained by Einstein using light's particle nature was called the photoelectric effect.

The **photoelectric effect** *is a phenomenon that occurs when light shined onto a metal surface causes the ejection of electrons from that metal.* It was observed that only certain frequencies of light are able to cause the ejection of electrons. If the frequency of the incident light is too low (red light, for example), then no electrons were ejected even if the intensity of the light was very high or it was shone onto the surface for a long time. If the frequency of the light was higher (green light, for example), then electrons were able to be ejected from the metal surface even if the intensity of the light was very low or it was shone for only a short time. This minimum frequency needed to cause electron ejection is referred to as the threshold frequency.

Classical physics was unable to explain the photoelectric effect. If classical physics applied to this situation, the electron in the metal could eventually collect enough energy to be ejected from the surface even if the incoming light was of low frequency. Einstein used the particle theory of light to explain the photoelectric effect as shown in **Figure** 5.8.

Consider the $E = h\nu$ equation. The E is the minimum energy that is required in order for the metal's electron to be ejected. If the incoming light's frequency, ν, is below the threshold frequency, there will never be enough energy to cause electron to be ejected. If the frequency is equal to or higher than the threshold frequency, electrons will be ejected. As the frequency increases beyond the threshold, the ejected electrons simply move faster. An increase in the intensity of incoming light that is above the threshold frequency causes the number of electrons that are ejected to increase, but they do not travel any faster. The photoelectric effect is applied in devices called photoelectric cells, which are commonly found in everyday items such as a calculator which uses the energy of light to generate electricity (**Figure** 5.9).

Run a simulation of the photoelectric effect at http://phet.colorado.edu/en/simulation/photoelectric .

Sample Problem 5.2: Quantized Energy

What is the energy of a photon of green light with a frequency of 5.75×10^{14} Hz?

Step 1: List the known quantities and plan the problem.

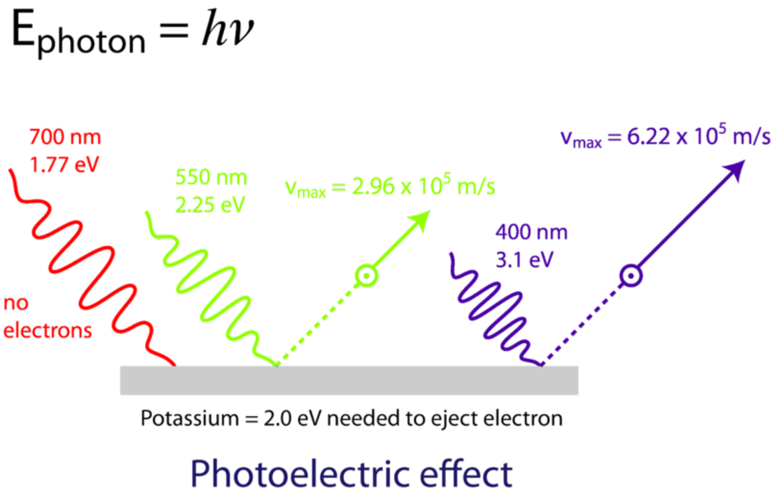


FIGURE 5.8

Low frequency light (red) is unable to cause ejection of electrons from the metal surface. At or above the threshold frequency (green) electrons are ejected. Even higher frequency incoming light (blue) causes ejection of the same number of electrons but with greater speed.

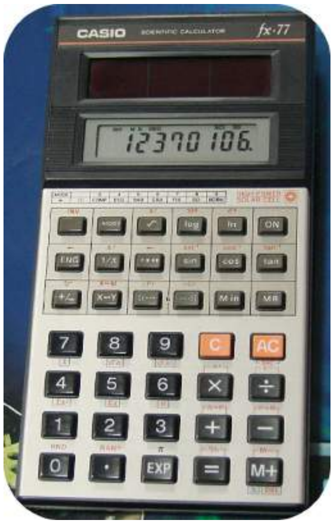


FIGURE 5.9

Photoelectric cells convert light energy into electrical energy which powers this calculator.

Known

- frequency (ν) = 5.75×10^{14} Hz
- Planck's constant (h) = 6.626×10^{-34} J•s

Unknown

- energy (E)

Apply the equation $E = h\nu$ to solve for the energy.

Step 2: Calculate

$$E = (6.626 \times 10^{-34}\ \mathrm{J \cdot s}) \times (5.75 \times 10^{14}\ \mathrm{Hz}) = 3.81 \times 10^{-19}\ \mathrm{J}$$

Step 3: Think about your result.

While the resulting energy may seem very small, this is for only one photon of light. Visible quantities of light consist of huge quantities of photons. Recall that a hertz is equal to a reciprocal second, so the units agree in the equation.

Practice Problem

3. A certain photon of radiation has an energy of 8.72×10^{-21} J. Calculate the frequency and wavelength (in m) of this radiation.

Atomic Emission Spectra

The electrons in an atom tend to be arranged in such a way that the energy of the atom is as low as possible. The **ground state** *of an atom is the lowest energy state of the atom.* When those atoms are given energy, the electrons absorb the energy and move to a higher energy level. These energy levels of the electrons in atoms are quantized, meaning again that the electron must move from one energy level to another in discrete steps rather than continuously. An **excited state** *of an atom is a state where its potential energy is higher than the ground state.* An atom in the excited state is not stable. When it returns back to the ground state, it releases the energy that it had previously gained in the form of electromagnetic radiation.

So how do atoms gain energy in the first place? One way is to pass an electric current through an enclosed sample of a gas at low pressure. Since the electron energy levels are unique for each element, every gas discharge tube will glow with a distinctive color depending on the identity of the gas (**Figure** 5.10).

FIGURE 5.10

Gas discharge tubes are enclosed glass tubes filled with a gas at low pressure through which an electric current is passed. Electrons in the gaseous atoms first become excited, and then fall back to lower energy levels, emitting light of a distinctive color in the process. Shown are gas discharge tubes of helium, neon, argon, krypton, and xenon.

"Neon" signs are familiar examples of gas discharge tubes. However, only signs that glow with the red-orange color seen in the figure are actually filled with neon. Signs of other colors contain different gases or mixtures of gases.

Scientists studied the distinctive pink color of the gas discharge created by hydrogen gas. When a narrow beam of this light was viewed through a prism, the light was separated into four lines of very specific wavelengths (and frequencies since λ and ν are inversely related). An **atomic emission spectrum** *is the pattern of lines formed when* *light passes through a prism to separate it into the different frequencies of light it contains.* **Figure** 5.11 shows the atomic emission spectrum of hydrogen.

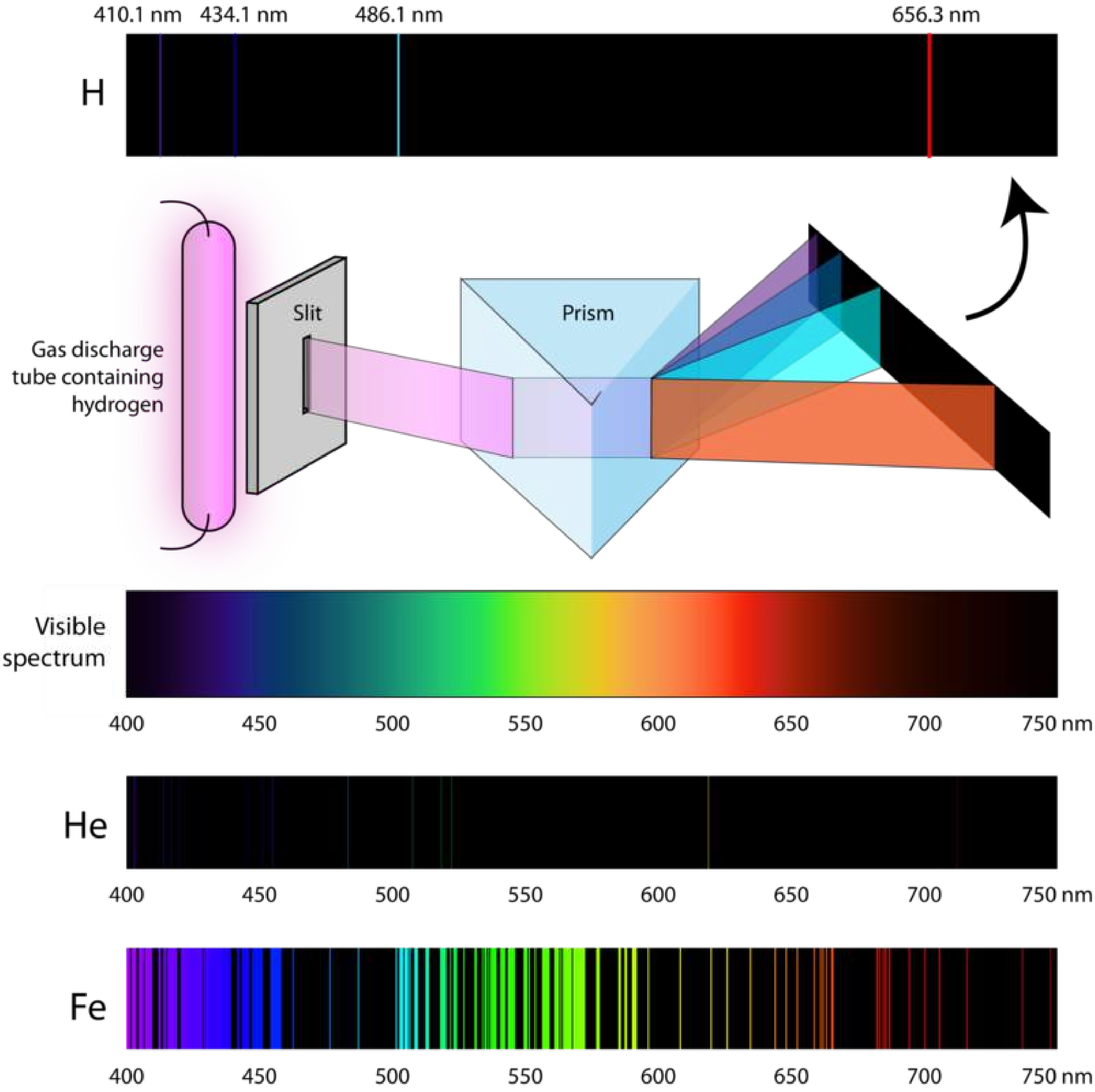


FIGURE 5.11

When light from a hydrogen gas discharge tube is passed through a prism, the light is split into four visible lines. Each of these spectral lines corresponds to a different electron transition from a higher energy state to a lower energy state. Every element has a unique atomic emission spectrum, as shown by the examples of helium (He) and iron (Fe).

The four visible lines of hydrogen's atomic emission spectrum with corresponding wavelengths are violet (410 nm), blue (434 nm), blue-green (486 nm), and red (656 nm). The pinkish color that our eyes see is a combination of these four colors. Every element has its own unique atomic emission spectrum which can be used to identify the gas, a technique which is used extensively in astronomy to identify the composition of distant stars.

Classical theory was unable to explain the existence of atomic emission spectra, also known as line-emission spectra. According to classical physics, a ground state atom would be able to absorb any amount of energy rather than only discrete amounts. Likewise, when the atoms relaxed back to a lower energy state, any amount of energy could be released. This would result in what is known a continuous spectrum, where all wavelengths and frequencies are represented. White light viewed through a prism and a rainbow are examples of continuous spectra. Atomic emission spectra were more proof of the quantized nature of light and led to a new model of the atom based on quantum theory.

Work with simulated discharge lamps at http://phet.colorado.edu/en/simulation/discharge-lamps .

The following video shows a lab demonstration of the atomic emission spectra: http://www.youtube.com/watch?v=955snB6HLB4 (1:39).

MEDIA

Click image to the left for more content.

Bohr Model of the Atom

Figure 5.12 helps explain the process that occurs when an atom is excited and relaxes back to a lower energy.

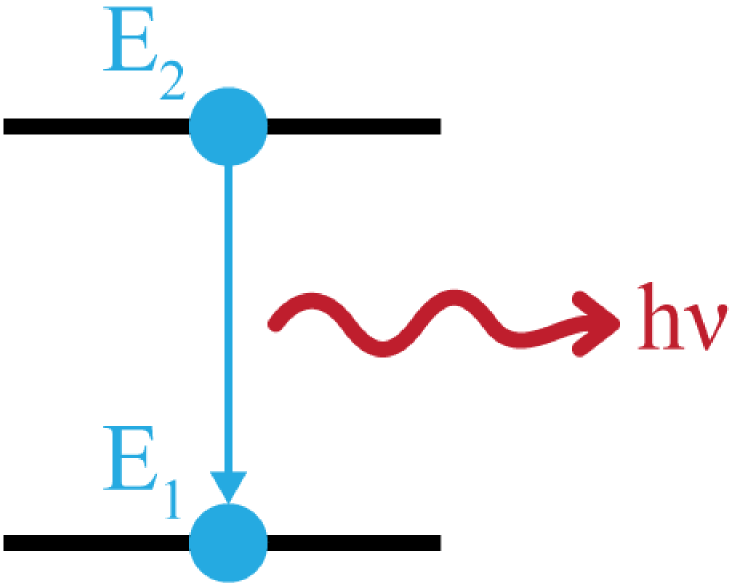


FIGURE 5.12

As an excited atom with energy equal to E_2 falls back down to energy E_1, it releases energy in the form of a photon of electromagnetic energy. The energy of the photon is $E_{photon} = E_2 - E_1 = h\nu$.

The energy states of an atom are indicated by E_1 and E_2, with E_2 being higher in energy. As the atom moves from the E_2 excited state down to the lower energy E_1 state, it loses energy by emitting a photon of radiation. The energy of that photon ($E_{photon} = h\nu$) is equal to the gap in energy between the two states or $E_2 - E_1$. The observation that the atomic emission spectrum of hydrogen consists of only specific frequencies of light indicates that the possible energy states of the hydrogen atom are fixed. This suggested that the electrons in a hydrogen atom were arranged into specific energy levels.

In 1913 Danish physicist Neils Bohr (1885-1962) proposed a model of the atom that explained the hydrogen atomic emission spectrum. According to the Bohr model, often referred to as a planetary model, the electrons encircle the nucleus of the atom in specific allowable paths called orbits. When the electron is in one of these orbits, its energy is fixed. The ground state of the hydrogen atom, where its energy is lowest, is when the electron is in the orbit that is closest to the nucleus. The orbits that are further from the nucleus are all of successively greater energy. The electron is not allowed to occupy any of the spaces in between the orbits. An everyday analogy to the Bohr model is the rungs of a ladder. As you move up or down a ladder, you can only occupy specific rungs and cannot be in the spaces in between rungs. Moving up the ladder increases your potential energy, while moving down the ladder decreases your energy.

Bohr's model explains the spectral lines of the hydrogen atomic emission spectrum. While the electron of the atom remains in the ground state, its energy is unchanged. When the atom absorbs one or more quanta of energy, the electron moves from the ground state orbit to an excited state orbit that is further away. Energy levels are designated with the variable n. The ground state is $n = 1$, the first excited state is $n = 2$, and so on. The energy that is gained by the atom is equal to the difference in energy between the two energy levels. When the atom relaxes back to a lower energy state, it releases energy that is again equal to the difference in energy of the two orbits (**Figure** 5.13).

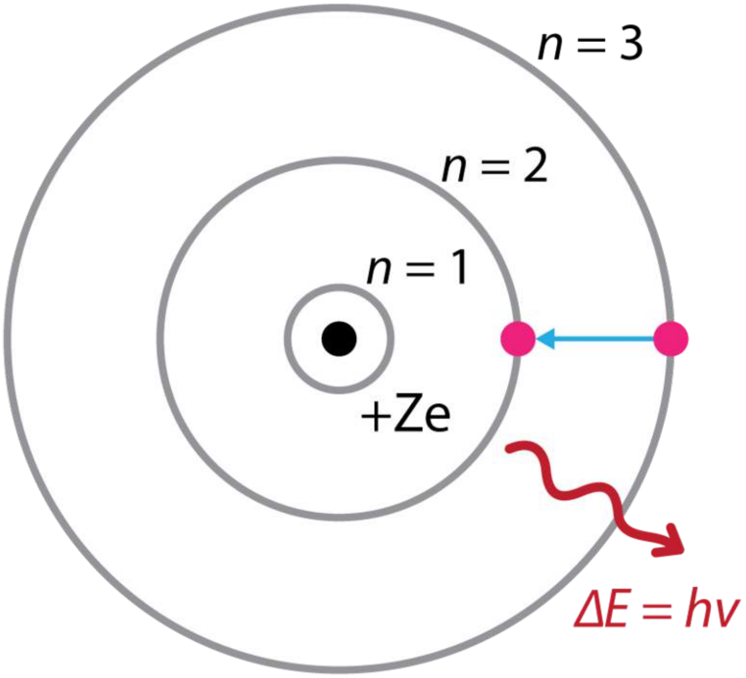


FIGURE 5.13

Bohr model of the atom: electron is shown transitioning from the n = 3 energy level to the n = 2 energy level. The photon of light that is emitted has a frequency that corresponds to the difference in energy between the two levels.

The change in energy, ΔE, then translates to light of a particular frequency being emitted according to the equation $\Delta E = h\nu$. Recall that the atomic emission spectrum of hydrogen had spectral lines consisting of four different frequencies. This is explained in the Bohr model by the realization that the electron orbits are not equally spaced. As the energy increases further and further from the nucleus, the spacing between the levels gets smaller and smaller.

Based on the wavelengths of the spectral lines, Bohr was able to calculate the energies that the hydrogen electron would have in each of its allowed energy levels. He then mathematically showed which energy level transitions corresponded to the spectral lines in the atomic emission spectrum (**Figure** 5.14).

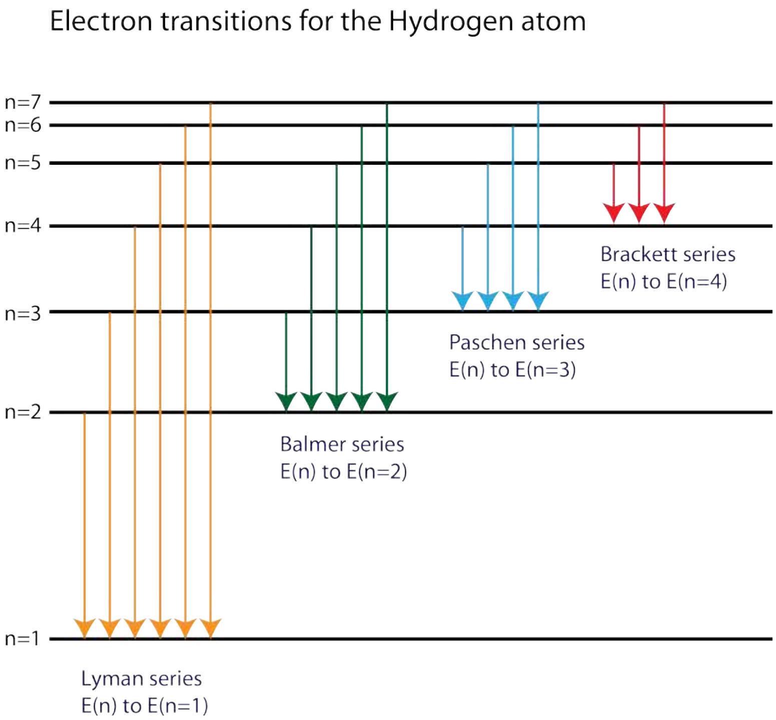


FIGURE 5.14

The electron energy level diagram for the hydrogen atom shows the electron transitions for the Lyman, Balmer, Paschen, and Brackett series. Bohr's atomic model mathematically accounted for the atomic emission spectrum of hydrogen.

He found that the four visible spectral lines corresponded to transitions from higher energy levels down to the second energy level ($n = 2$). This is called the Balmer series. Transitions ending in the ground state ($n = 1$) are called the Lyman series, but the energies released are so large that the spectral lines are all in the ultraviolet region of the spectrum. The transitions called the Paschen series and the Brackett series both result in spectral lines in the infrared region because the energies are too small.

Watch a simulation of the Bohr model of the hydrogen atom at http://www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/chapter8-Animations/ElectronOrbits.html .

Sample Problem 5.3: Spectral Lines of the Hydrogen Atom

In the hydrogen atom, the change in energy (ΔE) for the $n = 4$ to $n = 2$ electron transition is equal to 4.09×10^{-19} J. Calculate the wavelength (in nm) of the spectral line that results from this electron transition and identify its color.

Step 1: List the known quantities and plan the problem.

Known

- $\Delta E = 4.09 \times 10^{-19}$ J
- Planck's constant (h) = 6.626×10^{-34} J•s
- speed of light (c) = 3.00×10^{8} m/s
- conversion factor 1 m = 10^{9} nm

Unknown

- frequency (ν)
- wavelength (λ)

Apply the equation $\Delta E = h\nu$ to solve for the frequency of the emitted light. Then, use the equation $c = \lambda\nu$ to solve for the wavelength in m. Convert to nm.

Step 2: Calculate

$$\nu = \frac{\Delta E}{h} = \frac{4.09 \times 10^{-19}\text{ J}}{6.626 \times 10^{-24}\text{ J} \cdot \text{s}} = 6.17 \times 10^{14}\text{ Hz}$$
$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^{8}\text{ m/s}}{6.17 \times 10^{14}\text{ Hz}} = 4.86 \times 10^{-7}\text{ m}$$
$$4.86 \times 10^{-7}\text{ m} \times \left(\frac{10^{9}\text{ nm}}{1\text{ m}}\right) = 486\text{ nm}$$

Step 3: Think about your result.

The 486 nm spectral line corresponds to a blue-green color. Then n = 3 to n = 2 transition results in a smaller energy release. This in turn yields a lower frequency and the longer wavelength red spectral line. The n = 5 to n = 2 and n = 6 to n = 2 transitions yield the shorter wavelength lines.

Practice Problems

4. The energy change (ΔE) for the n = 2 to n = 1 transition of the Lyman series is 1.64×10^{-18} J. Calculate the wavelength of the resulting spectral line.
5. The visible red spectral line that results from the n = 3 to n = 2 transition of the hydrogen atom is of wavelength 656 nm. Calculate the energy change that produces this spectral line.

Bohr's model was a tremendous success in explaining the spectrum of the hydrogen atom. Unfortunately, when the mathematics of the model was applied to atoms with more than one electron, it was not able to correctly predict the frequencies of the spectral lines. While Bohr's model represented a great advancement in the atomic model and the concept of electron transitions between energy levels is valid, improvements were needed in order to fully understand all atoms and their chemical behavior.

Lesson Summary

- Light is one part of the entire electromagnetic spectrum and its wave nature can be described by its wavelength and frequency.
- The particle nature of light is illustrated by the photoelectric effect and atomic emission spectra.
- Quantum theory states that the electrons of an atom can only exist at discrete energy levels. When electrons transition from a high energy level to a low energy level, energy is released as electromagnetic radiation.
- In the Bohr model of the atom, electrons are allowed to be only at specific distances from the nucleus called orbits. Each orbit has a characteristic energy level.

Lesson Review Questions

Reviewing Concept

1. Answer the following:

   a. List five examples of electromagnetic radiation.
   b. What is the speed of all forms of electromagnetic radiations?
   c. List the colors of the visible spectrum in order from shortest to longest wavelength.

2. Answer the following:
   a. How are the wavelength and frequency of light related?
   b. How are the energy and frequency of light related?
   c. How are the energy and wavelength of light related?
3. Consider the following regions of the electromagnetic spectrum: (i) x-ray, (ii) infrared, (iii) microwave, (iv) visible, (v) radio wave, (vi) gamma ray, (vii) ultraviolet.
   a. Arrange them in order of increasing wavelength.
   b. Arrange them in order of increasing energy.
4. Answer the following:
   a. What is a quantum of energy?
   b. Explain how the stacking of poker chips is related to quantum theory.
5. What happens when a hydrogen atom absorbs a quantum of energy? Use the terms ground state and excited state in your answer.
6. Explain the difference between a continuous spectrum and an atomic emission spectrum.
7. Answer the following:
   a. What is the photoelectric effect?
   b. Why does the photoelectric effect support the idea that light can behave as a particle?
8. Where are electrons located according to the Bohr model?
9. Use the Bohr model to explain how the atomic emission spectrum of hydrogen is produced.
10. Answer the following:
    a. Which series in hydrogen's atomic emission spectrum is composed of visible light?
    b. What are the similarities of each line of that series?

Problems

11. What is the frequency of electromagnetic radiation that has a wavelength of 0.325 cm? In what region of the electromagnetic spectrum is this radiation located?
12. What is the frequency of a photon of light that has an energy of 2.95×10^{-17} J?
13. Answer the following:
    a. Using the equations $E = h\nu$ and $c = \lambda\nu$, derive an equation that expresses E as a function of λ.
    b. Use your equation from part a to determine the energy of a photon of yellow light that has a wavelength of 579 nm.
14. The change in energy of a certain electron transition in the Paschen series is 1.82×10^{-19} J. Calculate the wavelength (in nm) of the spectral line produced by this transition.

Further Reading / Supplemental Links

- We are able to see the visible spectrum without help. To see the infrared spectrum we need special equipment. Have fun watching a demonstration using an Infrared camera at http://www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/irvideo-lg.htm .
- Watch electrons in action lighting a fluorescent light bulb at http://www.youtube.com/watch?v=xFlWGWXjD0o .

- To understand waves better, you can visit these websites:
    - Wave on a String Simulation: http://phet.colorado.edu/en/simulation/wave-on-a-string
    - Radio Waves Electromagnetic Fields Simulation: http://phet.colorado.edu/en/simulation/radio-waves
    - Measuring the Speed of Light With Chocolate: http://education.jlab.org/frost/speed_of_light.html
    - Become Enlightened About Light: http://education.jlab.org/scienceseries/enlightened_about_light.html
    - Bending Light Refraction: http://www.youtube.com/watch?v=_CL8Qzc0O8c
    - Car Light Bends Animation: http://www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/chapter8-Animations/Car-LightBends.html
    - Explore how light interacts with the molecules in our atmosphere: http://phet.colorado.edu/en/simulation/molecules-and-light
- Compare the emission spectra of the sun with visible light at http://phet.colorado.edu/en/simulation/blackbody-spectrum .
- Get fired up at the interactive website: Anatomy of a Firework .
- Below are more links to simulations and videos of a lab activities about atomic spectra:
    - Identifying Salts: http://freezeray.com/flashFiles/identifyingSalts.htm
    - Cobalt Blue Glass Sodium Flame Test: http://www.youtube.com/watch?v=wTxPLqP5yGE
    - Flame Tests of Metal Ions: http://www.youtube.com/watch?v=PJ0Vz7Wv9Rg
    - Flame Tests of LiCl and $LiNO_3$: http://www.youtube.com/watch?v=JvuS3J5M9Cw

5.3 The Quantum Mechanical Model

Lesson Objectives

- Understand the de Broglie wave equation and how it illustrates the wave nature of the electron.
- Explain the difference between quantum mechanics and classical mechanics.
- Understand how the Heisenberg uncertainty principle and Schrödinger's wave equation led to the idea of atomic orbitals.
- Know the four quantum numbers and their significance to an atom's electron arrangement.
- Describe the interrelationships of principal energy level, sublevel, orbital and electron spin and how they relate to the number of electrons of an atom.

Lesson Vocabulary

- angular momentum quantum number
- Heisenberg uncertainty principle
- magnetic quantum number
- orbital
- principal quantum number
- spin quantum number
- quantum mechanical model
- quantum mechanics
- quantum numbers

Wave Nature of the Electron

Bohr's model of the atom was valuable in demonstrating how electrons were capable of absorbing and releasing energy and how atomic emission spectra were created. However, the model did not really explain why electrons should exist only in fixed circular orbits rather than being able to exist in a limitless number of orbits all with different energies. In order to explain why atomic energy states are quantized, scientists needed to rethink the way in which they viewed the nature of the electron and its movement.

de Broglie Wave Equation

Planck's investigation of the emission spectra of hot objects and the subsequent studies into the photoelectric effect had proven that light was capable of behaving both as a wave and as a particle. It seemed reasonable to wonder if electrons could also have a dual wave-particle nature. In 1924, French scientist Louis de Broglie (1892-1987) derived an equation that described the wave nature of any particle. Particularly, the wavelength (λ) of any moving object is given by:

$$\lambda = \frac{h}{mv}$$

In this equation, h is Planck's constant, m is the mass of the particle in kg, and v is the velocity of the particle in m/s. The problem below shows how to calculate the wavelength of the electron.

Sample Problem 5.4: de Broglie Equation

An electron of mass 9.11×10^{-31} kg moves at nearly the speed of light. Using a velocity of 3.00×10^{8} m/s, calculate the wavelength of the electron.

Step 1: List the known quantities and plan the problem.

Known

- mass (m) = 9.11×10^{-31} kg
- Planck's constant (h) = 6.626×10^{-34} J•s
- velocity (v) = 3.00×10^{8} m/s

Unknown

- wavelength (λ)

Apply the de Broglie wave equation $\lambda = \frac{h}{mv}$ to solve for the wavelength of the moving electron.

Step 2: Calculate

$$\lambda = \frac{h}{mv} = \frac{6.626 \times 10^{-34}\ \text{J} \cdot \text{s}}{(9.11 \times 10^{-31}\ \text{kg}) \times (3.00 \times 10^{8}\ \text{m/s})} = 2.42 \times 10^{-12}\ \text{m}$$

Step 3: Think about your result.

This very small wavelength is about $1/20^{th}$ of the diameter of a hydrogen atom. Looking at the equation, as the speed of the electron decreases, its wavelength increases. The wavelengths of everyday large objects with much greater masses should be very small.

Practice Problem

1. Calculate the wavelength of a 0.145 kg baseball thrown at a speed of 40 m/s.

The above practice problem results in an extremely short wavelength on the order of 10^{-34} m. This wavelength is impossible to detect even with advanced scientific equipment. Indeed, while all objects move with wavelike motion, we never notice it because the wavelengths are far too short. On the other hand, particles with measurable wavelengths are all very small. However, the wave nature of the electron proved to be a key development in a new understanding of the nature of the electron. An electron that is confined to a particular space around the nucleus of an atom can only move around that atom in such a way that its electron wave "fits" the size of the atom correctly (**Figure** 5.15). This means that the frequencies of electron waves are quantized. Based on the $E = h\nu$ equation, the quantized frequencies means that electrons can only exist in an atom at specific energies, as Bohr had previously theorized.

The study of motion of large objects such as baseballs is called mechanics, or more specifically classical mechanics. Because the quantum nature of the electron and other tiny particles moving at high speeds, classical mechanics is inadequate to accurately describe their motion. **Quantum mechanics** *is the study of the motion of objects that are*

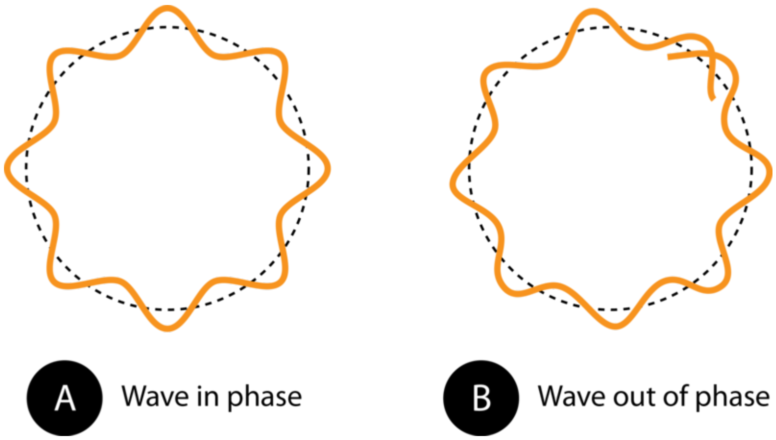


FIGURE 5.15

The circumference of the orbit in (A) allows the electron wave to fit perfectly into the orbit. This is an allowed orbit. In (B), the electron wave does not fit properly into the orbit, so this orbit is not allowed.

atomic or subatomic in size and thus demonstrate wave-particle duality. In classical mechanics, the size and mass of the objects involved effectively obscures any quantum effects so that such objects appear to gain or lose energies in any amounts. Particles whose motion is described by quantum mechanics gain or lose energy in the small pieces called quanta.

Heisenberg Uncertainty Principle

Another feature that is unique to quantum mechanics is the uncertainty principle. The **Heisenberg Uncertainty Principle** *states that it is impossible to determine simultaneously both the position and the velocity of a particle.* The detection of an electron, for example, would be made by way of its interaction with photons of light. Since photons and electrons have nearly the same energy, any attempt to locate an electron with a photon will knock the electron off course, resulting in uncertainty about where the electron is located (**Figure** 5.16). We do not have to worry about the uncertainty principle with large everyday objects because of their mass. If you are looking for something with a flashlight, the photons coming from the flashlight are not going to cause the thing you are looking for to move. This is not the case with atomic-sized particles, leading scientists to a new understanding about how to envision the location of the electrons within atoms.

You can see a funny, animated explanation of Heisenberg's Uncertainty Principle at http://video.pbs.org/video/18121247 .

Quantum Mechanical Model

In 1926, Austrian physicist Erwin Schrödinger (1887-1961) used the wave-particle duality of the electron to develop and solve a complex mathematical equation that accurately described the behavior of the electron in a hydrogen atom. The **quantum mechanical model** *of the atom comes from the solution to Schrödinger's equation.* Quantization of electron energies is a requirement in order to solve the equation. This is unlike the Bohr model, in which quantization was simply assumed with no mathematical basis.

Recall that in the Bohr model, the exact path of the electron was restricted to very well defined circular orbits around the nucleus. The quantum mechanical model is a radical departure from that. Solutions to the Schrödinger wave equation, called wave functions, give only the probability of finding an electron at a given point around the nucleus. Electrons do not travel around the nucleus in simple circular orbits.

The location of the electrons in the quantum mechanical model of the atom is often referred to as an electron cloud.

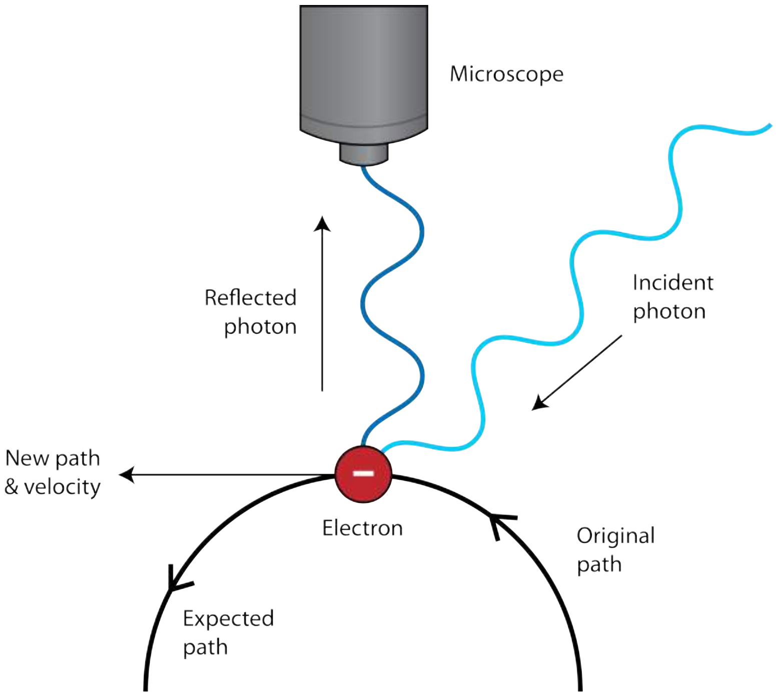


FIGURE 5.16

Heisenberg Uncertainty Principle: The observation of an electron with a microscope requires reflection of a photon off of the electron. This reflected photon causes a change in the path of the electron.

The electron cloud can be thought of in the following way. Imagine placing a square piece of paper on the floor with a dot in the circle representing the nucleus. Now take a marker and drop it onto the paper repeatedly, making small marks at each point the marker hits. If you drop the marker many, many times, the overall pattern of dots will be roughly circular. If you aim toward the center reasonably well, there will be more dots near the nucleus and progressively fewer dots as you move away from it. Each dot represents a location where the electron *could be* at any given moment. Because of the uncertainty principle, there is no way to know exactly where the electron is. An electron cloud has variable densities: a high density where the electron is most likely to be and a low density where the electron is least likely to be (**Figure** 5.17).

In order to specifically define the shape of the cloud, it is customary to refer to the region of space within which there is a 90% probability of finding the electron. This is called an **orbital**, *the three-dimensional region of space that indicates where there is a high probability of finding an electron.*

Atomic Orbitals and Quantum Numbers

Solutions of the Schrödinger wave equation give the energies that an electron is allowed to have. The mathematical representation of those energies results in regions of space called orbitals, which, as we will soon see, can have different sizes and shapes. In order to completely describe the orbitals and the electrons which occupy them, scientists use quantum numbers. **Quantum numbers** *specify the properties of the atomic orbitals and the electrons in those orbitals.* Understanding quantum numbers is helped by an analogy. Let's say you are attending a basketball game. Your ticket may specify a gate number, a section number, a row, and a seat number. No other ticket can have the same four parts to it. It may have the same gate, section, and seat number, but if so it would have to be in a different row. There are also four quantum numbers which describe each and every electron in every atom. No two electrons in a given atom can have the same four quantum numbers. We will describe each of these quantum numbers separately.

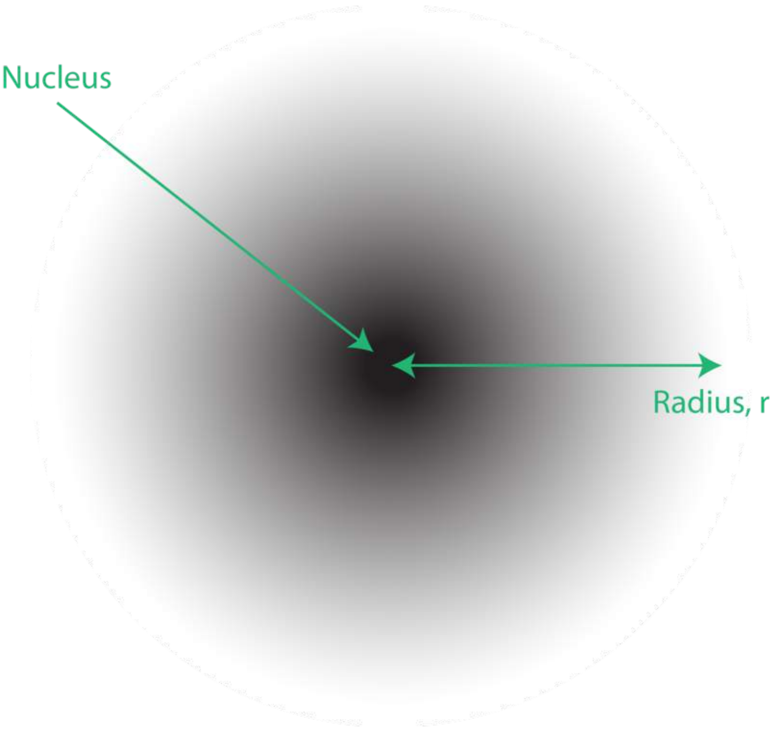


FIGURE 5.17

An electron cloud: the darker region nearer the nucleus indicates a high probability of finding the electron, while the lighter region further from the nucleus indicates a lower probability of finding the electron.

Principal Quantum Number

The **principal quantum number** *is symbolized by the letter n and is the principal or main energy level occupied by the electron.* The value of n begins with $n = 1$, which is the lowest in energy and is located closest to the nucleus. As the n value increases to $n = 2$, 3, and so on, the distance from the nucleus increases. The principal quantum number is essentially the same as the energy levels in the Bohr model of the atom that were used to explain atomic emission spectra. More than one electron may occupy a given principal energy level, but the specific number varies depending on which energy level.

Angular Momentum Quantum Number

The **angular momentum quantum number** *is symbolized by the letter l and indicates the shape of the orbital.* For each given principal energy level, orbitals of different shapes exist, and these are slightly different in energy and are referred to as energy sublevels. The number of sublevels varies depending on the value of n. Specifically, the number of sublevels possible is equal to the value of n. In other words, when $n = 1$, there is only one sublevel. When $n = 2$, there are two sublevels. The quantum number l is an integer which varies from 0 up to a value equal to n - 1. In other words, if $n = 1$, the only possible value of l is 0. If $n = 4$, then l can have a value of 0, 1, 2, or 3. Finally, each of the orbitals represented by the various sublevels has a letter designation: s, p, d, or f. If $l = 0$, the orbital is an s orbital. If $l = 1$, the orbital is a p orbital. See **Table** 5.1 for a summary.

TABLE 5.1: **Principal Energy Levels and Sublevels**

Principal Energy Level	Number of Possible Sublevels	Possible Angular Momentum Quantum Numbers	Orbital Designation by Principal Energy Level and Sublevel
$n = 1$	1	$l = 0$	1s

TABLE 5.1: (continued)

Principal Energy Level	Number of Possible Sublevels	Possible Angular Momentum Quantum Numbers	Orbital Designation by Principal Energy Level and Sublevel
$n = 2$	2	$l = 0$ $l = 1$	2*s* 2*p*
$n = 3$	3	$l = 0$ $l = 1$ $l = 2$	3*s* 3*p* 3*d*
$n = 4$	4	$l = 0$ $l = 1$ $l = 2$ $l = 3$	4*s* 4*p* 4*d* 4*f*

From the table you can see that in the 1^{st} principal energy level (n = 1) there is only one sublevel possible –an *s* sublevel. In the 2^{nd} principal energy level (n = 2), there are two sublevels possible –the *s* and *p* sublevels. This continues through the 3^{rd} and 4^{th} principal energy levels adding in the *d* and the *f* sublevels. In general for the n^{th} principal energy level there are *n* sublevels available. The order of the sublevels is always the same.

Magnetic Quantum Number

As mentioned above, each of the different orbital types has a different shape. The *s* orbitals are spherical in shape (**Figure** 5.18), the *p* orbitals are dumbbell shaped (**Figure** 5.19), and the *d* and *f* orbitals are more complex with multiple lobes (**Figure** 5.20).

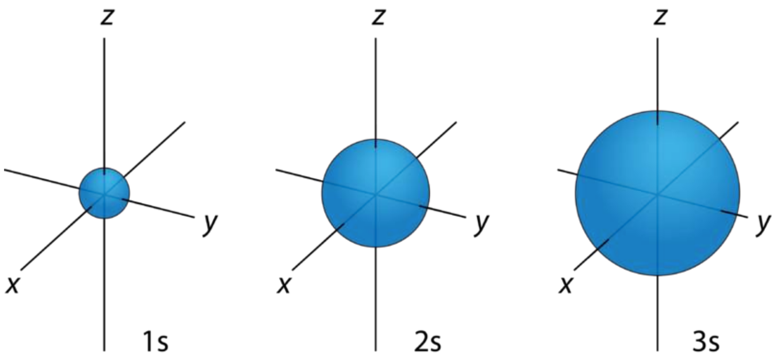


FIGURE 5.18

The *s* orbitals are spherical in shape and centered on the nucleus. The size of the orbital increases with the higher principal energy level.

View the *p* orbitals at http://www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/chapter8-Animations/P-orbitalDiagram.html .

The **magnetic quantum number** *is symbolized by the letter* m_l *and indicates the orientation of the orbital around the nucleus.* Because an *s* orbital is spherical in shape and centered on the nucleus, it only has one possible orientation. Thus, it has only one possible value of the magnetic quantum number and that is $m_l = 0$. Within each *s* sublevel, there is just one *s* orbital. The larger is the principal energy level (n), the larger is the size of the spherical *s* orbital. As seen in **Figure** 5.19, the dumbbell-shaped *p* orbitals have three possible orientations. In one orientation, called the p_x, the lobes of the orbital lie along the defined *x*-axis. In the p_y orbital, they lie along the *y*-axis, or at a 90° angle to the p_x orbital. Finally, the p_z orbital lies along the *z*-axis or 90° relative to the other two orbitals. In any *p* sublevel, there are always these three orbitals and that means that there are three possible values of the magnetic

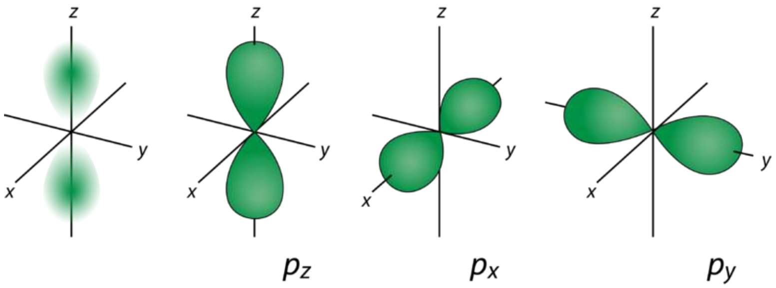


FIGURE 5.19

The *p* orbitals are dumbbell-shaped. (a) shows the electron density distribution, while (b) shows the orbitals that result. The three different *p* orbitals are identical in shape, but oriented along the *x*, *y*, or *z* axis.

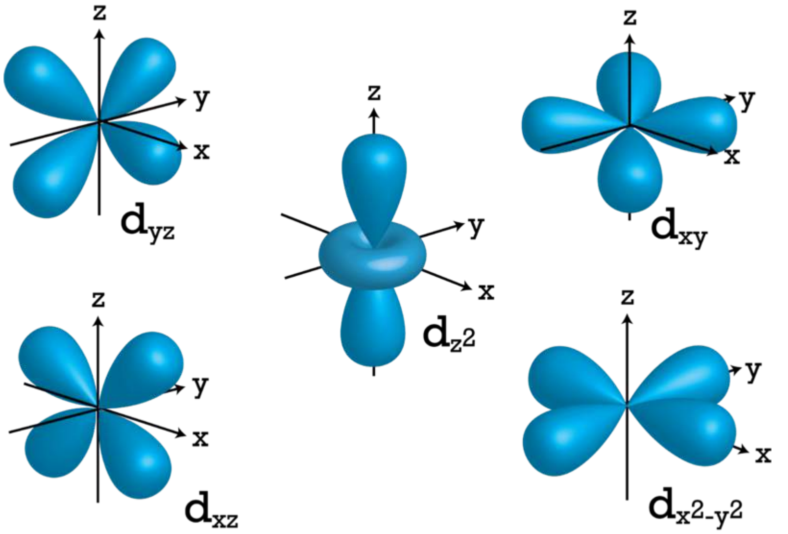


FIGURE 5.20

The five *d* orbitals have more complex shapes. Four of them have a 4-lobed appearance with different orientations, while the $d_z{}^2$ orbital has a more complex shape.

quantum number: $m_l = -1$, $m_l = 0$, and $m_l = +1$. There is no particular relationship between the coordinates (x, y, and z) and the m_l value.

There are five different *d* orbitals within each *d* sublevel. The corresponding magnetic quantum numbers are $m_l = -2$, $m_l = -1$, $m_l = 0$, $m_l = +1$, and $m_l = +2$. Finally, the pattern continues with the *f* sublevel containing seven possible *f* orbitals and m_l values ranging from -3 to $+3$. Ground states of all known elements can be described with these four sublevels.

Spin Quantum Number

Experiments show that electrons spin on their own internal axis, much as Earth does. The spinning of a charged particle creates a magnetic field. The orientation of that magnetic field depends upon the direction that the electron is spinning, either clockwise or counterclockwise. The **spin quantum number** *is symbolized by the letter* m_s *and indicated the direction of electron spin.* The values are $m_s = +\frac{1}{2}$ and $m_s = -\frac{1}{2}$. Each orbital, no matter its shape or orientation can hold a maximum of two electrons, which must have opposite spins. **Table** 5.2 shows a summary of energy levels and sublevels and the number of electrons that can possibly exist in the various states.

TABLE 5.2: Electron Arrangement Within Energy Levels

Principal Quantum Number (n)	Allowable Sublevels	Number of Orbitals per Sublevel	Number of Orbitals per Principal Energy Level	Number of Electrons per Sublevel	Number of Electrons per Principal Energy Level
1	*s*	1	1	2	2
2	*s* *p*	1 3	4	2 6	8
3	*s* *p* *d*	1 3 5	9	2 6 10	18
4	*s* *p* *d* *f*	1 3 5 7	16	2 6 10 14	32

Notice that the total number of allowable orbitals in each principal energy level (n) is equal to n^2. That is, when n = 1, there are 1^2 = 1 orbital possible. When n = 2, there are 2^2 = 4 orbitals possible, etc. Since each orbital holds two electrons, the number of electrons that can exist in a given principal energy level is equal to $2n^2$.

Lesson Summary

- de Broglie showed that electrons have both a wave nature and a particle nature.
- The behavior of atomic and subatomic sized particles is explained by quantum mechanics, where energy is gained and lost in small discrete amounts.
- The Heisenberg uncertainty principal showed that is not possible to know the location of an electron at any precise moment.
- The Schrödinger wave equation proved mathematically that the energy of an electron must be quantized.
- The quantum mechanical model of the atom describes the probability that an atom's electrons will be located within certain regions called orbitals.
- Electrons arrangements are governed by four quantum numbers. These correspond to the principal energy level, energy sublevel, orbital, and electron spin.

Lesson Review Questions

Reviewing Concepts

1. How is the wavelength of a moving object related to its mass?
2. Why is the de Broglie wave equation meaningful only for submicroscopic particles such as atoms and electrons and not for larger everyday objects?
3. How does the Heisenberg uncertainty principle affect the way in which electron locations are viewed in the quantum mechanical model as compared to the Bohr model?
4. What is an atomic orbital?
5. How many quantum numbers are used to describe each electron in an atom?

6. In what two ways is an electron that occupies the $n = 2$ principal energy level different than an electron which occupies the $n = 1$ principal energy level?
7. Identify which quantum number describes each of the following.
   a. the orientation of an orbital in space
   b. the direction of electron spin
   c. the main energy of an electron
   d. the shape of an orbital

Problems

8. What is the wavelength (in nm) of an electron moving at 250 m/s? In what region of the electromagnetic spectrum is this electron wave?
9. What are the possible values of l that an electron in the $n = 3$ principal energy level can have? Which sublevel does each of those l values represent?
10. What are the possible values of m_l that an electron in the p sublevel can have?
11. Which of the following combinations of principal energy level and sublevel *cannot* exist?
    a. $4d$
    b. $3f$
    c. $1p$
    d. $2s$
    e. $3p$
    f. $2d$
12. How is a p_z orbital different from a p_x orbital? How are they the same?
13. How many orbitals are found in each of the following?
    a. any s sublevel
    b. any f sublevel
    c. the $n = 2$ principal energy level
    d. the $4p$ sublevel
    e. the $n = 3$ principal energy level
14. How many electrons in each of the following?
    a. any orbital
    b. the $n = 2$ principal energy level
    c. the $3d$ sublevel
    d. the $n = 4$ principal energy level
    e. the $2p$ sublevel

5.4 Electron Arrangement in Atoms

Lesson Objectives

- Understand how to apply the Aufbau principle, Pauli exclusion principle, and Hund's rule to ground state electron configurations.
- Write correct orbital filling diagrams and electron configurations for all elements.
- Use the noble gas configuration shorthand method.
- Be able to determine the number of valence electrons and the number of unpaired electrons in any atom.
- Understand that some electron configurations are exceptions to the normal Aufbau process.

Lesson Vocabulary

- Aufbau principle
- electron configuration
- Hund's rule
- noble gas configuration
- Pauli exclusion principle
- valence electron

Electron Configurations

The quantum mechanical model provides what is now recognized as the modern and accepted model of the atom. An atom's **electron configuration** *is the arrangement of all of the electrons of that atom.* Since every element has a different number of electrons, each has a unique electron configuration. Recall that the natural state for all systems is to be in the lowest energy state possible. Thus, the ground state electron configuration for an element is the lowest-energy arrangement of electrons possible for that element. The basis for determining electron configurations is the quantum number guidelines learned in the previous lesson, "The Quantum Mechanical Model," along with a few basic rules.

Aufbau Principle

In order to create ground state electron configurations for any element, it is necessary to know the way in which the atomic sublevels are organized in order of increasing energy. **Figure** 5.21 shows the order of increasing energy of the sublevels.

The lowest energy sublevel is always the 1*s* sublevel, which consists of one orbital. The single electron of the hydrogen atom will occupy the 1s orbital when the atom is in its ground state. As we proceed with atoms with multiple electrons, those electrons are added to the next lowest sublevel: 2*s*, 2*p*, 3*s*, and so on. The **Aufbau principle** *states that an electron occupies orbitals in order from lowest energy to highest.* The Aufbau principle is sometimes

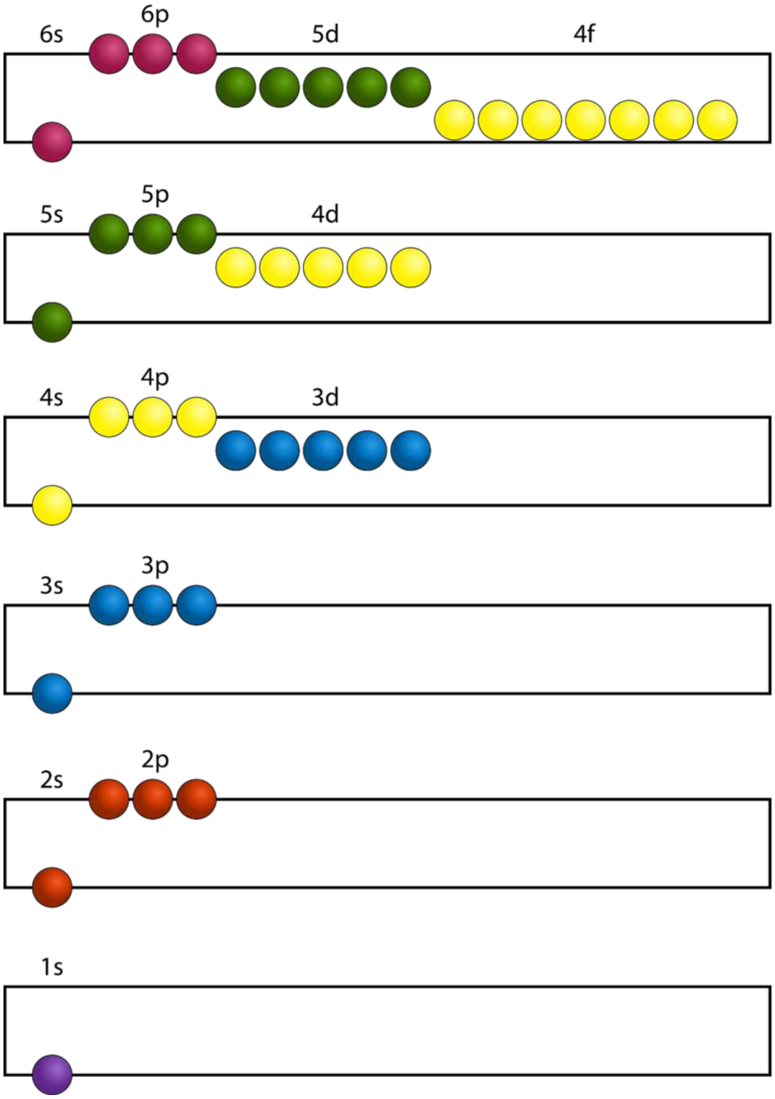


FIGURE 5.21

Electrons are added to atomic orbitals in order from low energy to high according to the Aufbau principle. Principal energy levels are color coded, while sublevels are grouped together and each circle represents an orbital capable of holding two electrons.

referred to as the "building-up" principle. It is worth noting that in reality atoms are not built by adding protons and electrons one at a time and that this method is merely an aid for us to understand the end result.

As seen in **Figure** 5.21, the energies of the sublevels in different principal energy levels eventually begin to overlap. After the $3p$ sublevel, it would seem logical that the $3d$ sublevel should be the next lowest in energy. However, the $4s$ sublevel is slightly lower in energy than the $3d$ sublevel and thus fills first. Following the filling of the $3d$ sublevel is the $4p$, then the $5s$ and the $4d$. Note that the $4f$ sublevel does not fill until just after the $6s$ sublevel. **Figure** 5.22 is a useful and simple aid for keeping track of the order of fill of the atomic sublevels.

Pauli Exclusion Principle

Recall that every orbital, no matter which type, is capable of containing two electrons and that each must have opposite spin. This leads to the **Pauli exclusion principle**, *which states that no two electrons in an atom can have the same set of four quantum numbers*. The energy of the electron is specified by the principal, angular momentum, and magnetic quantum numbers. The two values of the spin quantum number allow each orbital to hold two electrons. **Figure** 5.23 shows how the electrons are indicated in a diagram.

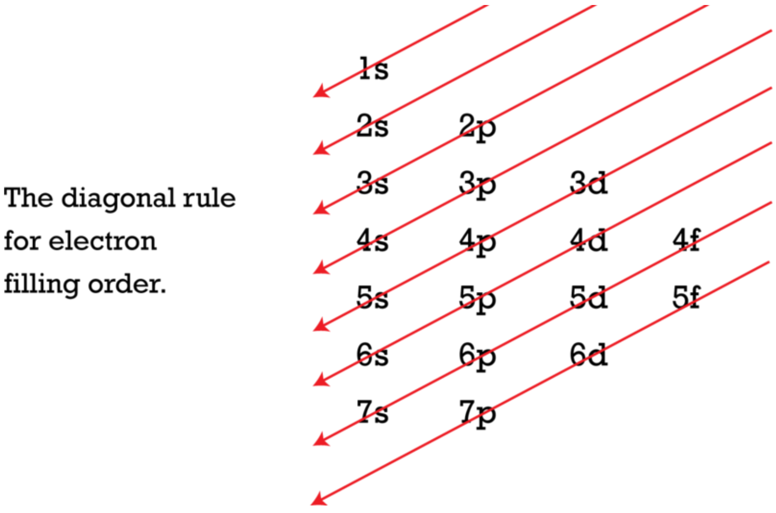


FIGURE 5.22
The Aufbau principle is illustrated in the diagram by following each red arrow in order from top to bottom: 1*s*, 2*s*, 2*p*, 3*s*, etc.

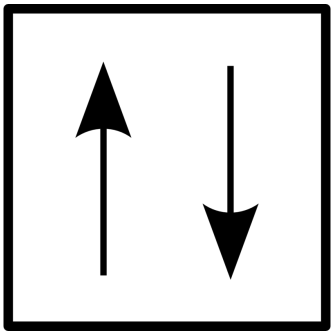


FIGURE 5.23
In an orbital filling diagram, a square represents an orbital, while arrows represent electrons. An arrow pointing upward represents one spin direction, while an arrow pointing downward represents the other spin direction.

View an animation of electron spin at http://www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/chapter8-Animations/ElectronSpin.html .

Hund's Rule

The last of the three rules for constructing electron arrangements requires electrons to be placed one at a time in a set of orbitals within the same sublevel. This minimizes the natural repulsive forces that one electron has for another. **Hund's rule** *states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron and that each of the single electrons must have the same spin*. **Figure** 5.24 shows how a set of three *p* orbitals is filled with one, two, three, and four electrons.

Orbital Filling Diagrams

An orbital filling diagram is the more visual way to represent the arrangement of all the electrons in a particular atom. In an orbital filling diagram, the individual orbitals are shown as circles (or squares) and orbitals within a sublevel are drawn next to each other horizontally. Sublevels can be shown as in **Figure** 5.21 with increasing energy proceeding up the page, or to save space all sublevels can simply be shown horizontally one after the other. Each sublevel is labeled by its principal energy level and sublevel. Electrons are indicated by arrows inside the circles. An arrow pointing upwards indicates one spin direction, while a downward pointing arrow indicates the other direction. The orbital filling diagrams for hydrogen, helium, and lithium are shown below.

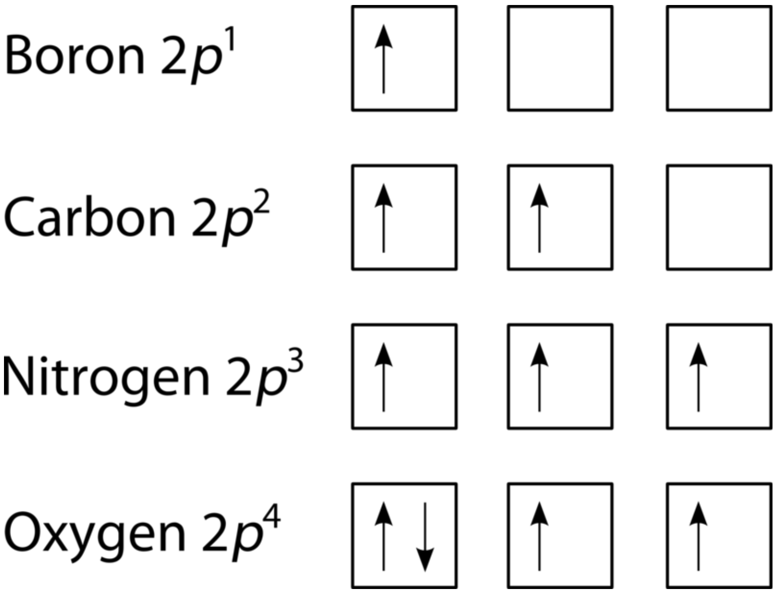
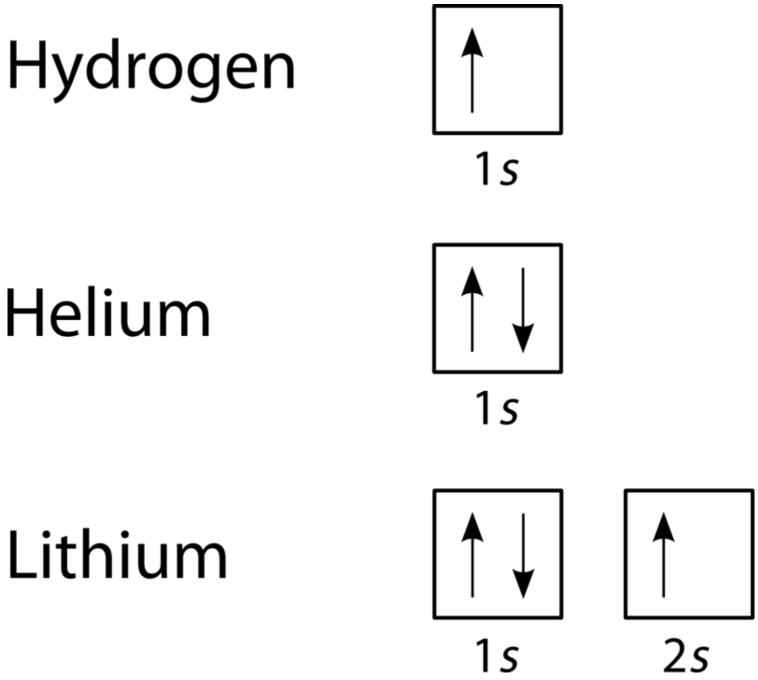


FIGURE 5.24

The $2p$ sublevel, for the elements boron (Z = 5), carbon (Z = 6), nitrogen (Z = 7), and oxygen (Z = 8): According to Hund's rule, as electrons are added to a set of orbitals of equal energy, one electron enters each orbital before any orbital receives a second electron.



According to the Aufbau process, sublevels and orbitals are filled with electrons in order of increasing energy. Since the s sublevel consists of just one orbital, the second electron simply pairs up with the first electron as in helium. The next element is lithium and necessitates the use of the next available sublevel, the $2s$.

Electron Configuration Notation

Electron configuration notation eliminates the circles and arrows of orbital filling diagrams. Each occupied sublevel designation is written followed by a superscript that is the number of electrons in that sublevel. For example, the hydrogen configuration is $1s^1$, while the helium configuration is $1s^2$. Multiple occupied sublevels are written one after another. The electron configuration of lithium is $1s^22s^1$. The sum of the superscripts in an electron configuration is equal to the number of electrons in that atom, which is in turn equal to its atomic number.

Sample Problem 5.5: Orbital Filling Diagrams and Electron Configurations

Draw the orbital filling diagram for carbon and write its electron configuration.

Step 1: List the known quantities and plan the problem.

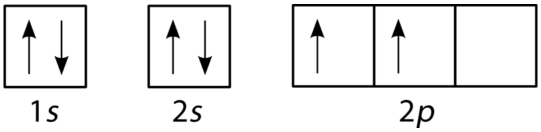
Known

- atomic number of carbon, $Z = 6$

Use the order of fill diagram to draw an orbital filling diagram with a total of six electrons. Follow Hund's rule. Write the electron configuration.

Step 2: Construct diagram

- Orbital filling diagram:



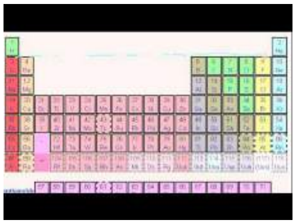
- Electron configuration: $1s^2 2s^2 2p^2$

Step 3: Think about your result.

Following the $2s$ sublevel is the $2p$, and p sublevels always consist of three orbitals. All three orbitals need to be drawn even if one or more is unoccupied. According to Hund's rule, the sixth electron enters the second of those p orbitals and with the same spin as the fifth electron.

You can watch video lectures on this topic from Khan Academy:

- Introduction to Orbitals at http://www.youtube.com/watch?v=yBrp8uvNAhI (13:38)



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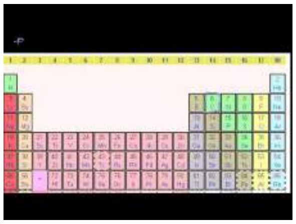
Click image to the left for more content.

- More on Orbitals and Electron Configuration at http://www.youtube.com/watch?v=FmQoSenbtnU (14:31)



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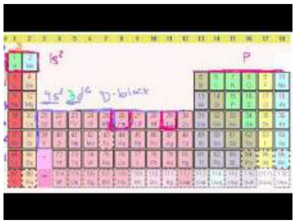
- Electron Configurations at http://www.youtube.com/watch?v=RJlEH5Jz80w (10:04)



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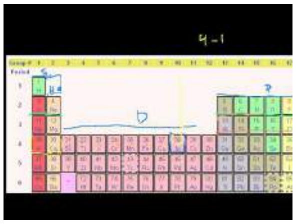
- Electron Configurations 2 at http://www.youtube.com/watch?v=YURReI6OJsg (10:18)



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Second Period Elements

Periods refer to the horizontal rows of the periodic table. Looking at a periodic table you will see that the first period contains only the elements hydrogen and helium. This is because the first principal energy level consists of only the s sublevel and so only two electrons are required in order to fill the entire principal energy level. Each time a new principal energy level begins, as with the third element lithium, a new period is started on the periodic table. As one moves across the second period, electrons are successively added. With beryllium ($Z = 4$), the $2s$ sublevel is complete and the $2p$ sublevel begins with boron ($Z = 5$). Since there are three $2p$ orbitals and each orbital holds two electrons, the $2p$ sublevel is filled after six elements. **Table** 5.3 shows the electron configurations of the elements in the second period.

TABLE 5.3: **Electron Configurations of Second-Period Elements**

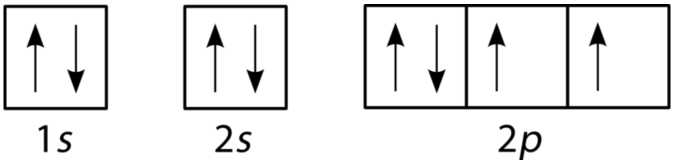
Element Name	Symbol	Atomic Number	Electron Configuration
Lithium	Li	3	$1s^2 2s^1$
Beryllium	Be	4	$1s^2 2s^2$
Boron	B	5	$1s^2 2s^2 2p^1$
Carbon	C	6	$1s^2 2s^2 2p^2$
Nitrogen	N	7	$1s^2 2s^2 2p^3$
Oxygen	O	8	$1s^2 2s^2 2p^4$
Fluorine	F	9	$1s^2 2s^2 2p^5$
Neon	Ne	10	$1s^2 2s^2 2p^6$

Upon reaching the element neon, the last electron to add to the atom is the p^6 electron and so the sublevel and the principal energy level are now filled. In the study of chemical reactivity, we will find that the electrons in the outermost principal energy level are very important and so they are given a special name. **Valence electrons** *are the electrons in the highest occupied principal energy level of an atom.* In the second period elements listed above, the two electrons in the $1s$ sublevel are called inner-shell electrons and are not involved directly in the element's reactivity or in the formation of compounds. Lithium has a single electron in the second principle energy level and so we say that lithium has one valence electron. Beryllium has two valence electrons . How many valence electrons does boron have? You must recognize that the second principal energy level consists of both the $2s$ and the $2p$ sublevels and so the answer is three. In fact, the number of valence electrons goes up by one for each step across a period until the last element is reached. Neon, with its configuration ending in s^2p^6, has eight valence electrons.

The magnetic properties of elements are related to the number of electrons in the atom that are unpaired. For example, hydrogen has only a single electron, so it is necessarily unpaired. Helium has no unpaired electrons because its second electron is matched up, with opposite spin, in the same orbital as the first electron. Proceeding across the second period, we come up with the following numbers of unpaired electrons:

Li = 1, Be = 0, B = 1, C = 2, N = 3, O = 2, F = 1, Ne = 0

In order to correctly identify the number of unpaired electrons, you may find it necessary to construct the orbital filling diagram. Hund's rule must be followed and that will affect the number of unpaired electrons. Oxygen's orbital filling diagram serves as an example:



The four electrons in the $2p$ sublevel are ordered such that two are paired up in the first orbital, while the single electrons in the second and third orbitals are unpaired.

Third Period Elements

Sodium, element number eleven, is the first element in the third period of the periodic table. Its electron configuration is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons of the sodium atom are the inner-shell electrons and the configuration of just those ten electrons is exactly the same as the configuration of the element neon ($Z = 10$). This provides the basis for a shorthand notation for electron configurations called the noble gas configuration. The elements that are found in the last column of the periodic table are an important group of elements that are called the noble gases. They are helium, neon, argon, krypton, xenon, and radon. A **noble gas configuration** *of an atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons.* So for sodium, we make the substitution of [Ne] for the $1s{=}^2 2s^2 2p^6$ part of the configuration. Sodium's noble gas configuration becomes [Ne]$3s^1$. **Table** 5.4 shows the noble gas configurations of the third period elements.

TABLE 5.4: **Electron Configurations of Third-Period Elements**

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Sodium	Na	11	[Ne]$3s^1$
Magnesium	Mg	12	[Ne]$3s^2$
Aluminum	Al	13	[Ne]$3s^2 3p^1$
Silicon	Si	14	[Ne]$3s^2 3p^2$
Phosphorus	P	15	[Ne]$3s^2 3p^3$
Sulfur	S	16	[Ne]$3s^2 3p^4$
Chlorine	Cl	17	[Ne]$3s^2 3p^5$
Argon	Ar	18	[Ne]$3s^2 3p^6$

Again, the number of valence electrons increases from one to eight across the third period.

Fourth and Fifth Period Elements

The element potassium begins the fourth period. The last electron in the potassium atom goes into the $4s$ sublevel, which fills before the $3d$ sublevel. From this point onward, it is important to consult the diagram in **Figure** 5.22 in order to follow the Aufbau process correctly. The fourth period elements are shown in **Table** 5.5.

TABLE 5.5: Electron Configurations of Fourth-Period Elements

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Potassium	K	19	$[\mathrm{Ar}]4s^1$
Calcium	Ca	20	$[\mathrm{Ar}]4s^2$
Scandium	Sc	21	$[\mathrm{Ar}]3d^1 4s^2$
Titanium	Ti	22	$[\mathrm{Ar}]3d^2 4s^2$
Vanadium	V	23	$[\mathrm{Ar}]3d^3 4s^2$
Chromium	Cr	24	$[\mathrm{Ar}]3d^5 4s^1$
Manganese	Mn	25	$[\mathrm{Ar}]3d^5 4s^2$
Iron	Fe	26	$[\mathrm{Ar}]3d^6 4s^2$
Cobalt	Co	27	$[\mathrm{Ar}]3d^7 4s^2$
Nickel	Ni	28	$[\mathrm{Ar}]3d^8 4s^2$
Copper	Cu	29	$[\mathrm{Ar}]3d^{10} 4s^1$
Zinc	Zn	30	$[\mathrm{Ar}]3d^{10} 4s^2$
Gallium	Ga	31	$[\mathrm{Ar}]3d^{10} 4s^2 4p^1$
Germanium	Ge	32	$[\mathrm{Ar}]3d^{10} 4s^2 4p^2$
Arsenic	As	33	$[\mathrm{Ar}]3d^{10} 4s^2 4p^3$
Selenium	Se	34	$[\mathrm{Ar}]3d^{10} 4s^2 4p^4$
Bromine	Br	35	$[\mathrm{Ar}]3d^{10} 4s^2 4p^5$
Krypton	Kr	36	$[\mathrm{Ar}]3d^{10} 4s^2 4p^6$

Beginning with scandium ($Z = 21$), the $3d$ sublevel begins to fill. Note that it is customary for the lower principal energy level (the 3^{rd}) to be written first in the electron configuration even though it fills after the outer 4^{th} principal energy level. Titanium and vanadium follow scandium. We would expect that the next element, chromium, would have an outer configuration of $3d^4 4s^2$. However, in this case, one of the $4s$ electrons is shifted to the last of the empty $3d$ orbitals, resulting in an outer configuration of $3d^5 4s^1$. This is of slightly lower energy and thus is more stable than the arrangement we would expect according to the Aufbau principle. The reason is because having six unpaired electrons instead of four minimizes electron-electron repulsions.

Proceeding onward, manganese (Mn) pairs up the single electron in the $4s$ orbital. Iron (Fe), cobalt (Co), and nickel (Ni) follow, with electrons now pairing up in the $3d$ orbitals. Copper (Cu) is another element with an unexpected configuration. This time, the second $4s$ electron is shifted into the last available spot in the $3d$ sublevel, completely filling it. This is the lowest energy configuration for Cu. Exceptional electron configurations like those of Cr and Cu occur frequently among the heavier elements and are not always easy to explain. It is important to be aware of them, but not necessary to memorize every exception. From gallium, Ga, through the noble gas krypton, Kr, electrons fill the $4p$ sublevel.

Recall that valence electrons are only those electrons in the outermost principal energy level. For elements 21-30, electrons are being added to the $3d$ sublevel which is not the outermost energy level since the $4s$ electrons entered those atoms first. Therefore, the atoms of elements 21-30 have two valence electrons, with the exception of chromium and copper which have only one valence electron. Because the d and f sublevels always fill behind s

sublevels of a higher principal energy level, they can never be valence electrons. The maximum number of valence electrons possible is eight, from a full set of *s* and *p* orbitals.

The fifth period consists of the elements rubidium (Rb), through xenon (Xe). The pattern of sublevel filling is the same as in the fourth period: 5*s* followed by 4*d* followed by 5*p*. Several more exceptional configurations occur as seen in **Table** 5.6.

TABLE 5.6: **Electron Configurations of Fifth-Period Elements**

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Rubidium	Rb	37	$[Kr]5s^1$
Strontium	Sr	38	$[Kr]5s^2$
Yttrium	Y	39	$[Kr]4d^1 5s^2$
Zirconium	Zr	40	$[Kr]4d^2 5s^2$
Niobium	Nb	41	$[Kr]4d^4 5s^1$
Molybdenum	Mo	42	$[Kr]4d^4 5s^1$
Technetium	Tc	43	$[Kr]4d^5 5s^2$
Ruthenium	Ru	44	$[Kr]4d^7 5s^1$
Rhodium	Rh	45	$[Kr]4d^8 5s^1$
Palladium	Pd	46	$[Kr]4d^{10}$
Silver	Ag	47	$[Kr]4d^{10} 5s^1$
Cadmium	Cd	48	$[Kr]4d^{10} 5s^2$
Indium	In	49	$[Kr]4d^{10} 5s^2 5p^1$
Tin	Sn	50	$[Kr]4d^{10} 5s^2 5p^2$
Antimony	Sb	51	$[Kr]4d^{10} 5s^2 5p^3$
Tellurium	Te	52	$[Kr]4d^{10} 5s^2 5p^4$
Iodine	I	53	$[Kr]4d^{10} 5s^2 5p^5$
Xenon	Xe	54	$[Kr]4d^{10} 5s^2 5p^6$

Sixth and Seventh Period Elements

The sixth period contains 32 elements and begins with cesium (Cs) filling the 6*s* sublevel. The 6*s* sublevel is followed by the 4*f*, then the 5*d* and finally the 6*p*. See **Table** 5.7 for the electron configurations of sixth period elements, noting again the large number of exceptions.

TABLE 5.7: **Electron Configurations of Sixth-Period Elements**

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Cesium	Cs	55	$[Xe]6s^1$
Barium	Ba	56	$[Xe]6s^2$
Lanthanum	La	57	$[Xe]5d^1 6s^2$
Cerium	Ce	58	$[Xe]4f^1 5d^1 6s^2$
Praseodymium	Pr	59	$[Xe]4f^3 6s^2$
Neodymium	Nd	60	$[Xe]4f^4 6s^2$
Promethium	Pm	61	$[Xe]4f^5 6s^2$
Samarium	Sm	62	$[Xe]4f^6 6s^2$
Europium	Eu	63	$[Xe]4f^7 6s^2$
Gadolinium	Gd	64	$[Xe]4f^7 5d^1 6s^2$
Terbium	Tb	65	$[Xe]4f^9 6s^2$

TABLE 5.7: (continued)

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Dysprosium	Dy	66	[Xe]$4f^{10}6s^2$
Holmium	Ho	67	[Xe]$4f^{11}6s^2$
Erbium	Er	68	[Xe]$4f^{12}6s^2$
Thulium	Tm	69	[Xe]$4f^{13}6s^2$
Ytterbium	Yb	70	[Xe]$4f^{14}6s^2$
Lutetium	Lu	71	[Xe]$4f^{14}5d^16s^2$
Hafnium	Hf	72	[Xe]$4f^{14}5d^26s^2$
Tantalum	Ta	73	[Xe]$4f^{14}5d^36s^2$
Tungsten	W	74	[Xe]$4f^{14}5d^46s^2$
Rhenium	Re	75	[Xe]$4f^{14}5d^56s^2$
Osmium	Os	76	[Xe]$4f^{14}5d^66s^2$
Iridium	Ir	77	[Xe]$4f^{14}5d^76s^2$
Platinum	Pt	78	[Xe]$4f^{14}5d^96s^1$
Gold	Au	79	[Xe]$4f^{14}5d^{10}6s^1$
Mercury	Hg	80	[Xe]$4f^{14}5d^{10}6s^2$
Thallium	Tl	81	[Xe]$4f^{14}5d^{10}6s^26p^1$
Lead	Pb	82	[Xe]$4f^{14}5d^{10}6s^26p^2$
Bismuth	Bi	83	[Xe]$4f^{14}5d^{10}6s^26p^3$
Polonium	Po	84	[Xe]$4f^{14}5d^{10}6s^26p^4$
Astatine	At	85	[Xe]$4f^{14}5d^{10}6s^26p^5$
Radon	Rn	86	[Xe]$4f^{14}5d^{10}6s^26p^6$

The seventh period starts with francium (Fr) and the 7*s* sublevel. This is followed by the 5*f*, and 6*d*. The seventh period is incomplete and consists largely of artificial and very unstable elements. Electron configurations for the known seventh period elements are listed in **Table** 5.8.

TABLE 5.8: **Electron Configurations of Seventh-Period Elements**

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Francium	Fr	87	[Rn]$7s^1$
Radium	Ra	88	[Rn]$7s^2$
Actinium	Ac	89	[Rn]$6d^17s^2$
Thorium	Th	90	[Rn]$6d^27s^2$
Protactinium	Pa	91	[Rn]$5f^26d^17s^2$
Uranium	U	92	[Rn]$5f^36d^17s^2$
Neptunium	Np	93	[Rn]$5f^46d^17s^2$
Plutonium	Pu	94	[Rn]$5f^67s^2$
Americium	Am	95	[Rn]$5f^77s^2$
Curium	Cm	96	[Rn]$5f^76d^17s^2$
Berkelium	Bk	97	[Rn]$5f^97s^2$
Californium	Cf	98	[Rn]$5f^{10}7s^2$
Einsteinium	Es	99	[Rn]$5f^{11}7s^2$
Fermium	Fm	100	[Rn]$5f^{12}7s^2$
Mendelevium	Md	101	[Rn]$5f^{13}7s^2$
Nobelium	No	102	[Rn]$5f^{14}7s^2$
Lawrencium	Lr	103	[Rn]$5f^{14}6d^17s^2$
Rutherfordium	Rf	104	[Rn]$5f^{14}6d^27s^2$

TABLE 5.8: (continued)

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Dubnium	Db	105	$[\mathrm{Rn}]5f^{14}6d^{3}6s^{2}$
Seaborgium	Sg	106	$[\mathrm{Rn}]5f^{14}6d^{4}7s^{2}$
Bohrium	Bh	107	$[\mathrm{Rn}]5f^{14}6d^{5}7s^{2}$
Hassium	Hs	108	$[\mathrm{Rn}]5f^{14}6d^{6}7s^{2}$
Meitnerium	Mt	109	$[\mathrm{Rn}]5f^{14}6d^{7}7s^{2}$

Lesson Summary

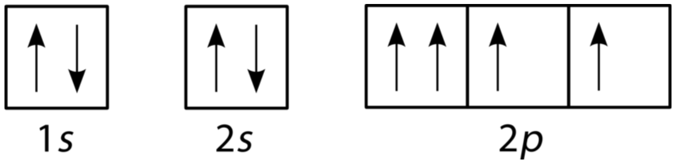
- Electrons occupy atomic orbitals in the ground state of atoms according to the Aufbau principle, the Pauli exclusion principle, and Hund's rule.
- Orbital filling diagrams are drawn to show how electrons fill up energy levels and orbitals from low energy to high.
- The electron configuration is unique for each element.
- Valence electrons are the electrons in the outermost principal energy level and there is a maximum of eight.
- Unpaired electrons are important to an atom's magnetic properties.
- Some electron configurations such as those of chromium and copper, do not strictly follow the Aufbau principle.

Lesson Review Questions

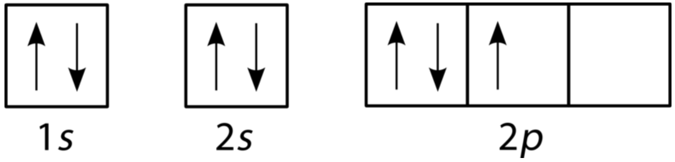
Reviewing Concepts

1. What do the superscripts in an electron configuration represent?
    a. What is the atomic number of an element with the full electron configuration of $1s^22s^22p^23s^23p^63d^{10}4s^24p^5$?
2. Arrange these sublevels in order of increasing energy: $3p$, $5s$, $4f$, $2s$, $3d$.
3. Which rule is violated with the electron configuration $1s^22s^22p^63s^23p^63d^2$? Explain.
4. Which rule is violated by the orbital filling diagrams below? Explain.

a.



b.



Problems

5. Using only the atomic number, construct orbital filling diagrams for the following elements. Note: none are exceptions to the Aufbau principle.
    a. boron ($Z = 5$)

   b. sulfur ($Z = 16$)
   c. nickel ($Z = 28$)
   d. rubidium ($Z = 37$)

6. Using only the atomic number, write full electron configurations for the following elements.
   a. fluorine ($Z = 9$)
   b. calcium ($Z = 20$)
   c. zirconium ($Z = 40$)
   d. europium ($Z = 63$)
7. Write the noble gas configurations for the same four elements from question #6.
8. How many electrons are in the second principal energy level of each element below?
   a. sodium
   b. nitrogen
   c. beryllium
9. How many completely filled principal energy levels in atoms of each element?
   a. argon
   b. ruthenium
   c. barium
10. How many principal energy levels are occupied in atoms of the elements listed in question #9?
11. How many sublevels are occupied in an atom of aluminum? How many are completely filled?
12. Which elements have completely filled outermost *s* and *p* sublevels?
13. Fill in the following table (**Table** 5.9). Use a periodic table to find the atomic numbers. All follow the conventional Aufbau order of filling sublevels.

TABLE 5.9: **Table for Problem 13**

	O	K	Fe	Kr	Cd	I	Pu
Total number of occupied principal energy levels							
Number of completely filled principal energy levels							
Total number of occupied sublevels							
Number of completely filled sublevels							
Number of unpaired electrons							

TABLE 5.9: (continued)

	O	K	Fe	Kr	Cd	I	Pu
Number of valence electrons							

NC Standards addressed in this chapter

Chm.1.1.2 Analyze an atom in terms of the location of electrons.

• Analyze diagrams related to the Bohr model of the hydrogen atom in terms of allowed, discrete energy levels in the emission spectrum. • Describe the electron cloud of the atom in terms of a probability model.

• Relate the electron configurations of atoms to the Bohr and electron cloud models.

Chm.1.1.3 Explain the emission of electromagnetic radiation in spectral form in terms of the Bohr model.

• Understand that energy exists in discrete units called quanta. • Describe the concepts of excited and ground state of electrons in the atom:

1. When an electron gains an amount of energy equivalent to the energy difference, it moves from its ground state to a higher energy level.

2. When the electron moves to a lower energy level, it releases an amount of energy equal to the energy difference in these levels as electromagnetic radiation (emissions spectrum).

• Articulate that this electromagnetic radiation is given off as photons.

• Understand the inverse relationship between wavelength and frequency, and the direct relationship between energy and frequency.

• Use the "Bohr Model for Hydrogen Atom" and "Electromagnetic Spectrum" diagrams from the Reference Tables to relate color, frequency, and wavelength of the light emitted to the energy of the photon.

• Explain that Niels Bohr produced a model of the hydrogen atom based on experimental observations. This model indicated that:

1. an electron circles the nucleus only in fixed energy ranges called orbits;

2. an electron can neither gain or lose energy inside this orbit, but could move up or down to another orbit;

3. that the lowest energy orbit is closest to the nucleus.

• Describe the wave/particle duality of electrons.

Chm.1.3 Understand the physical and chemical properties of atoms based on their position on the Periodic Table. Electron configurations/valence electrons/ionization energy/electronegativity

• Write electron configurations, including noble gas abbreviations (no exceptions to the general rules). Included here are extended arrangements showing electrons in orbitals.

• Identify s, p, d, and f blocks on Periodic Table.

• Identify an element based on its electron configuration. (Students should be able to identify elements which follow the general rules, not necessarily those which are exceptions.)

• Determine the number of valence electrons from electron configurations.

• Predict the number of electrons lost or gained and the oxidation number based on the electron configuration of an atom. • Define ionization energy and know group and period general trends for ionization energy. Explain the reasoning behind the trend.

• Apply trends to arrange elements in order of increasing or decreasing ionization energy.

- Define electronegativity and know group and period general trends for electronegativity. Explain the reasoning behind the trend.
- Apply trends to arrange elements in order of increasing or decreasing electronegativity.

5.5 References

1. CK-12 Foundation - Zachary Wilson. . CC-BY-NC-SA 3.0
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CHAPTER 6 The Periodic Table

Chapter Outline

Getting and staying organized: it's a challenge. Everywhere you look, humans are constantly trying to organize things in the most efficient way possible. Here in this hardware store, the myriad items for sale need to be organized so that the shoppers can find what they need as quickly and easily as possible. This organization works on multiple levels such as by aisle number or by manufacturer or by price. Likewise, the periodic table is a marvelously organized tool that chemists use to keep track of chemical elements. The placement of each element on the table easily tells the chemist valuable information about that element's physical and chemical properties. In this chapter you will learn how the periodic table was developed and how it can be used to make learning chemistry easier.

Opening image courtesy of HazelBlk. CC–BY 3.0.

6.1 History of the Periodic Table

Lesson Objectives

- Explain early attempts at organizing chemical elements
- Understand how Mendeleev organized his periodic table
- Know the improvements that Moseley made on Mendeleev's table
- Know the periodic law
- Describe components of the modern periodic table: periods, groups, metals, nonmetals, and metalloids

Lesson Vocabulary

- group
- metal
- metalloid
- nonmetal
- period
- periodic law
- periodic table

Early Attempts to Organize Elements

By the year 1700, only a handful of elements had been identified and isolated. Several of these, such as copper and lead, had been known since ancient times. As scientific methods improved, the rate of discovery dramatically increased (**Figure** 6.1).

With the ever-increasing number of elements, chemists recognized that there may be some kind of systematic way to organize the elements. The question was: how?

A logical way to begin to group elements together was by their chemical properties. In other words, putting elements in separate groups based on how they reacted with other elements. In 1829, a German chemist, Johann Dobereiner (1780-1849), placed various groups of three elements into groups called triads. One such triad was lithium, sodium, and potassium. Triads were based on both physical as well as chemical properties. Dobereiner found that the atomic masses of these three elements, as well as other triads, formed a pattern. When the atomic masses of lithium and potassium were averaged together $(6.94 + 39.10)/2 = 23.02$, it was approximately equal to the atomic mass of sodium (22.99). These three elements also displayed similar chemical reactions, such as vigorously reacting with the members of another triad: chlorine, bromine, and iodine (**Figure** 6.2).

While Dobereiner's system would pave the way for future ideas, a limitation of the triad system was that not all of the known elements could be classified in this way.

English chemist John Newlands (1838-1898) ordered the elements in increasing order of atomic mass and noticed that every eighth element exhibited similar properties. He called this relationship the *Law of Octaves*. Unfortunately,

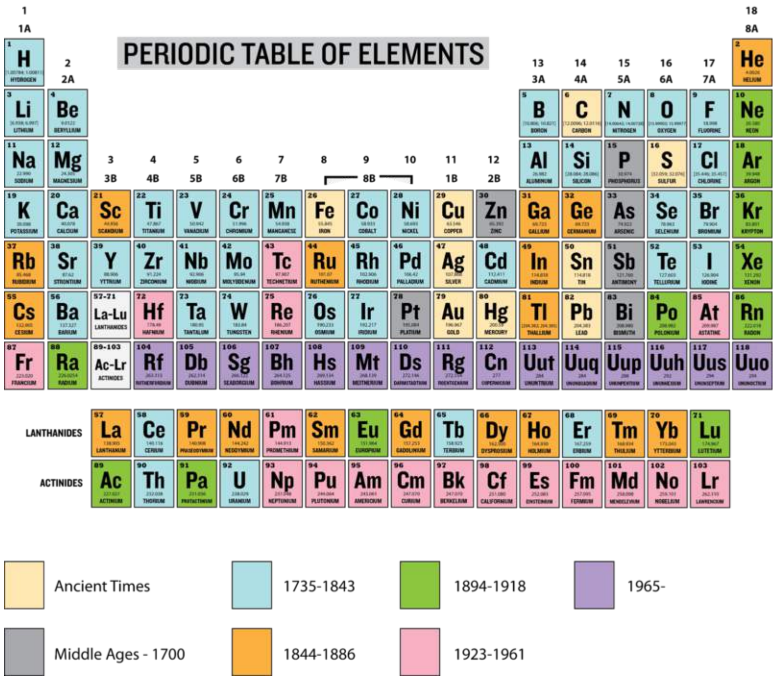


FIGURE 6.1
Dates of discovery of the chemical elements.

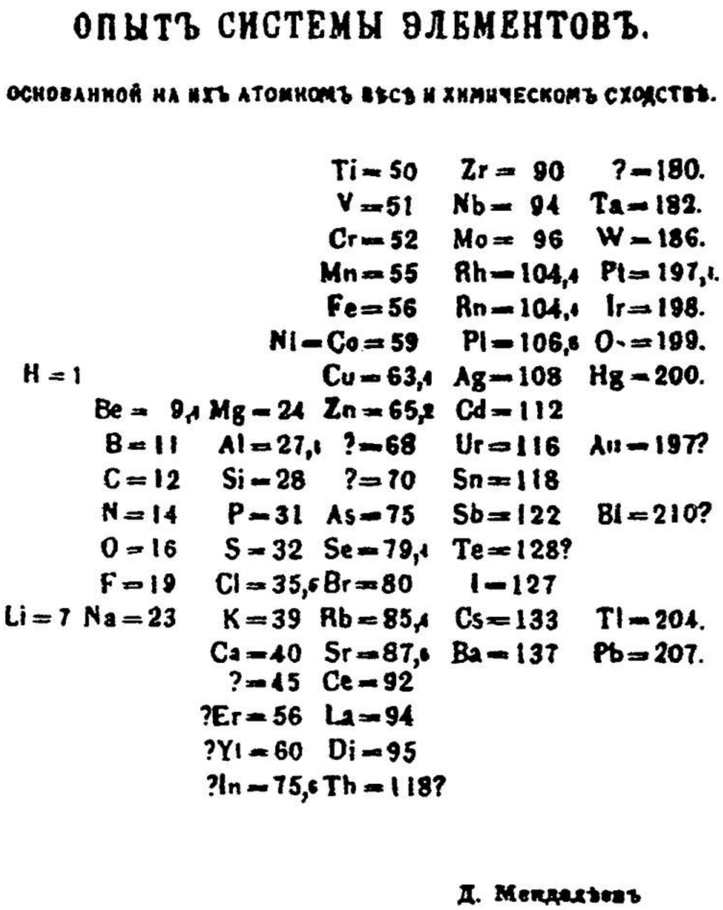
FIGURE 6.2
The elements chlorine, bromine, and iodine: Chlorine is a greenish-yellow gas. Bromine is a dark orange liquid. Iodine is a shiny blue-black solid. Though different in appearance, they have very similar chemical properties.

there were some elements that were missing and the law did not seem to hold for elements that were heavier than calcium. Newlands's work was largely ignored and even ridiculed by the scientific community in his day. It was not until years later that another more extensive periodic table effort would gain much greater acceptance and the pioneering work of John Newlands would be appreciated.

Mendeleev's Periodic Table

In 1869, Russian chemist and teacher Dmitri Mendeleev (1836-1907) published a periodic table of the elements. The following year, German chemist Lothar Meyer independently published a very similar table. Mendeleev is generally given more credit than Meyer because his table was published first and because of several key insights that he made regarding the table.

Mendeleev was writing a chemistry textbook for his students and wanted to organize all of the known elements at that time according to their chemical properties. He famously organized the information for each element on to separate note cards that were then easy to rearrange as needed. He discovered that when he placed them in order of increasing atomic mass, certain similarities in chemical behavior repeated at regular intervals. This type of a repeating pattern is called *periodic*. A pendulum that swings back and forth in a given time interval is periodic, as is the movement of the moon around the Earth. **Figure** 6.3 shows an early version of Mendeleev's table.



ОПЫТЪ СИСТЕМЫ ЭЛЕМЕНТОВЪ.

ОСНОВАННОЙ НА ИХЪ АТОМНОМЪ ВѢСѢ И ХИМИЧЕСКОМЪ СХОДСТВѢ.

			Ti=50	Zr=90	?=180.
			V=51	Nb=94	Ta=182.
			Cr=52	Mo=96	W=186.
			Mn=55	Rh=104,4	Pt=197,4.
			Fe=56	Rn=104,4	Ir=198.
			Ni=Co=59	Pl=106,6	Os=199.
H=1			Cu=63,4	Ag=108	Hg=200.
	Be=9,4	Mg=24	Zn=65,2	Cd=112	
	B=11	Al=27,4	?=68	Ur=116	Au=197?
	C=12	Si=28	?=70	Sn=118	
	N=14	P=31	As=75	Sb=122	Bi=210?
	O=16	S=32	Se=79,4	Te=128?	
	F=19	Cl=35,5	Br=80	I=127	
Li=7	Na=23	K=39	Rb=85,4	Cs=133	Tl=204.
		Ca=40	Sr=87,6	Ba=137	Pb=207.
		?=45	Ce=92		
		?Er=56	La=94		
		?Yt=60	Di=95		
		?In=75,6	Th=118?		

Д. Менделѣевъ

FIGURE 6.3

Mendeleev's first published periodic table shows elements arranged in vertical columns according to increasing atomic mass. The atomic mass is the number which follows each symbol. Elements with question marks were unknown at the time, but were discovered at a later date.

In this table, atomic mass increases from top to bottom of vertical columns, with successive columns going left to right. As a result, elements that are in the same horizontal row are groups of elements that were known to exhibit similar chemical properties. One of Mendeleev's insights is illustrated by the elements tellurium (Te) and iodine (I). Notice that tellurium is listed before iodine even though its atomic mass is higher. Mendeleev reversed the order because he knew that the properties of iodine were much more similar to those of fluorine (F), chlorine (Cl), and bromine (Br) than they were to oxygen (O), sulfur (S), and selenium (Se). He simply assumed that there was an error in the determination of one or both of the atomic masses. As we will see shortly, this turned out not to be the case, but Mendeleev was indeed correct to group these two elements as he did.

Notice that there are several places in the table that have no chemical symbol, but are instead labeled with a question mark. Between zinc (Zn) and arsenic (As) are two such missing elements. Mendeleev believed that elements with atomic masses of 68 and 70 would eventually be discovered and that they would fit chemically into each of those spaces. Listed in **Table** 6.1 are other properties that Mendeleev predicted for the first of these two missing elements, which he called "eka-aluminum," compared with the element gallium.

TABLE 6.1: Mendeleev's Predictions for Eka-Aluminum

	Eka-Aluminum (Ea)	Gallium (Ga)
Atomic mass	68 amu	69.9 amu
Melting point	Low	30.15°C
Density	5.9 g/cm^3	5.94 g/cm^3
Formula of oxide	Ea_2O_3	Ga_2O_3

The element gallium was discovered four years after the publication of Mendeleev's table, and its properties matched up remarkably well with eka-aluminum, fitting into the table exactly where he had predicted. This was also the case with the element that followed gallium, which was named eventually named germanium.

Mendeleev's periodic table gained wide acceptance with the scientific community and earned him credit as the discoverer of the periodic law. Element number 101, synthesized in 1955, is named mendelevium after the founder of the periodic table. It would, however, be several years after Mendeleev died before the several discrepancies with the atomic masses could be explained and before the reasons behind the repetition of chemical properties could be fully explained.

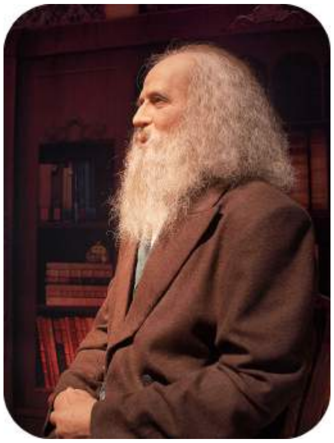


FIGURE 6.4

A wax sculpture of Mendeleev.

Interactive periodic table websites:

- Periodic History: http://freezeray.com/flashFiles/periodicHistory.htm
- Discovery Dates: http://freezeray.com/flashFiles/discoveryDates.htm

The Periodic Law

Recall that Rutherford's experiments which resulted in the discovery of the nucleus occurred in 1911, long after Mendeleev's periodic table was developed. Just two years later, in 1913, English physicist Henry Moseley (1887-1915) examined x-ray spectra of a number of chemical elements. His results led to the definition of atomic number as the number of protons contained in the nucleus of each atom. He then realized that the elements of the periodic table should be arranged in order of increasing atomic number rather than increasing atomic mass.

When ordered by atomic number, the discrepancies within Mendeleev's table disappeared. Tellurium has an atomic number of 52, while iodine has an atomic number of 53. So even though tellurium does indeed have a greater atomic mass than iodine, it is properly placed before iodine in the periodic table. Mendeleev and Moseley are credited with being most responsible for the modern **periodic law**: *When elements are arranged in order of increasing atomic* *number, there is a periodic repetition of their chemical and physical properties.* The result is the periodic table as we know it today (**Figure** 6.5). Each new horizontal row of the periodic table corresponds to the beginning of a new period because a new principal energy level is being filled with electrons. Elements with similar chemical properties appear at regular intervals, within the vertical columns called groups.

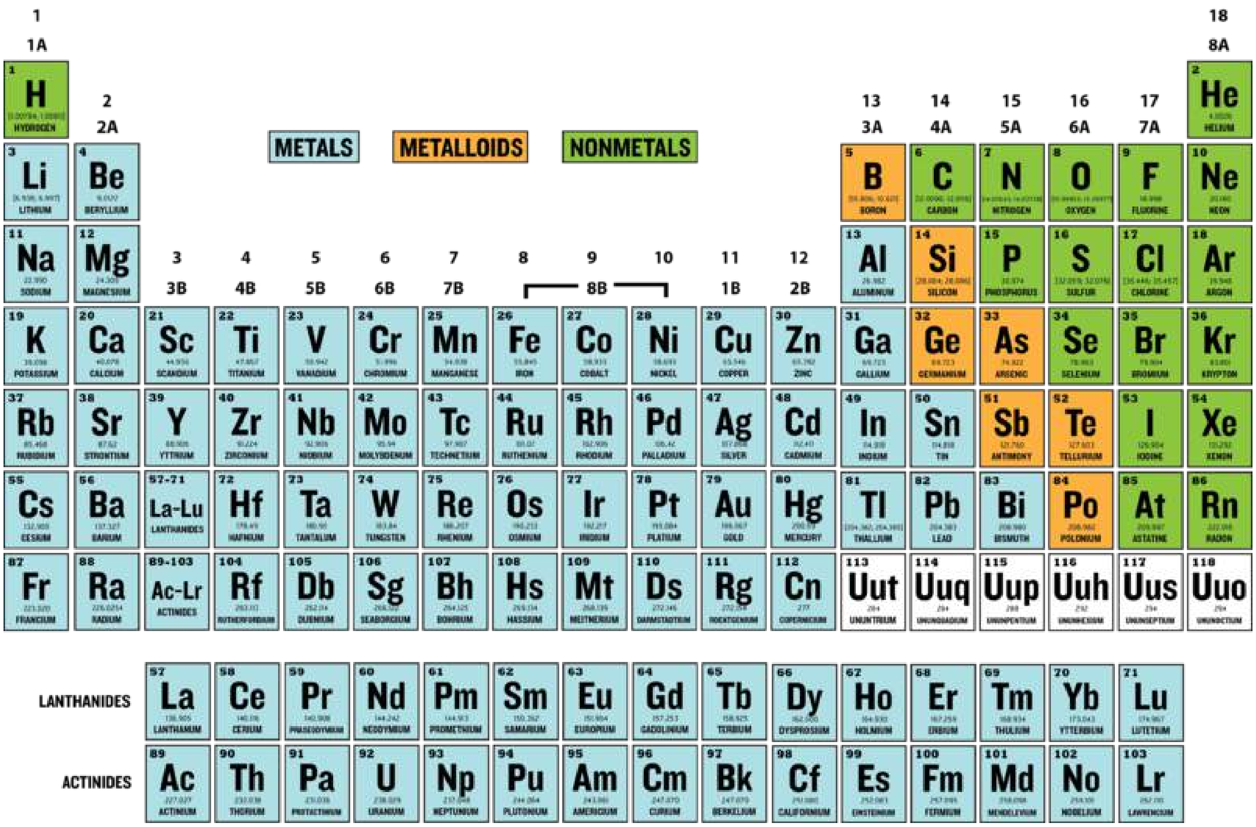


FIGURE 6.5

The periodic table of the elements.

The Modern Periodic Table

The periodic table has undergone extensive changes in the time since it was originally developed by Mendeleev and Moseley. Many new elements have been discovered, while others have been artificially synthesized. Each fits properly into a group of elements with similar properties. The **periodic table** *is an arrangement of the elements in order of their atomic numbers so that elements with similar properties appear in the same vertical column or group.*

Figure 6.5 shows the most commonly used form of the periodic table. Each square shows the chemical symbol of the element along with its name. Notice that several of the symbols seem to be unrelated to the name of the element: Fe for iron, Pb for lead, etc. Most of these are the elements that have been known since ancient times and have symbols based on their Latin names. The atomic number of each element is written above the symbol. Each square on this version of the periodic table also shows the average atomic mass of the element.

A **period** *is a horizontal row of the periodic table.* There are seven periods in the periodic table, with each one beginning at the far left. A new period begins when a new principal energy level begins filling with electrons. Period 1 has only two elements (hydrogen and helium), while periods 2 and 3 have 8 elements. Periods 4 and 5 have 18 elements. Periods 6 and 7 have 32 elements because the two bottom rows that are separated from the rest of the table belong to those periods. They are pulled out in order to make the table itself fit more easily onto a single page.

A **group** *is a vertical column of the periodic table*. There are a total of 18 groups. There are two different numbering systems that are commonly used to designate groups and you should be familiar with both. The traditional system used in the United States involves the use of the letters A and B. The first two groups are 1A and 2A, while the last six groups are 3A through 8A. The middle groups use B in their titles. Unfortunately, there was a slightly different system in place in Europe. To eliminate confusion the International Union of Pure and Applied Chemistry (IUPAC) decided that the official system for numbering groups would be a simple 1 through 18 from left to right. Many periodic tables show both systems simultaneously

Metals, Nonmetals, and Metalloids

Elements can be classified in a number of different ways. Classifying by period and/or group is important because it is based on their electron configuration. Another way is to classify elements based on physical properties. Three broad classes of elements are metals, nonmetals, and metalloids.

A **metal** *is an element that is a good conductor of heat and electricity*. Metals are also malleable, which means that they can be hammered into very thin sheets without breaking. They are ductile, which means that they can be drawn into wires. When a fresh surface of any metal is exposed, it will be very shiny because it reflects light well. This is called luster. All metals are solid at room temperature with the exception of mercury (Hg), which is a liquid. Melting points of metals display a very wide variance. The melting point of mercury is −39°C, while the highest melting metal is tungsten (W), with a melting point of 3422°C. On the periodic table of **Figure** 6.5, the metals are blue and are located to the left of the bold stair-step line. About 80 percent of the elements are metals (see examples in **Figure** 6.6).

FIGURE 6.6

The elements mercury, gold, and copper show the properties of metals. Mercury (*left*) is the only liquid metal, but has high luster. Gold (*middle*) is malleable and can be formed into very thin sheets called gold leaf. Copper is very ductile and a good conductor. Copper (*right*) is used extensively in electrical wiring.

A **nonmetal** *is an element that is generally a poor conductor of heat and electricity*. Most properties of nonmetals are the opposite of metals. There is a wider variation in properties among the nonmetals than among the metals, as seen in **Figure** 6.7. Nonmetals exist in all three states of matter. The majority are gases, such as nitrogen and oxygen. Bromine is a liquid. A few are solids, such as carbon and sulfur. In the solid state, nonmetals are brittle, meaning that they will shatter if struck with a hammer. The solids are not lustrous. Melting points are generally much lower than those of metals. On the periodic table, the nonmetals are the green squares and are located to the right of the stair-step line.

A **metalloid** *is an element that has properties that are intermediate between those of metals and nonmetals*. Silicon is a typical metalloid (**Figure** 6.8). It has luster like a metal, but is brittle like a nonmetal. Silicon is used extensively in computer chips and other electronics because its electrical conductivity is in between that of a metal and a nonmetal.

FIGURE 6.7

Nonmetals have properties that are unlike those of metals. Sulfur (*left*) is brittle and its distinctive yellow color lacks luster. Bromine (*center*) is the only liquid nonmetal and must be carefully handled due to its toxicity. Helium (*right*), a colorless and unreactive gas, is lighter than air and thus is used in blimps.

Metalloids can also be called semimetals. On the periodic table, the elements colored red, which generally border the stair-step line, are considered to be metalloids. Notice that aluminum borders the line, but it is considered to be a metal since all of its properties are like those of metals.

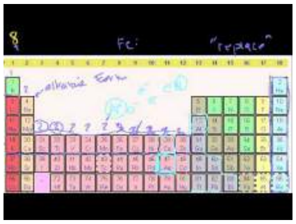
FIGURE 6.8

Silicon is an example of a metalloid, having some properties of a metal and some of a nonmetal.

Explore some interactive periodic tables at:

- Periodic Table: http://chemistry.about.com/library/blperiodictable.htm?nl=1
- It's Elemental (NOVA): http://www.pbs.org/wgbh/nova/physics/periodic-table.html
- It's Elemental! Jefferson Lab: http://education.jlab.org/itselemental/index.html

Watch this Khan Academy video on Groups of the Periodic Table (11:51):



MEDIA

Click image to the left for more content.

Lesson Summary

- Mendeleev arranged chemical elements into a periodic table by arranging them in order of increasing atomic mass.
- According to the periodic law, physical and chemical properties of elements are periodic functions of their atomic numbers.
- In the modern periodic table, elements are arranged into periods and groups. Elements in the same group have similar properties.
- Elements can be classified into the categories of metals, nonmetals, or metalloids.

Lesson Review Questions

Reviewing Concepts

1. Compare Dobereiner's triads to the modern periodic table. What aspects of his idea agreed with the modern table and what aspects did not?
2. Why did Mendeleev leave empty spaces in his periodic table?
3. Explain the significance of the discovery of gallium to Mendeleev's periodic table.
4. What discovery occurred after Mendeleev's periodic table was published that allowed Moseley to improve it?
5. Identify each physical property as more characteristic of a metal or of a nonmetal.
    a. luster
    b. malleability
    c. brittle
    d. lower melting points
    e. good conductor of electric current

Problems

6. Use **Figure** 6.1 to answer the following.
    a. How many elements were known by 1700?
    b. How many elements were discovered between 1735 and 1843?
7. The atomic mass of beryllium (Be) is 9.01 amu. The atomic mass of calcium (Ca), which is in the same group as beryllium, is 40.08 amu. What would you predict to be the atomic mass of magnesium (Mg), which is also in the same group but is between Be and Ca?
8. Use the periodic table to put the following elements into pairs that would be expected to have similar chemical properties: Ba, S, Br, Ca, K, F, Se, Na
9. Identify the elements below as a metal, nonmetal, or metalloid.
    a. phosphorus
    b. boron
    c. cesium
    d. xenon
    e. bismuth
10. Write the name and symbol of the element located in each position on the periodic table.
    a. period 2 and group 16
    b. period 5 and group 9

c. period 4 and group 5A
d. period 6 and group 12

Further Reading / Supplemental Links

Games:

- Element Concentration: http://education.jlab.org/elementconcentration/index.html
- Element Crossword Puzzles: http://education.jlab.org/elementcrossword/index.html
- Element Flash Cards: http://education.jlab.org/elementflashcards/index.html
- Element Hangman: http://education.jlab.org/elementhangman/index.html
- Element Matching Game: http://education.jlab.org/elementmatching/index.html
- Element Word Scramble: http://education.jlab.org/elementwordscramble/index.html

Interactive web sites:

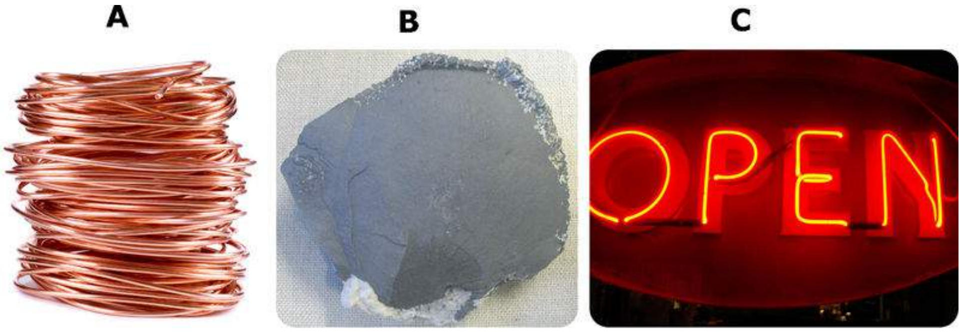
- Did you know you have rocks in your kitchen? Visit Rocks In Your Cabinet at http://www.oercommons.org/courses/kids-in-the-hall-of-planet-earth/view .
- Create an Atomic Mobile: http://www.amnh.org/ology/einstein#features/stufftodo_einstein/atommobile.php?TB_iframe=true&height=550&width=740 .

Tutorial and video:

- Islands of Stability: http://www.pbs.org/wgbh/nova/physics/stability-elements.html

6.2 The Periodic Table and Chemical Families

The Modern Periodic Table and Chemical Families



Look at substances A–C in the photos above. They look very different from one another, but they have something important in common. All three are elements, or pure substances. Can you identify which elements they are? For ideas, listen to the amazing elements song at the URL below. The singer rapidly names all of the known elements while pictures of the elements flash by. Even if the video doesn't help you name the elements pictured above, it will certainly impress you with the need to organize the large number of elements that have been discovered.

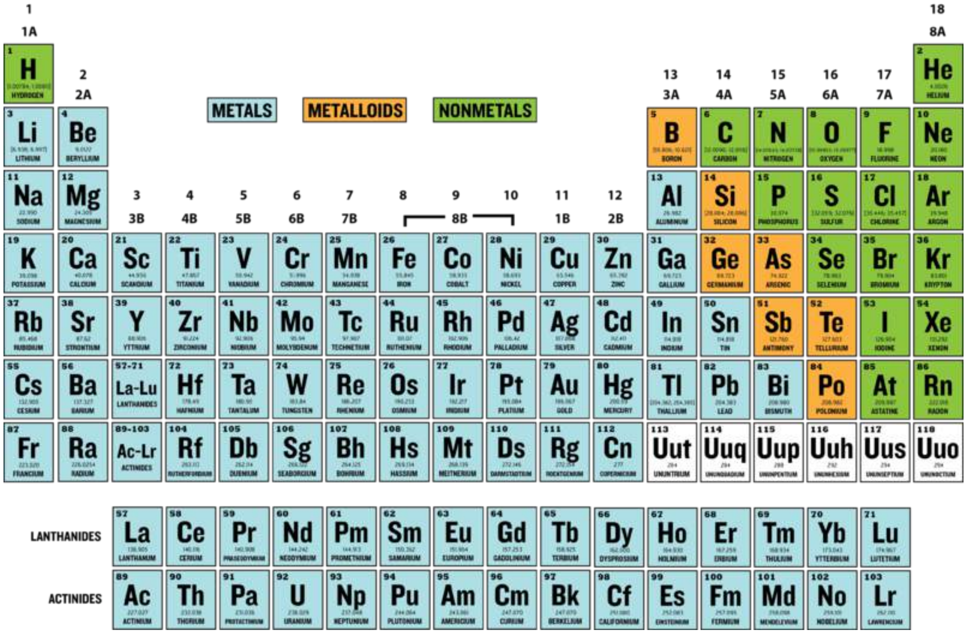
http://www.youtube.com/watch?v=DYW50F42ss8 (1:25)

The First Periodic Table

In the 1860s, a scientist named Dmitri Mendeleev also saw the need to organize the elements. He created a table in which he arranged all of the elements by increasing atomic mass from left to right across each row. When he placed eight elements in each row and then started again in the next row, each column of the table contained elements with similar properties. He called the **columns of elements groups.** Mendeleev's table is called a **periodic table** and the **rows are called periods**. That's because the table keeps repeating from row to row, and **periodic means "repeating."**

The Modern Periodic Table

A periodic table is still used today to organize the elements. You can see a simple version of the modern periodic table in the **Figure** below. The modern table is based on Mendeleev's table, except the modern table **arranges the elements by increasing atomic number** instead of atomic mass. **Atomic number is the number of protons** in an atom, and this number is unique for each element. The modern table has more elements than Mendeleev's table because many elements have been discovered since Mendeleev's time. You can explore an interactive version of the modern periodic table at this URL:http://www.ptable.com/.



METALS METALLOIDS NONMETALS

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57-71 La-Lu LANTHANIDES	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103 Ac-Lr ACTINIDES	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

LANTHANIDES	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
ACTINIDES	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Reading the Table

In the table above, each element is represented by its chemical symbol, which consists of one or two letters. The first letter of the symbol is always written in upper case, and the second letter—if there is one—is always written in lower case. For example, the symbol for copper is Cu. It stands for cuprum, which is the Latin word for copper. The number above each symbol in the table is its unique atomic number. Notice how the atomic numbers increase from left to right and from top to bottom in the table.

Q: Find the symbol for copper in the periodic table above. What is its atomic number? What does this number represent?

A: The atomic number of copper is 29. This number represents the number of protons in each atom of copper. (Copper is the element that makes up the coil of wire in photo A of the opening sequence of photos.)

Periods of the Modern Periodic Table

Rows of the modern periodic table are called **periods**, as they are in Mendeleev's table. From left to right across a period, each element has one more proton than the element before it. Some periods in the modern periodic table are longer than others. For example, period 1 contains only two elements: hydrogen (H) and helium (He). In contrast, periods 6 and 7 are so long that many of their elements are placed below the main part of the table. They are the elements starting with lanthanum (La) in period 6 and actinium (Ac) in period 7. Some elements in period 7 have not yet been named. They are represented by temporary three-letter symbols, such as Uub. The number of each period represents the number of energy levels that have electrons in them for atoms of each element in that period.

Q: Find calcium (Ca) in the periodic table above. How many energy levels have electrons in them for atoms of calcium?

A: Calcium is in period 4, so its atoms have electrons in them for the first four energy levels.

Groups of the Modern Periodic Table

Columns of the modern table are called **groups**, as they are in Mendeleev's table. However, the modern table has many more groups—18 compared with just 8 in Mendeleev's table. **Elements in the same group have similar properties.** For example, all elements in group 18 are colorless, odorless gases, such as neon (Ne). (Neon is the element inside the light in opening photo C.) In contrast, all elements in group 1 are very reactive solids. They react explosively with water, as you can see in the video and **Figure** below.

http://www.youtube.com/watch?v=QSZ-3wScePM (1:09)

The alkali metal sodium (Na) reacting with water.

Classes of Elements

All elements can be classified in one of three classes: **metals, metalloids, or nonmetals.** Elements in each class share certain basic properties. For example, elements in the metals class can conduct electricity, whereas elements in the nonmetals class generally cannot. Elements in the metalloids class fall in between the metals and nonmetals in their properties. An example of a metalloid is arsenic (As). (Arsenic is the element in opening photo B.) In the periodic table above, elements are color coded to show their class. As you move from left to right across each period of the table, the elements change from metals to metalloids to nonmetals.

Q: To which class of elements does copper (Cu) belong: metal, metalloid, or nonmetal? Identify three other elements in this class.

A: In the periodic table above, the cell for copper is colored yellow. This means that copper belongs to the metals class. Other elements in the metals class include iron (Fe), sodium (Na), and gold (Au). It is apparent from the table that the majority of elements are metals.

You can see videos about all of the elements in the modern periodic table at the URL below. Choose an element from each class and watch the video about it.

http://www.periodicvideos.com/index.htm

Summary

The modern periodic table is used to organize all the known elements. Elements are arranged in the table by increasing atomic number.

In the modern periodic table, each element is represented by its chemical symbol. The number above each symbol is its atomic number. Atomic numbers increase from left to right and from top to bottom in the table.

Rows of the periodic table are called periods. From left to right across a period, each element has one more proton than the element before it.

Columns of the periodic table are called groups. Elements in the same group have similar properties.

All elements can be classified in one of three classes: metals, metalloids, or nonmetals. Elements in each class share certain basic properties. From left to right across each period of the periodic table, elements change from metals to metalloids to nonmetals.

Vocabulary

group: Column of the periodic table, which contains elements with similar properties.

period: Row of the periodic table that contains elements ranging from metals on the left to metalloids and to nonmetals on the right.

periodic table: Table of elements arranged by increasing atomic number (modern periodic table) or by increasing atomic mass (Mendeleev's periodic table).

Practice

Practice using the modern periodic table by playing the element math game at the URL below. Be sure to check your answers. For any questions you answer incorrectly, click on the "Tell me more about..." tab to see where you went wrong.

http://education.jlab.org/elementmath/

Review

What is the modern periodic table?

Compare and contrast the periods and groups of the modern periodic table.

In the modern periodic table above, find the element named lead (Pb). How many protons do atoms of lead have? To which class of elements does lead belong?

Which groups of the modern periodic table contain elements that are classified as metalloids?

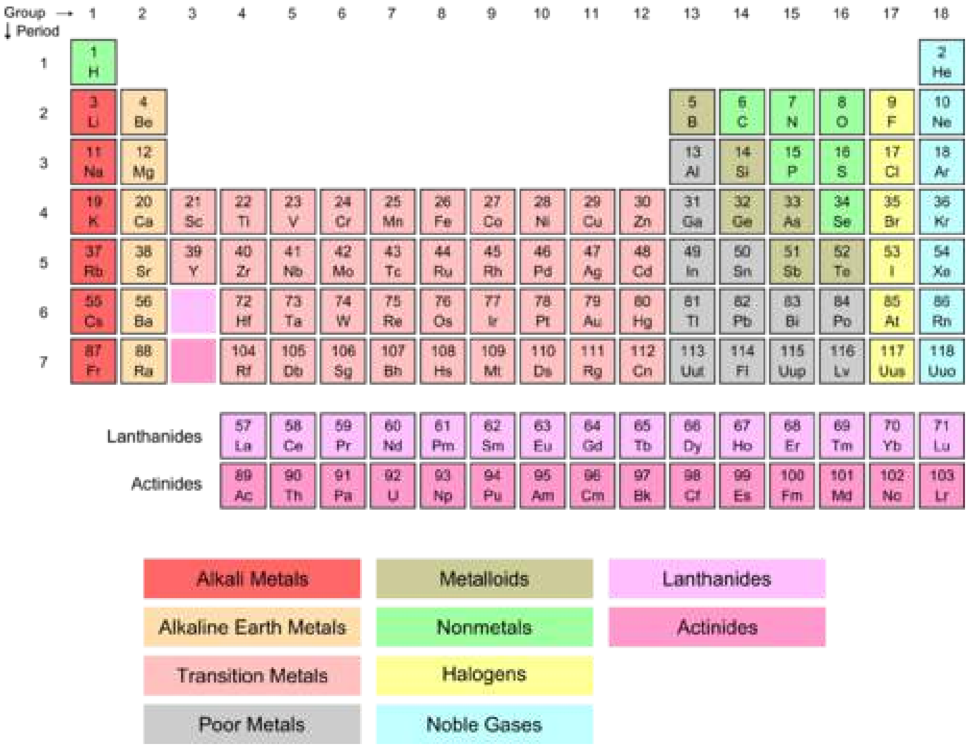
Chemical Families: Group One

Some chemistry students just enjoy learning about the science, while others are intrigued by the violent reactions that sometimes can occur. Many chemistry classes have been enlivened by the demonstration of how reactive sodium is with water. In some instances, the demonstration has gone off safely. Unfortunately, in other situations students and instructor have incurred serious injury due to their failure to observe proper safety precautions.

One of the values of the periodic table is the ability to make predictions about the behavior of individual elements. By knowing which group an element is in, we can determine the number of reactive electrons and say something about how that element will behave.

Hydrogen and the Alkali Metals

The periodic table is arranged on the basis of **atomic numbers (number of protons in the nucleus).** One of the valuable consequences of this arrangement is that we can learn a lot about the electron distribution in these atoms. The colors in the table below indicate the different groupings of atoms based on the location and number of electrons in the atom.



Group → ↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
Lanthanides				57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides				89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Alkali Metals	Metalloids	Lanthanides
Alkaline Earth Metals	Nonmetals	Actinides
Transition Metals	Halogens	
Poor Metals	Noble Gases	

If we look at Group I (red column), we see that it is labeled **alkali metals**. Also note the green H above the alkali metals. All these elements have a similar configuration of outer-shell electrons. In each case, there is one electron in the outer orbital.In the case of hydrogen (one proton in the nucleus and one electron), the arrangement is very simple. The lone electron exists in the first energy level around the nucleus. For lithium, there are two electrons in the first energy level and one electron in the second energy level. The same pattern holds for sodium and potassium.

TABLE 6.2:

Element	**Symbol**	**# electrons in Energy Levels**
hydrogen	H	1
lithium	Li	2,1
sodium	Na	2,8,1
potassium	K	2,8,8,1

This one electron is very easily removed during chemical reactions. The group I elements react rapidly with oxygen to produce metal oxides. They are very soft metals, with the exception of cesium which is a liquid at room temperature. The alkali metals also react readily with water to produce hydrogen gas and metal hydroxides in the following video.

http://youtu.be/QSZ-3wScePM

Li reacts with water to produce hydrogen gas. Sodium also reacts the same way, just more rapidly. Potassium reacts rapidly with water producing hydrogen gas and heat which ignites the hydrogen gas. Rubidium and cesium react yet more vigorously and explode on contact with water.

Summary

Group I (alkali metals and H) elements all have one electron in their outer shell.

This electron is in a s orbital.

The Group I metals are all very reactive with water.

Practice

Use the link below to answer the following questions:

http://chemed.chem.wisc.edu/chempaths/GenChem-Textbook/Group-IA-Alkali-Metals-544.html

How is hydrogen similar to the alkali metals? How is it different?

Why don't we know much about francium (atomic number 87)?

Describe the physical properties of the alkali metals.

Review

What group are the alkali metals and hydrogen in?

What is the outer shell electron configuration in this group?

Which alkali metal is a liquid at room temperature?

How reactive are the alkali metals with oxygen?

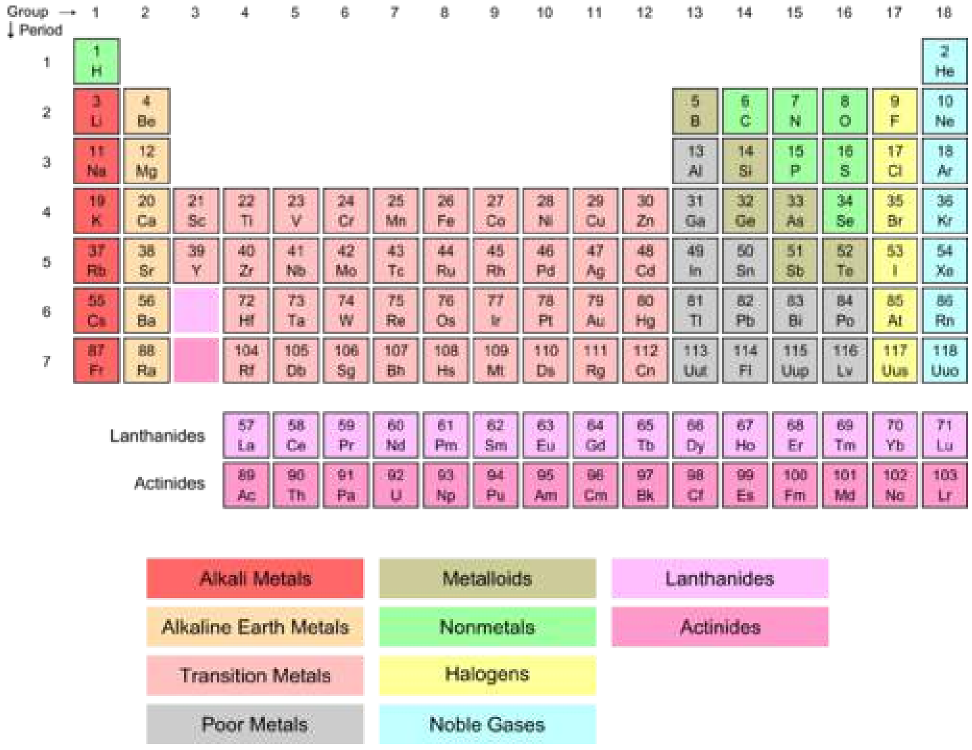
How reactive are these metals with water?

Chemical Families: Group Two

We take a lot of chemistry for granted. Very few of us think about the chemistry of bone or oyster shells. Both of these materials have large amounts of calcium compounds in them and play important roles in maintaining the structure of the organism. The shell provides a solid surrounding for the oyster. Bones give support to the body so the person can move around and not just be a soft mass of tissue.

Alkaline Earth Metals

Group II elements are referred to as " **alkaline earth**" metals (tan column below). The name " **alkaline**" comes from the fact that compounds of these elements form **basic** (pH greater than 7) or alkaline solutions when dissolved in water. If the **Group I** elements all have **one** electron in their outer orbital, we can predict that the **Group II** elements will have **two electrons** in that outer shell.



Group → ↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo

Lanthanides	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
Actinides	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

The beryllium atom, the first element of Group II, has an atomic number of four. The atom has the two electrons in the first shell(filled) as well as two electrons in the second shell, giving a total of four electrons. Note that there are two electrons in the outer shell, a structure that is characteristic of the Group II elements. Barium (atomic number 56) has the same outer shell structure of two electrons, even though the internal electron structure for barium is quite complicated.

The Group II elements tend to be less reactive than their Group I counterparts. The need to remove two electrons in order for the material to react means more energy is needed for electron removal. However, these elements are reactive enough that they do not exist in their elemental forms in nature, but are present as compounds.

Uses of Alkaline Earth Compounds

Since magnesium burns brightly, it is used in flares and fireworks. Magnesium alloys with aluminum provide light weight and sturdy materials for airplanes, missiles, and rockets. Several antacids use magnesium hydroxide to neutralize excess stomach acid.

Calcium compounds are widely found in limestone, marble, and chalk. Calcium is an important constituent of cement. Other uses include calcium chloride as a deicer and limestone as a white pigment in paints and toothpaste.

Strontium is widely used in fireworks and magnets. Barium compounds can be used in paints, filler for rubber, plastic, and resins, and as a contrast medium for X-rays. Many beryllium compounds are toxic, but these materials have been employed in metal alloys.

Summary

The alkaline earth elements are in Group II of the periodic table.

These elements each have two electrons in their outer shell.

The alkaline earth elements are less reactive than the alkali metals.

Practice

Use the link below to answer the following questions:

http://www.rsc.org/chemsoc/visualelements/pages/data/intro_groupii_data.html

What colour are all the alkaline earth elements?

In what compounds is magnesium found in nature?

In what compounds is calcium found in nature?

Review

Why are these elements known as “alkaline earth” elements?

How many electrons are in the outer shell of the alkaline earth elements?

Are the alkaline earth elements more or less reactive than the alkali metals? Explain your answer?

Is radium usually considered as part of the alkaline earth category in terms of chemistry? Explain your answer.

6.3 Electron Configuration and the Periodic Table

Lesson Objectives

- Understand the relationship between electron sublevels and the length of periods of the periodic table.
- Identify each block of the periodic table and be able to determine which block each element belongs to based on its electron configuration.
- Describe the relationship between outer electron configuration and group number. Be able to determine the number of valence electrons.
- Identify which groups are identified by the common names: alkali metals, alkaline earth metals, halogens, and noble gases.
- Locate transition elements, lanthanides, and actinides.

Lesson Vocabulary

- actinide
- alkali metal
- alkaline earth metal
- halogen
- inner transition element
- lanthanide
- noble gas
- representative (main-group) elements
- transition element

The development of the periodic table was largely based on elements that display similar chemical behavior. In the modern table, these elements are found in vertical columns called groups. In this lesson, you will see how the form of the periodic table is related to electron configurations, which in turn influences chemical reactivity.

Periods and Blocks

There are seven horizontal rows of the periodic table, called periods. The length of each period is determined by the number of electrons that are capable of occupying the sublevels that fill during that period, as seen in **Table** 6.3.

TABLE 6.3: **Period Length and Sublevels in the Periodic Table**

Period	Number of Elements in Period	Sublevels in Order of Fill
1	2	$1s$
2	8	$2s$ $2p$
3	8	$3s$ $3p$
4	18	$4s$ $3d$ $4p$

TABLE 6.3: (continued)

Period	Number of Elements in Period	Sublevels in Order of Fill
5	18	5*s* 4*d* 5*p*
6	32	6*s* 4*f* 5*d* 6*p*
7	32	7*s* 5*f* 6*d* 7*p*

Recall that the four different sublevels (*s*, *p*, *d*, *f*) each consist of a different number of orbitals. The *s* sublevel has one orbital, the *p* sublevel has three orbitals, the *d* sublevel has five orbitals, and the *f* sublevel has seven orbitals. In the first period, only the 1*s* sublevel is being filled. Since all orbitals can hold two electrons, the entire first period consists of just two elements. In the second period, the 2*s* sublevel, with two electrons, and the 2*p* sublevel with six electrons, are being filled. Consequently, the second period contains eight elements. The third period is similar to the second, filling the 3*s* and 3*p* sublevels. Notice that the 3*d* sublevel does not actually fill until after the 4*s* sublevel. This results in the fourth period containing 18 elements due to the additional 10 electrons that are contributed by the *d* sublevel. The fifth period is similar to the fourth. After the 6*s* sublevel fills, the 4*f* sublevel with its 14 electrons fills. This is followed by the 5*d* and the 6*p*. The total number of elements in the sixth period is 32. The later elements in the seventh period are still being created. So while there are a possible of 32 elements in the period, the current number is slightly less.

The period to which a given element belongs can easily be determined from its electron configuration. As an example, consider the element nickel (Ni). Its electron configuration is $[\mathrm{Ar}]3d^8 4s^2$. The highest occupied principal energy level is the fourth, indicated by the 4 in the $4s^2$ portion of the configuration. Therefore, nickel can be found in the fourth period of the periodic table. **Figure** 6.9 shows a version of the periodic table which includes abbreviated electron configurations.

Based on electron configurations, the periodic table can be divided into blocks denoting which sublevel is in the process of being filled. The *s*, *p*, *d*, and *f* blocks are illustrated below in **Figure** 6.10.

Figure 6.10 also illustrates how the *d* sublevel is always one principal level behind the period in which that sublevel occurs. In other words, the 3*d* sublevels fills during the fourth period. The *f* sublevel is always two levels behind. The 4*f* sublevel belongs to the sixth period.

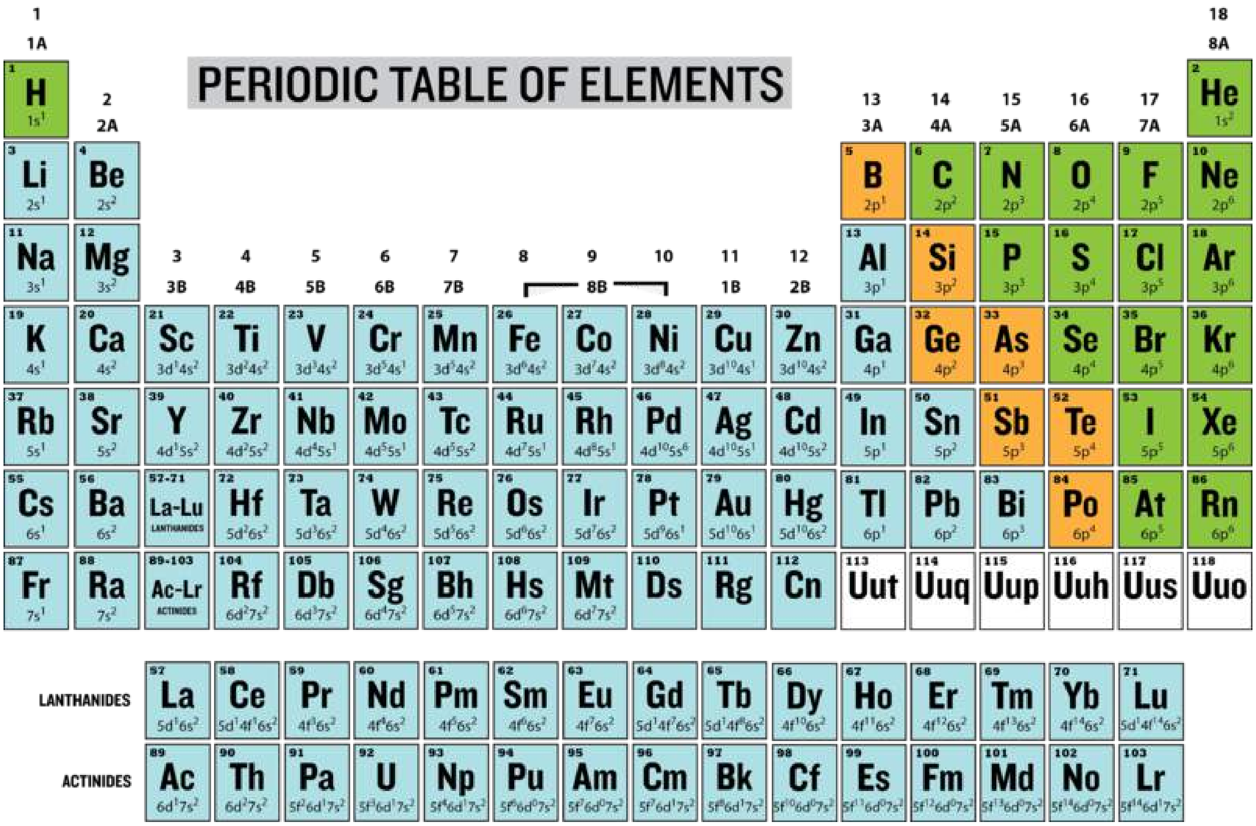
Representative Elements

We will now examine each of these blocks in more detail. The *s* and *p* sublevels are always filling during the period which corresponds to that element's highest principal energy level. That is, the second period is where the 2*s* and 2*p* sublevels fill. The *s-block elements and the p-block elements are together called the* **representative** or **main-group elements**.

The

The *s* block consists of the elements in Group 1 and Group 2. These groups consist of highly reactive metals. The *elements in Group 1 (lithium, sodium, potassium, rubidium, cesium, and francium) are called the* **alkali metals**. All of the alkali metals have a single *s* electron in their outermost principal energy. Recall that such electrons are called valence electrons. The general form of the electron configuration of each of the alkali metals is ns^1, where the n refers to the highest occupied principal energy level. As an example, the electron configuration of lithium (Li), the alkali metal of Period 2 is $1s^2 2s^1$. This single valence electron is what gives the alkali metals their extreme reactivity. **Figure** 6.11 shows the element sodium.

All alkali metals are very soft and can be cut easily with a knife. Their high reactivity means that they must be stored under oil in order to prevent them from reacting with air or with water vapor. The reactions between alkali metals



PERIODIC TABLE OF ELEMENTS

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H $1s^1$																	2 He $1s^2$
3 Li $2s^1$	4 Be $2s^2$											5 B $2p^1$	6 C $2p^2$	7 N $2p^3$	8 O $2p^4$	9 F $2p^5$	10 Ne $2p^6$
11 Na $3s^1$	12 Mg $3s^2$											13 Al $3p^1$	14 Si $3p^2$	15 P $3p^3$	16 S $3p^4$	17 Cl $3p^5$	18 Ar $3p^6$
19 K $4s^1$	20 Ca $4s^2$	21 Sc $3d^1 4s^2$	22 Ti $3d^2 4s^2$	23 V $3d^3 4s^2$	24 Cr $3d^5 4s^1$	25 Mn $3d^5 4s^2$	26 Fe $3d^6 4s^2$	27 Co $3d^7 4s^2$	28 Ni $3d^8 4s^2$	29 Cu $3d^{10} 4s^1$	30 Zn $3d^{10} 4s^2$	31 Ga $4p^1$	32 Ge $4p^2$	33 As $4p^3$	34 Se $4p^4$	35 Br $4p^5$	36 Kr $4p^6$
37 Rb $5s^1$	38 Sr $5s^2$	39 Y $4d^1 5s^2$	40 Zr $4d^2 5s^2$	41 Nb $4d^4 5s^1$	42 Mo $4d^5 5s^1$	43 Tc $4d^5 5s^2$	44 Ru $4d^7 5s^1$	45 Rh $4d^8 5s^1$	46 Pd $4d^{10} 5s^6$	47 Ag $4d^{10} 5s^1$	48 Cd $4d^{10} 5s^2$	49 In $5p^1$	50 Sn $5p^2$	51 Sb $5p^3$	52 Te $5p^4$	53 I $5p^5$	54 Xe $5p^6$
55 Cs $6s^1$	56 Ba $6s^2$	57-71 La-Lu LANTHANIDES	72 Hf $5d^2 6s^2$	73 Ta $5d^3 6s^2$	74 W $5d^4 6s^2$	75 Re $5d^5 6s^2$	76 Os $5d^6 6s^2$	77 Ir $5d^7 6s^2$	78 Pt $5d^9 6s^1$	79 Au $5d^{10} 6s^1$	80 Hg $5d^{10} 6s^2$	81 Tl $6p^1$	82 Pb $6p^2$	83 Bi $6p^3$	84 Po $6p^4$	85 At $6p^5$	86 Rn $6p^6$
87 Fr $7s^1$	88 Ra $7s^2$	89-103 Ac-Lr ACTINIDES	104 Rf $6d^2 7s^2$	105 Db $6d^3 7s^2$	106 Sg $6d^4 7s^2$	107 Bh $6d^5 7s^2$	108 Hs $6d^6 7s^2$	109 Mt $6d^7 7s^2$	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

LANTHANIDES	57 La $5d^1 6s^2$	58 Ce $5d^1 4f^1 6s^2$	59 Pr $4f^3 6s^2$	60 Nd $4f^4 6s^2$	61 Pm $4f^5 6s^2$	62 Sm $4f^6 6s^2$	63 Eu $4f^7 6s^2$	64 Gd $5d^1 4f^7 6s^2$	65 Tb $5d^1 4f^8 6s^2$	66 Dy $4f^{10} 6s^2$	67 Ho $4f^{11} 6s^2$	68 Er $4f^{12} 6s^2$	69 Tm $4f^{13} 6s^2$	70 Yb $4f^{14} 6s^2$	71 Lu $5d^1 4f^{14} 6s^2$
ACTINIDES	89 Ac $6d^1 7s^2$	90 Th $6d^2 7s^2$	91 Pa $5f^2 6d^1 7s^2$	92 U $5f^3 6d^1 7s^2$	93 Np $5f^4 6d^1 7s^2$	94 Pu $5f^6 6d^0 7s^2$	95 Am $5f^7 6d^0 7s^2$	96 Cm $5f^7 6d^1 7s^2$	97 Bk $5f^8 6d^1 7s^2$	98 Cf $5f^{10} 6d^0 7s^2$	99 Es $5f^{11} 6d^0 7s^2$	100 Fm $5f^{12} 6d^0 7s^2$	101 Md $5f^{13} 6d^0 7s^2$	102 No $5f^{14} 6d^0 7s^2$	103 Lr $5f^{14} 6d^1 7s^2$

FIGURE 6.9

This periodic table shows the outer electron configurations of the elements.

and water is particularly vigorous and spectacular. The reaction rapidly produces large quantities of hydrogen gas. Alkali metals also react easily with most nonmetals. All of the alkali metals are far too reactive to be found in nature in their "free" or uncombined states. For example, all naturally occurring sodium exists as one or many compounds of sodium such as sodium chloride –table salt.

The elements in Group 2 (beryllium, magnesium, calcium, strontium, barium, and radium) are called the **alkaline earth metals** (see **Figure** 6.12). These elements have two valence electrons, both belonging to the outermost *s* sublevel. The general electron configuration of all alkaline earth metals is ns^2. The alkaline earth metals are still too reactive to exist in nature as free elements, but are less reactive than the alkali metals. They are also harder, stronger, and denser than the alkali metals. They also make many compounds with nonmetals.

Watch video experiments of *s* block elements:

- Reaction of Lithium with Water at http://www.youtube.com/watch?v=Vxqe_ZOwsHs
- Reaction of Lithium with Oxygen at http://www.youtube.com/watch?v=mXv38UvP_tQ
- Reaction of Sodium with Water at http://www.youtube.com/watch?v=dmcfsEEogxs
- Reaction of Sodium with Oxygen at http://www.youtube.com/watch?v=hxrQrOlWGH0
- Reaction of Potassium with Water at http://www.youtube.com/watch?v=oqMN3y8k9So
- Reaction of Potassium with Oxygen at http://www.youtube.com/watch?v=pT0w6BRfDGs
- Reaction of Magnesium with Oxygen at http://www.youtube.com/watch?v=NnFzHt6l4z8
- Reaction of Calcium with Water at http://www.youtube.com/watch?v=i-rFsFwdkTU

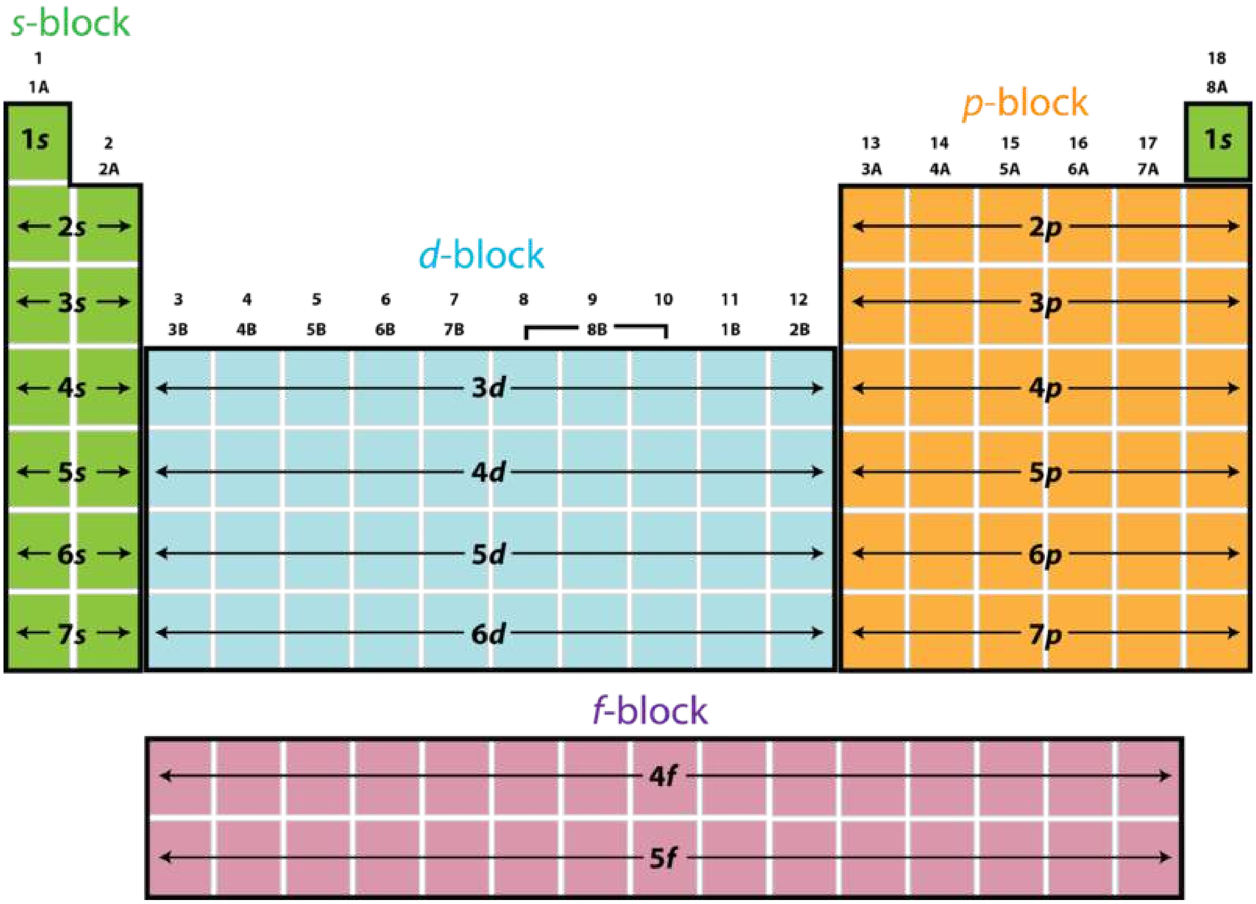


FIGURE 6.10
A block diagram of the periodic table shows which sublevels are being filled at any point.

FIGURE 6.11
Sodium, like all alkali metals, is very soft. A fresh surface, exposed from cutting, exhibits luster that is quickly lost as the sodium reacts with air.

FIGURE 6.12
Alkaline earth metals: beryllium, magnesium, calcium, strontium, and barium. Strontium and barium react with air and must be stored in oil.

Hydrogen and Helium

Looking at the block diagram (**Figure** 6.10), you may be wondering why hydrogen and helium were not included in the alkali metal and alkaline earth metal groups. Though hydrogen, with its $1s^1$ configuration, appears as though it should be similar to the rest of Group 1, it does not share the properties of that group. Hydrogen is a unique element which is not reasonably included with any other group of the periodic table. Some periodic tables even separate hydrogen's square from the rest of Group 1 to indicate its solitary status.

Helium has a configuration of $1s^2$, which would seem to place it with the alkaline earth metals. However, it is instead placed in Group 18 at the far right of the periodic table. This group, called the noble gases, are very unreactive because their outermost s and p sublevels are completely filled. Helium, being in Group 1, does not have a p sublevel. Its filled $1s$ sublevel makes it very similar to the other members of Group 18.

The

The p block consists of the elements in groups 13-18 except for helium. The p sublevel always fills after the s sublevel of a given principal energy level. Therefore, the general electron configuration for an element in the p block is ns^2np^{1-6}. For example, the electron configuration of elements in Group 13 is ns^2np^1. This continues for the remainder of the groups. *The elements of Group 18 (helium, neon, argon, krypton, xenon, and radon) are called the* **noble gases**. They are an especially important group of the periodic table because they are almost completely unreactive owing to their completely filled outermost s and p sublevels. As noted above, helium is somewhat of an exception with a configuration of $1s^2$, while all of the other noble gases have configurations of ns^2np^6. The noble gases were not a part of Mendeleev's periodic table because they had not yet been discovered. In 1894, English physicist Lord Rayleigh and Scottish chemist Sir William Ramsay detected argon as a small percentage of the atmosphere. Discovery of the other noble gases soon followed. The group was originally called the *inert gases* because it was thought that it was impossible for any of them to react and form compounds. Beginning in the early 1960s, several compounds of xenon and highly reactive fluorine were synthesized. The name of the group was changed to noble gases.

The number of valence electrons in elements of the p block is equal to the group number minus 10. As an example, sulfur is located in Group 16 and has 16 –10 = 6 valence electrons. Since sulfur is located in period 3, its outer electron configuration is $[Ne]3s^23p^4$. In the older system of labeling groups with A and B designations, the representative elements are designated 1A through 8A. The number of valence electrons is equal to the group number preceding the A. Sulfur is a member of Group 6A.

The properties of the *p* block elements exhibit a wide variation. The stair-step line separating metals from nonmetals runs through the *p* block. As a result, there are 8 metals, all 7 metalloids, and all 15 nonmetals. Note that there is some variation among different periodic tables over how to classify the rare elements polonium and astatine. The metals of the *p* block are much more stable than the *s* block metals. Aluminum and tin are frequently used in packaging. Lead (**Figure** 6.13) is used in car batteries, bullets, and as radiation shielding.

FIGURE 6.13

Lead blocks are used in radiation shielding.

The elements of Group 17 (fluorine, chlorine, bromine, iodine, and astatine) are called the **halogens**. The halogens all have the general electron configuration ns^2np^5, giving them seven valence electrons. They are one electron short of having the full outer *s* and *p* sublevel, which makes them very reactive. They undergo especially vigorous reactions with the reactive alkali metals. As elements, chlorine and fluorine are gases at room temperature, bromine is a dark orange liquid, and iodine is a dark purple-gray solid. Astatine is so rare that its properties are mostly unknown.

Watch video experiments of *p* block elements:

- Superconductors! at http://education.jlab.org/frost/superconductor.html
- The Flying Ring at http://education.jlab.org/frost/ring_fling.html
- Superconducting BAYCO Bits at http://education.jlab.org/scienceseries/superconducitng_bayco_bits.html
- Halogen Activity Series at http://www.dlt.ncssm.edu/core/Chapter8-Atomic_Str_Part2/bleachoverlay-lg.htm
- Reaction of Bromine with Aluminum at http://www.youtube.com/watch?v=zZAiYc3WVXM
- Halogen Displacement Reactions at http://www.youtube.com/watch?v=cbFCWFksYoM

Transition Elements

Transition elements *are the elements that are found in Groups 3-12 on the periodic table*. The term refers to the fact that the *d* sublevel, which is in the process of being filled, is in a lower principal energy level than the *s* sublevel filled before it. For example, the electron configuration of scandium, the first transition element, is $[\text{Ar}]3d^{14}s^2$. Remember that the configuration is reversed from the fill order –the 4*s* filled before the 3*d* begins. Because they are all metals, the transition elements are often called the transition metals. As a group, they display typical metallic properties and are less reactive than the metals in Groups 1 and 2. Some of the more familiar ones are so unreactive that they can be found in nature in their free, or uncombined state (**Figure** 6.14). These include platinum, gold, and silver.

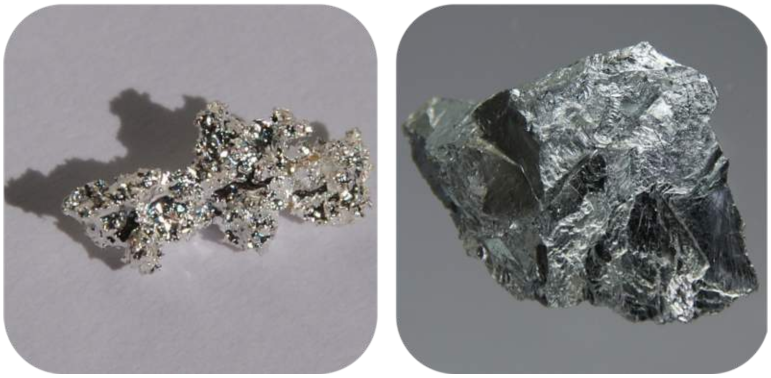


FIGURE 6.14

Silver (*left*) and chromium (*right*) are two typical transition metals.

Compounds of many transition elements are distinctive for being widely and vividly colored. Electron transitions that occur within the *d* sublevel release energies that result in the emission of visible light of varied wavelengths (**Figure** 6.15).

FIGURE 6.15

Transition metal compounds dissolved in water exhibit a wide variety of bright colors. From left to right are shown solutions of cobalt(II) nitrate, potassium dichromate, potassium chromate, nickel(II) chloride, copper(II) sulfate, and potassium permanganate.

The

The transition elements found in Groups 3-12 are also referred to as the *d* block, since the *d* sublevel is in the process of being filled. Since there are five *d* orbitals that can accommodate ten electrons, there are ten elements in each period of the *d* block. The general electron configuration for elements in the *d* block is $(n - 1)d^{1-10}ns^2$. The *d* sublevel is always in a principal energy level that is one lower than that of the *s* sublevel. For example, the configuration of zirconium (Zr) is $[\text{Kr}]4d^2 5s^2$. The group number can easily be determined from the combined number of electrons in the *s* and *d* sublevel. Zirconium is in Period 5 and Group 4. Recall from an earlier chapter that there are several deviations from the expected order of filling the *d* sublevel that cannot always be easily understood. The element cobalt (Co) is in Period 4 and Group 9. It has the expected electron configuration of $[\text{Ar}]3d^7 4s^2$. Directly below cobalt in the group is the element rhodium (Rh). However, its configuration is $[\text{Kr}]4d^8 5s^1$, meaning that one of its 5*s* electrons has moved to the 4*d* sublevel. The total of nine electrons still allows you to predict that rhodium is a member of Group 9.

Because electrons in the *d* sublevel do not belong to the outermost principal energy level, they are not valence electrons. Most *d* block elements have two valence electrons, which are the two electrons from the outermost *s* sublevel. Rhodium is an example of a transition metal with only one valence electron because its configuration deviates from the expected filling order.

The

The first of the f sublevels to begin filling is the 4f sublevel. It fills after the 6s sublevel, meaning that f sublevels are two principal energy levels behind. The general electron configuration for elements in the f block is $(n - 2)f^{1-14}ns^2$. The seven orbitals of the f sublevel accommodate 14 electrons, so the f block is 14 elements in length. It is pulled out of the main body of the period table and is shown at the very bottom. Because of that, the elements of the f block do not belong to a group, being wedged in between Groups 3 and 4. The **lanthanides** *are the 14 elements from cerium (atomic number 58) to lutetium (atomic number 71).* The 4f sublevel is in the process of being filled for the lanthanides. They are all metals and are similar in reactivity to the Group 2 alkaline earth metals.

The **actinides** *are the 14 elements from thorium (atomic number 90) to lawrencium (atomic number 103).* The 5f sublevel is in the process of being filled. The actinides are all radioactive elements and only the first four have been found naturally on Earth. All of the others have only been artificially made in the laboratory. *The lanthanides and actinides together are sometimes called the* **inner transition elements**.

Sample Problem 6.1: Electron Configurations and the Periodic Table

The electron configurations of atoms of four different elements are shown below. Without consulting the periodic table, name the period, group, and block where this element is located. Determine the number of valence electrons for each. Then, using a periodic table, name the element and identify it as a metal, nonmetal, or metalloid.

1. $[\mathrm{Kr}]4d^{10}5s^25p^3$
2. $[\mathrm{Rn}]5f^77s^2$
3. $[\mathrm{Ar}]4s^2$
4. $[\mathrm{Xe}]4f^{14}5d^66s^2$

Step 1: Plan the problem.

The period is the highest occupied principal energy level. The group is the vertical column. The block depends on which sublevel is in the process of being filled. The valence electrons are those in the outermost principal energy level. For name and type of element, use a periodic table.

Step 2: Solutions

1. The highest occupied principal energy level is the fifth, so this element is in Period 5. The group is found by adding $10 + 2 + 3$ from the configuration. The element is in Group 15. Since the p sublevel is not filled, it is in the p block. There are five electrons in the outermost energy level so it has 5 valence electrons. The element is antimony, a metalloid.
2. The element is in Period 7. Since the f sublevel is incompletely filled, it is part of the f block and also does not belong to a group. It has 2 valence electrons. The element is americium, a metal from the actinides.
3. The element is in Period 4 and Group 2. Even though the 4s sublevel is filled, the last electron went into that sublevel, making it a member of the s block. It has 2 valence electrons. The element is calcium, a metal.
4. The element is in Period 6. In determining the group, the f electrons can be ignored since they do not affect groups. So $6 + 2 = 8$ and the element is in Group 8. The incompletely filled d sublevel makes it a member of the d block. It has 2 valence electrons. The element is osmium, a metal (specifically, a transition metal).

Step 3: Think about your result.

Once you have identified each element, you can use the periodic table which contains the electron configurations to make sure it has been identified correctly.

Practice Problems

1. For each of the following, identify the period, group, and block. Determine the number of valence electrons. Identify the element by name and type (metal, nonmetal, or metalloid).
   a. $[\mathrm{Kr}]4d^{10}5s^2$
   b. $[\mathrm{Ne}]3s^2 3p^2$
2. For the two elements in problem 1 above, how many unpaired electrons are in the atom?
3. Which two elements have 3 unpaired electrons in the $3d$ sublevel?

You can watch video lectures on this topic from Kahn Academy. These videos combine orbital configurations (from the chapter *Electrons in Atoms*) with locations on the periodic table:

- Introduction to Orbitals at http://www.youtube.com/watch?v=yBrp8uvNAhI (13:38)
- More on Orbitals and Electron Configuration at http://www.youtube.com/watch?v=FmQoSenbtnU (14:31)
- Electron Configurations at http://www.youtube.com/watch?v=RJlEH5Jz80w (10:04)
- Electron Configurations 2 at http://www.youtube.com/watch?v=YURReI6OJsg (10:18)

Lesson Summary

- An element's placement in the periodic table is determined by its electron configuration.
- Chemical properties of elements can largely be explained by the outer electron configuration.
- The periodic table is divided into 4 blocks (s, p, d, f) based on what sublevel is in the process of being filled.
- Alkali metals, alkaline earth metals, halogens, and noble gases are common names given to several groups.
- Transition elements are members of the d block, while the f block consists of the lanthanides and the actinides.

Lesson Review Problems

Reviewing Concepts

1. How many elements are in the second period? The fourth? The sixth?
2. Use a periodic table to identify the block that each of these elements would be found.
   a. rubidium
   b. holmium
   c. palladium
   d. tellurium
3. Give the common name of each of the following groups.
   a. Group 17
   b. Group 2
   c. Group 1
   d. Group 18
4. Where on the periodic table are the most reactive metals?
5. Where on the periodic table are the most reactive nonmetals?
6. Why are the noble gases almost completely unreactive?

7. Which groups are called the representative elements?
8. What common name is given to the set of elements located in the *d* block?

Problems

9. Without referring to the periodic table, write the full electron configuration of the following elements.
    a. the element in Period 2 and Group 15
    b. the alkaline earth metal in Period 3
    c. the noble gas in Period 3
10. Identify the following elements by name.
    a. the halogen in Period 5
    b. the alkali metal in Period 4
    c. the element in Period 4 that has two electrons in the *p* sublevel
    d. the second period element with six valence electrons
    e. the lanthanide with three *f* sublevel electrons
    f. the element in Period 6 with all five *d* orbitals half-filled
11. Which element has 5 occupied principal energy levels and chemical properties that are similar to gallium?
12. Without looking at the periodic table, identify the period, group, and block for elements that have the electron configurations below. Determine the number of valence electrons. Using the periodic table, identify the element by name and type (metal, nonmetal, or metalloid).
    a. $[\text{Ar}]3d^{10}4s^2 4p^6$
    b. $[\text{Kr}]4d^6 5s^1$
    c. $[\text{He}]2s^2 2p^1$
    d. $[\text{Xe}]4f^{10}6s^2$
    e. $[\text{Xe}]4f^{14}5d^3 6s^2$
    f. $[\text{Ne}]3s^1$
13. Which two elements have one unpaired electron in the $4p$ sublevel?

6.4 Periodic Trends

Lesson Objectives

- Learn the periodic trends for atomic radius.
- Know the relationship of group number to valence electrons.
- Describe how ions are formed.
- Learn the periodic trends for ionization energy.
- Explain how multiple ionization energies are related to noble gas electron configurations.
- Describe electron affinity.
- Predict the effect that ion formation has on the size of an atom.
- Learn the periodic trends for electronegativity.

Lesson Vocabulary

- anion
- atomic radius
- cation
- electron affinity
- electronegativity
- ion
- ionization energy

So far, you have learned that the elements are arranged in the periodic table according to their atomic number and that elements in vertical groups share similar electron configurations and chemical properties. In this lesson, we will explore various measurable properties of elements and the variance of those properties among elements. Specifically, we will examine trends within periods and groups. A trend is a general increase or decrease in a particular measurable quantity. For example, as the calendar moves from August to December in the northern hemisphere, the trend is for the average daily temperature to decrease. That doesn't mean that the temperature drops every single day, just that the overall direction is generally downward.

Atomic Radius

The size of an atom is defined by the edge of its orbital. However, orbital boundaries are fuzzy and in fact are variable under different conditions. In order to standardize the measurement of atomic radii, the distance between the nuclei of two identical atoms bonded together is measured. The **atomic radius** *is defined as one-half the distance between the nuclei of identical atoms that are bonded together* (**Figure** 6.16).

Atomic radii have been measured for elements, as shown in **Figure** 6.17. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be 74 pm. Therefore, the atomic radius of a hydrogen atom is 74/2 = 37 pm.

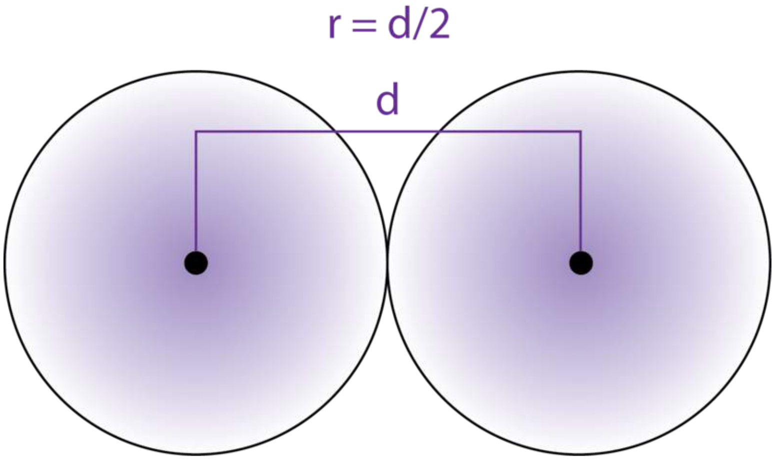


FIGURE 6.16

The atomic radius (*r*) of an atom can be defined as one half the distance (*d*) between two nuclei in a diatomic molecule.

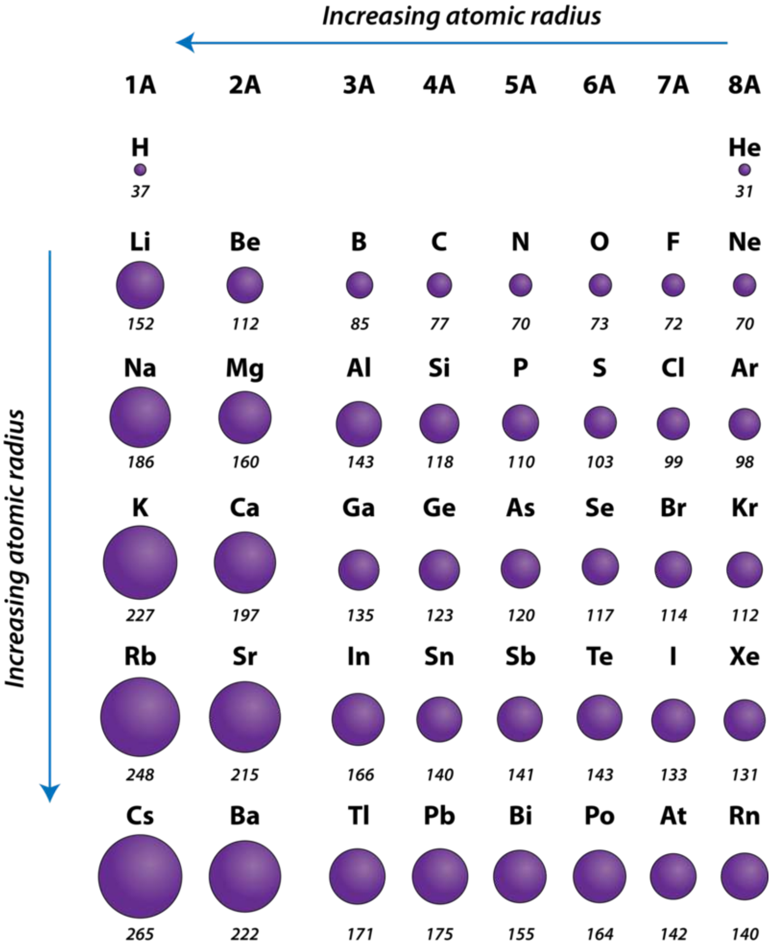


FIGURE 6.17

Atomic radii of the representative elements measured in picometers.

Periodic Trend

As you can see from **Figure** 6.17, the atomic radius of atoms generally decreases from left to right across a period. There are some small exceptions, such as the oxygen radius being slightly greater than the nitrogen radius. Within a period, protons are added to the nucleus as electrons are being added to the same principal energy level. These electrons are gradually pulled closer to the nucleus because of its increased positive charge. Since the force of attraction between nuclei and electrons increases, the size of the atoms decreases. The effect lessens as one moves further to the right in a period because of electron-electron repulsions that would otherwise cause the atom's size to increase.

Group Trend

As **Figure** 6.17 shows, the atomic radius of atoms generally increases from top to bottom within a group. As the atomic number increases down a group, there is again an increase in the positive nuclear charge. However, there is also an increase in the number of occupied principle energy levels. Higher principal energy levels consist of orbitals which are larger in size than the orbitals from lower energy levels. The effect of the greater number of principal energy levels outweighs the increase in nuclear charge and so atomic radius increases down a group.

Figure 6.18 shows a graph of atomic radius plotted versus atomic number. Each successive period is shown in a different color. You can see that as the atomic number increases within a period, the atomic radius decreases. As atomic number increases within a group, the atomic radius increases.

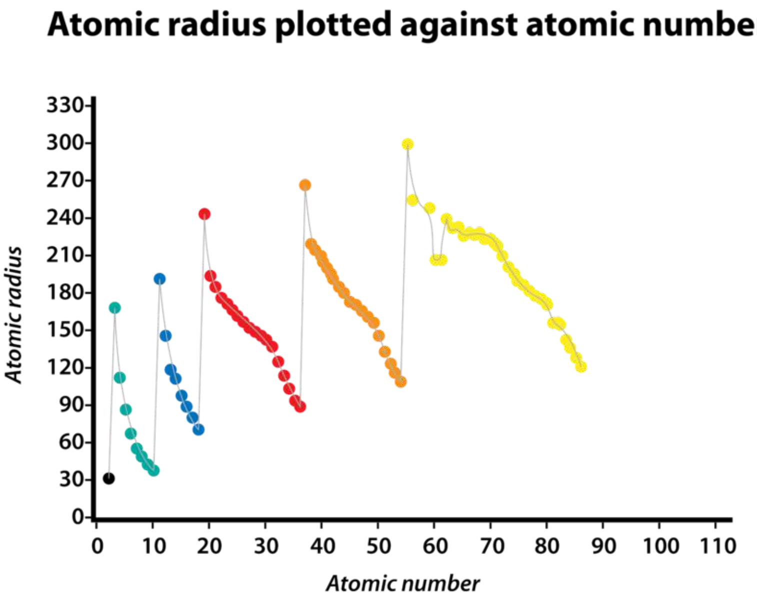


FIGURE 6.18
Graph of the atomic radii plotted against the atomic number.

Valence Electrons

We have previously defined the valence electrons as those in the outermost principal energy level. Since the outermost principal energy level consists only of *s* and *p* sublevels, the maximum number of valence electrons is eight. **Table** 6.4 shows the relationship of valence electrons to group number for the representative elements.

TABLE 6.4: **Valence Electrons of Representative Elements**

Group Number	Outer Electron Configuration	Number of Valence Electrons
1	ns^1	1
2	ns^2	2
13	ns^2np^1	3

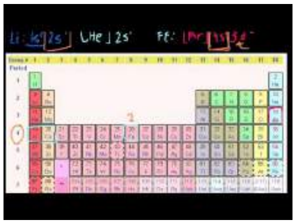
TABLE 6.4: (continued)

Group Number	Outer Electron Configuration	Number of Valence Electrons
14	ns^2np^2	4
15	ns^2np^3	5
16	ns^2np^4	6
17	ns^2np^5	7
18	ns^2np^6	8

You can see that the number of valence electrons is constant within a group because of the identical outer electron configuration. This is the reason why elements within a group share similar chemical properties.

Watch these video lectures:

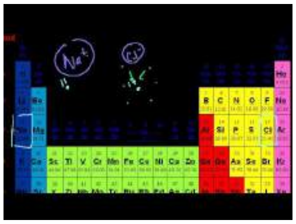
- Valence Electrons (15:23) at http://www.youtube.com/watch?v=1TZA171yxY4



MEDIA

Click image to the left for more content.

- Ion Size (3:04) at http://www.youtube.com/watch?v=HBi8xjMchZc



MEDIA

Click image to the left for more content.

Forming Ions

Many chemical compounds consist of particles called ions. An **ion** *is an atom or group of bonded atoms that has a positive or negative charge*. For our purposes, we will presently only consider monatomic ions, which are single atoms with an electrical charge. How do atoms obtain this charge? There are essentially two possibilities: (1) gaining or losing positively-charged protons, or (2) gaining or losing negatively-charged electrons. The nucleus of an atom is very stable and the number of protons is unchangeable in chemical reactions. As we have already seen, however, the electrons are capable of movement within an atom. Electron energy level transitions are very common and responsible for atomic emission spectra. It is the electrons which are either gained or lost to make ions (**Figure** 6.19).

When an atom loses one or more electrons it becomes positively charged because it now has more protons than electrons. *A positively charged ion is called a* **cation**. The charge for a cation is written as a numerical superscript after the chemical symbol, followed by a plus sign. If the ion carries a single unit of charge the number "1" is assumed and not written. For example, a sodium atom that loses one electron becomes a sodium ion, written as

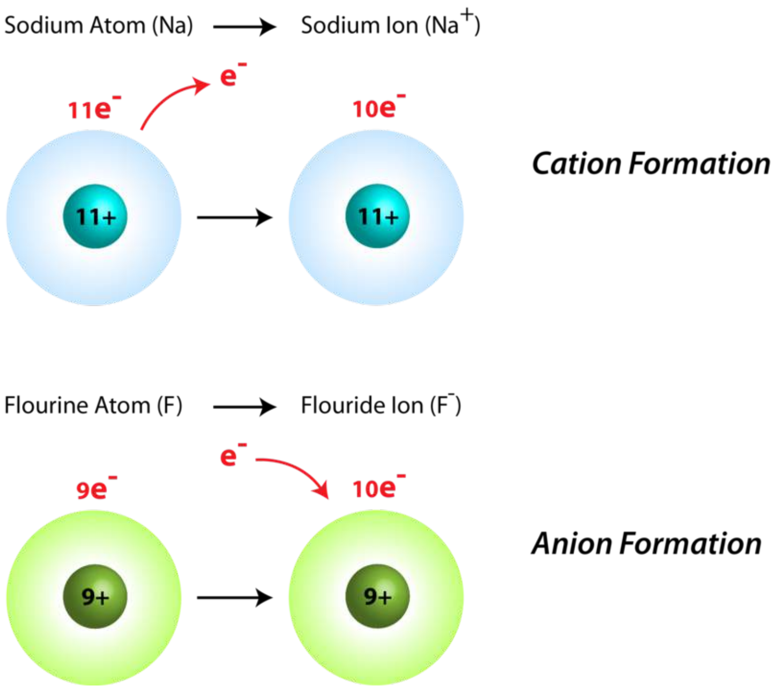


FIGURE 6.19

Na^+. A magnesium atom that loses two electrons becomes a magnesium ion, written as Mg^{2+}. The magnesium ion carries a 2+ charge because it would now have two more protons than electrons.

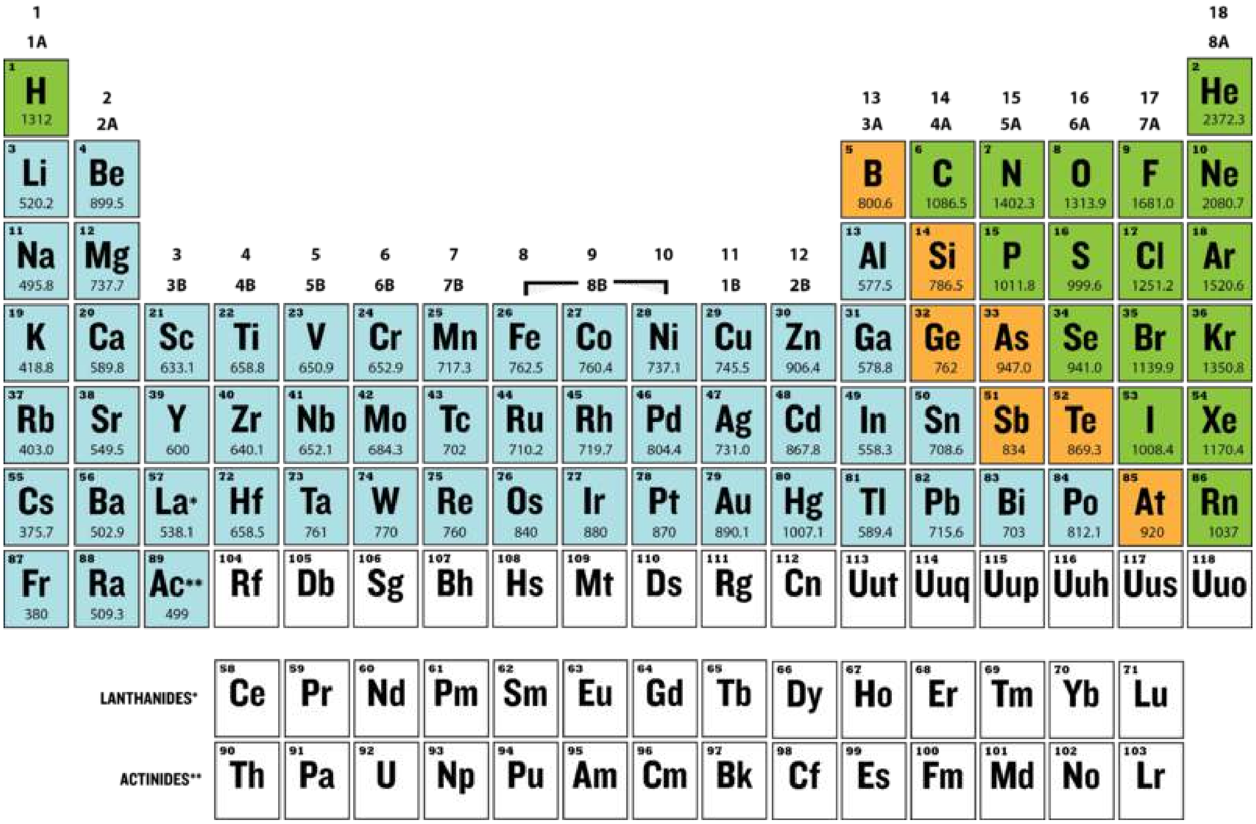
When an atom gains one or more electrons it becomes negatively charged because it now has more electrons than protons. *A negatively charged ion is called an* **anion**. The anion charge is written in the same way as for cations except with a minus sign. A chlorine atom that gains one electron becomes a chloride ion, written as Cl^-. Note that the names of monatomic anions have been given an "-ide" suffix. A sulfur atom that gains two electrons becomes a sulfide ion, written as S^{2-}.

Ionization Energy

To make an electron jump from a lower energy level to a higher energy level, there must be an input of energy. It stands to reason then, that removing the electron from the atom entirely requires even more energy. This is called an ionization process. **Ionization energy** *is the energy required to remove an electron from an atom*. An equation can be written to illustrate this process for a sodium atom.

$$\text{Na} + \text{energy} \rightarrow \text{Na}^+ + e^-$$

The equation shows that energy added to a sodium atom results in a sodium ion plus the removed electron (e^-). The electron that is removed from an atom is always a valence electron, since this electron is in the outermost principal energy level and is furthest from the nucleus. Elements have differing ionization energies (**Figure** 6.20), which are influenced by the size of the atom, the nuclear charge, and the electron energy levels. Ionization energies are measured in units of kilojoules per mole or kJ/mol.



1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1312																	2 He 2372.3
3 Li 520.2	4 Be 899.5											5 B 800.6	6 C 1086.5	7 N 1402.3	8 O 1313.9	9 F 1681.0	10 Ne 2080.7
11 Na 495.8	12 Mg 737.7											13 Al 577.5	14 Si 786.5	15 P 1011.8	16 S 999.6	17 Cl 1251.2	18 Ar 1520.6
19 K 418.8	20 Ca 589.8	21 Sc 633.1	22 Ti 658.8	23 V 650.9	24 Cr 652.9	25 Mn 717.3	26 Fe 762.5	27 Co 760.4	28 Ni 737.1	29 Cu 745.5	30 Zn 906.4	31 Ga 578.8	32 Ge 762	33 As 947.0	34 Se 941.0	35 Br 1139.9	36 Kr 1350.8
37 Rb 403.0	38 Sr 549.5	39 Y 600	40 Zr 640.1	41 Nb 652.1	42 Mo 684.3	43 Tc 702	44 Ru 710.2	45 Rh 719.7	46 Pd 804.4	47 Ag 731.0	48 Cd 867.8	49 In 558.3	50 Sn 708.6	51 Sb 834	52 Te 869.3	53 I 1008.4	54 Xe 1170.4
55 Cs 375.7	56 Ba 502.9	57 La* 538.1	72 Hf 658.5	73 Ta 761	74 W 770	75 Re 760	76 Os 840	77 Ir 880	78 Pt 870	79 Au 890.1	80 Hg 1007.1	81 Tl 589.4	82 Pb 715.6	83 Bi 703	84 Po 812.1	85 At 920	86 Rn 1037
87 Fr 380	88 Ra 509.3	89 Ac** 499	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

LANTHANIDES*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
ACTINIDES**	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

FIGURE 6.20

Periodic table showing the first ionization energies of elements, measured in kJ/mol.

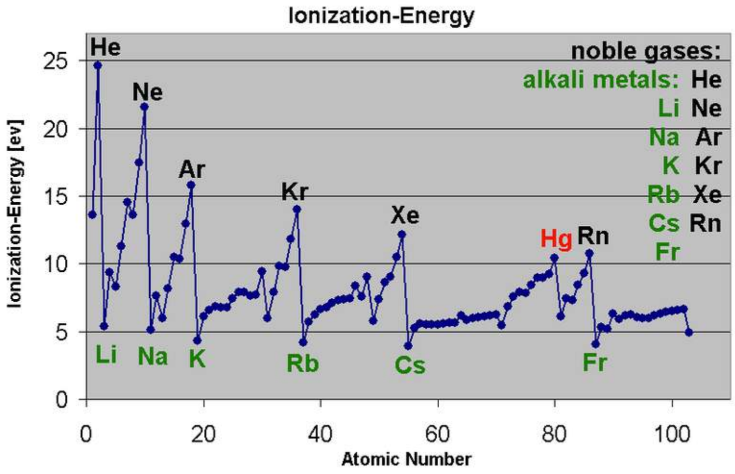


FIGURE 6.21

Graph of first ionization energy plotted against atomic number

Period Trend

As can be seen from **Figures** 6.20 and 6.21, the ionization energy of atoms generally increases from left to right across each row of the periodic table. The reason for this increase in ionization energy is the increase in nuclear charge. As more protons are added to the nucleus, the attraction for electrons increases. Because the electron that is to be removed is held more tightly to the nucleus, it is harder to remove it. This results in a larger ionization energy.

There are several exceptions to the general increase in ionization energy across a period. The elements of Group 13 (B, Al, etc.) have a lower ionization energy than the elements of Group 2 (Be, Mg, etc.). This is an illustration of a concept called *electron shielding*. Outer electrons are partially shielded from the attractive force of the protons in the nucleus by inner electrons (Figure 6.22).

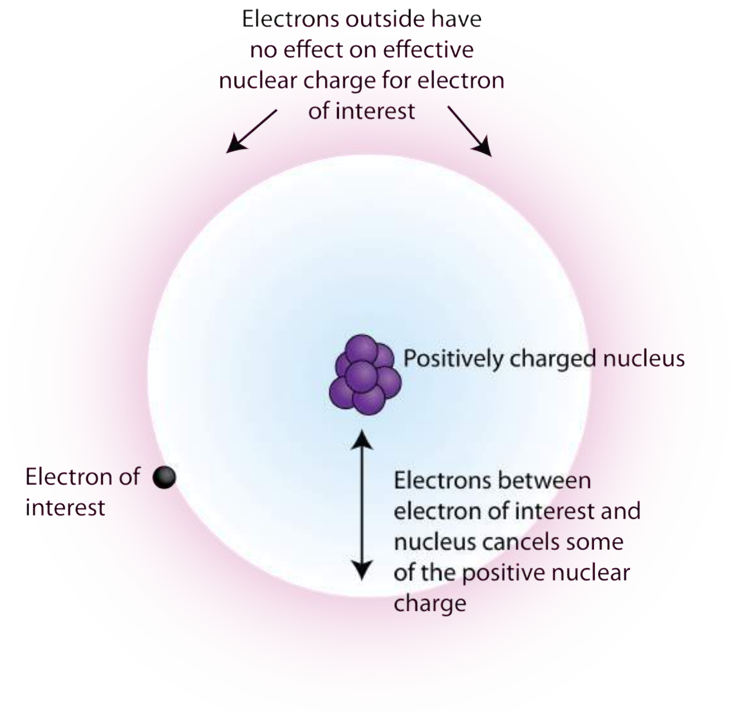


FIGURE 6.22

The shielding effect is shown by the interior electron cloud (in green) shielding the outer electron of interest from the full attractive force of the nucleus. A larger shielding effect results in a decrease in ionization energy.

To explain how shielding works, consider a lithium atom. It has three protons and three electrons –two in the first principal energy level and its valence electron in the second. The valence electron is partially shielded from the attractive force of the nucleus by the two inner electrons. Removing that valence electron becomes easier because of the shielding effect. There is also a shielding effect that occurs between sublevels within the same principal energy level. Specifically, an electron in the *s* sublevel is capable of shielding electrons in the *p* sublevel of the same principal energy level. This is because of the spherical shape of the *s* orbital. The reverse is not true –electrons in *p* orbitals do not shield electrons in *s* orbitals (**Figure** 6.23).

The electron being removed from an Al atom is a 3*p* electron, which is shielded by the two 3*s* electrons as well as all the inner core electrons. The electron being removed from a Mg atom is a 3*s* electron, which is only shielded by the inner core electrons. Since there is a greater degree of electron shielding in the Al atom, it is slightly easier to remove the valence electron and its ionization energy is less than that of Mg. This is despite the fact that the nucleus of the Al atom contains one more proton than the nucleus of the Mg atom.

There is another anomaly between Groups 15 and 16. Atoms of Group 16 (O, S, etc.) have lower ionization energies than atoms of Group 15 (N, P, etc.). Hund's rule is behind the explanation. In a nitrogen atom, there are three

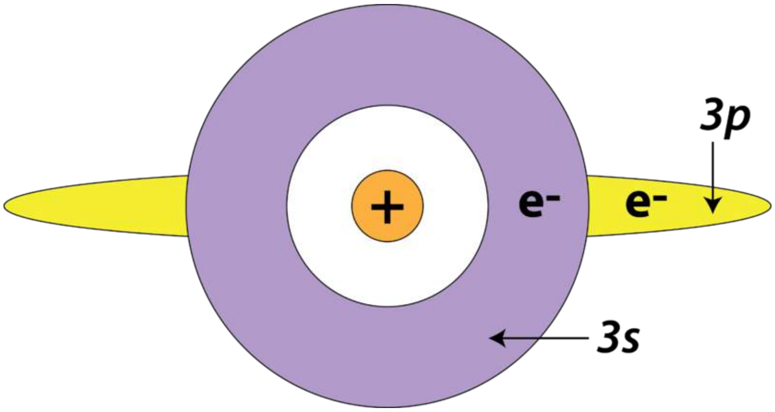


FIGURE 6.23

The spherical 3*s* orbital exhibits a shielding effect on the dumbbell shaped 3*p* orbital, that is of slightly higher energy. This reduces the ionization energy of a 3*p* electron compared to a 3*s* electron.

electrons in the 2*p* sublevel and each are unpaired. In an oxygen atom, there are four electrons in the 2*p* sublevel and so one orbital contains a pair of electrons. It is that second electron in the orbital that is removed in the ionization of an oxygen atom. Since electrons repel each other, it is slightly easier to remove the electron from the paired set in the oxygen atom than it is to remove an unpaired electron from the nitrogen atom.

Group Trend

The ionization energy of the representative elements generally decreases from top to bottom within a group. This trend is explained by the increase in size of the atoms within a group. The valence electron that is being removed in each atom is further from the nucleus for a larger atom. The attractive force between the valence electron and the nucleus weakens as the distance between them increases. This results in a lower ionization energy for the larger atoms within a group, even though the nuclear charge is increased. The shielding effect within a group increases because of a larger number of inner electrons. For all of the alkali metals, the single valence electron is shielded by all of the inner core electrons. This increase in shielding effect also makes it easier to remove the valence electron from the larger atom.

Multiple Ionizations

So far, we have described the first ionization energy of atoms and its trends. However, in many cases multiple electrons can be removed from atoms. If an atom loses two electrons, it acquires a 2+ charge. If an atom loses three electrons, it acquires a 3+ charge, and so on. The ionization energies required are called the second ionization energy (IE_2), third ionization energy (IE_3), etc. The first six ionization energies are shown for the elements of the first three periods in **Table** 6.5.

TABLE 6.5: **Ionization Energies (kJ/mol) of the First 18 Elements**

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
H	1312					
He	2373	5251				
Li	520	**7300**	11,815			
Be	899	1757	**14,850**	21,005		
B	801	2430	3660	**25,000**	32,820	
C	1086	2350	4620	6220	**38,000**	47,261
N	1400	2860	4580	7500	9400	**53,000**
O	1314	3390	5300	7470	11,000	13,000

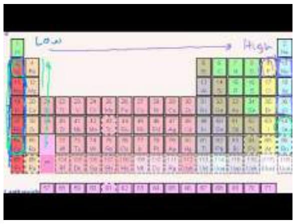
TABLE 6.5: (continued)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
F	1680	3370	6050	8400	11,000	15,200
Ne	2080	3950	6120	9370	12,200	15,000
Na	496	**4560**	6900	9540	13,400	16,600
Mg	738	1450	**7730**	10,500	13,600	18,000
Al	578	1820	2750	**11,600**	14,800	18,400
Si	786	1580	3230	4360	**16,000**	20,000
P	1012	1904	2910	4960	6240	**21,000**
S	1000	2250	3360	4660	6990	8500
Cl	1251	2297	3820	5160	6540	9300
Ar	1521	2666	3900	5770	7240	8800

Notice that the second ionization energy of an element is always higher than the first, the third is always higher than the second, and so on. This is because after one ionization, a positively charged ion is formed. Now there is a greater overall attractive force on the remaining electrons since the protons now outnumber the electrons. So to remove a second electron is more difficult.

The first ionization energies for the noble gases (He, Ne, Ar) are higher than those of any other element within that period. The noble gases have a full outer *s* and *p* sublevel, which gives them extra stability and means that they are unreactive. The stability of the noble gas electron configuration applies to other elements as well. Consider the element lithium, with its $1s^2 2s^1$ electron configuration. Being an alkali metal, its first ionization energy is very low. After it loses its valence electron (the 2*s* electron), it becomes a lithium ion, Li^+, and now has the electron configuration of $1s^2$. This is the electron configuration of the noble gas helium. The second ionization energy of lithium (shaded above) shows an extremely large jump compared to the first because the removal of a second electron requires breaking apart the noble gas electron configuration. The pattern continues across each period of the table. Beryllium shows a large jump after IE_2, boron after IE_3, and so on.

Watch a video lecture about Ionization Energy (12:13) at http://www.youtube.com/watch?v=ywqg9PorTAw .



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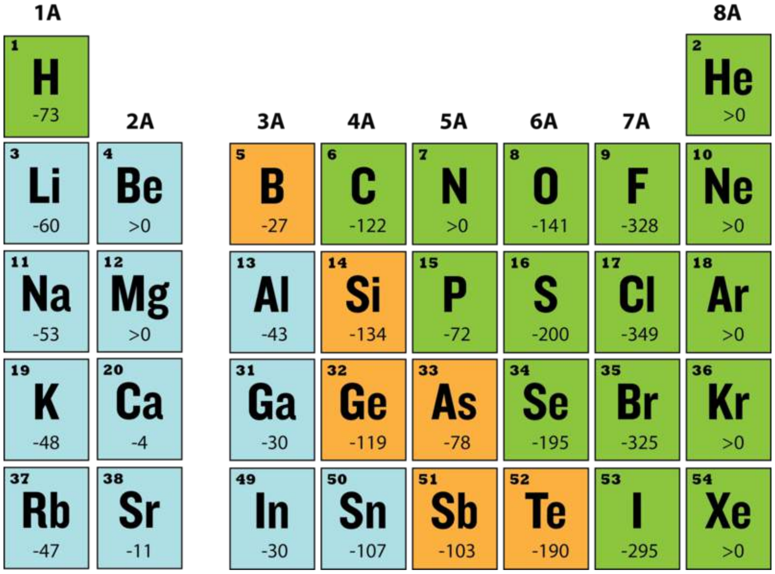
Electron Affinity

In most cases, the formation of an anion by the addition of an electron to a neutral atom releases energy. This can be shown for chloride ion formation below:

$$Cl + e^- \rightarrow Cl^- + \text{energy}$$

The energy change that occurs when a neutral atom gains an electron is called its **electron affinity**. When energy is released in a chemical reaction or process, that energy is expressed as a negative number. **Figure** 6.24 shows electron affinities in kJ per mole for the representative elements.

The elements of the halogen group (Group 17) gain electrons most readily, as can be seen from their large negative electron affinities. This means that more energy is released in the formation of a halide ion than for the anions of any



1A	2A	3A	4A	5A	6A	7A	8A
1 H -73							2 He >0
3 Li -60	4 Be >0	5 B -27	6 C -122	7 N >0	8 O -141	9 F -328	10 Ne >0
11 Na -53	12 Mg >0	13 Al -43	14 Si -134	15 P -72	16 S -200	17 Cl -349	18 Ar >0
19 K -48	20 Ca -4	31 Ga -30	32 Ge -119	33 As -78	34 Se -195	35 Br -325	36 Kr >0
37 Rb -47	38 Sr -11	49 In -30	50 Sn -107	51 Sb -103	52 Te -190	53 I -295	54 Xe >0

FIGURE 6.24

Electron affinities (in kJ/mol) of the first five periods of the representative elements. Electron affinities are negative numbers because energy is released.

other elements. Considering electron configuration, it is easy to see why. The outer configuration of all halogens is ns^2np^5. The addition of one more electron gives the halide ions the same electron configuration as a noble gas, which we have seen is particularly stable.

Period and group trends for electron affinities are not nearly as regular as for ionization energy. In general, electron affinities increase (become more negative) from left to right across a period and decrease (become less negative) from top to bottom down a group. However, there are many exceptions, owing in part to inherent difficulties in accurately measuring electron affinities.

Ionic Radius

Figure 6.25 shows the comparison of ion sizes to atom sizes for Groups 1, 2, 13, 16 and 17. The atoms are shown in gray. Groups 1, 2, and 13 are metals and form cations, shown in red. Groups 16 and 17 are nonmetals and form anions, shown in blue.

The removal of electrons always results in a cation that is considerably smaller than the parent atom. When the valence electron(s) are removed, the resulting ion has one fewer occupied principal energy level, so the electron cloud that remains is smaller. Another reason is that the remaining electrons are drawn closer to the nucleus because the protons now outnumber the electrons.

The addition of electrons always results in an anion that is larger than the parent atom. When the electrons outnumber the protons, the overall attractive force that the protons have for the electrons is decreased. The electron cloud also spreads out because more electrons results in greater electron-electron repulsions.

Trends

Period and group trends for ion radii are similar to the trends for atomic radii and for the same reasons. Left to right across the second period, the cations decrease in size because of greater nuclear charge. Starting in Group 15, a nitrogen atom becomes more stable by gaining three electrons to become a nitride ion, N^{3-}, and acquire a noble gas electron configuration. The nitride ion is larger than the previous cations, but the anions then decrease in size in

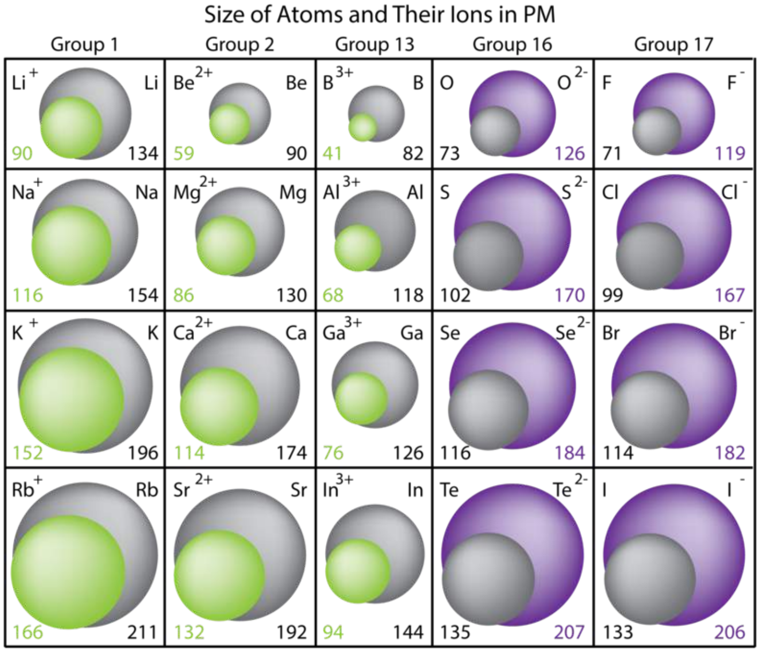


FIGURE 6.25

Atomic and ionic radii of the elements in Periods 1-5 and Groups 1,2, 13, 16, and 17: Atoms are shown in gray. The most common ion is shown in either red (for cations) or blue (for anions).

Groups 16 and 17. Both types of ions increase in size from top to bottom within a group due to an increase in the number of occupied principal energy levels.

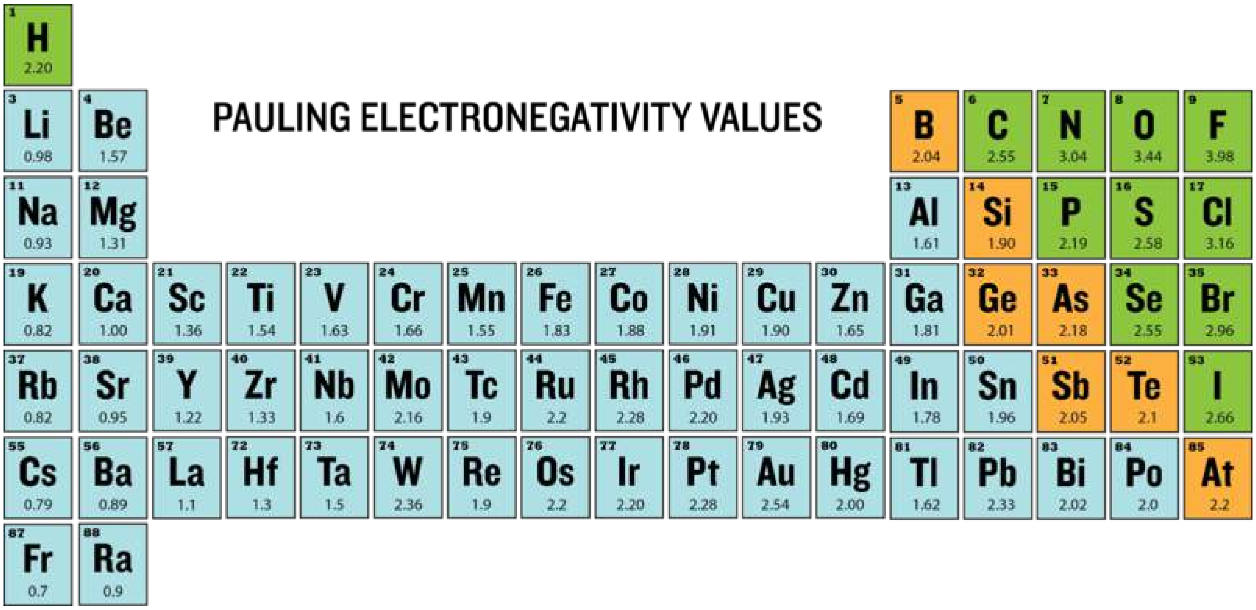
Electronegativity

Valence electrons of both atoms are always involved when those two atoms come together to form a chemical bond. Chemical bonds are the basis for how elements combine with one another to form compounds. When these chemical bonds form, atoms of some elements have a greater ability to attract the valence electrons involved in the bond than other elements. **Electronegativity** *is a measure of the ability of an atom to attract the electrons when the atom is part of a compound.* Electronegativity differs from electron affinity because electron affinity is the actual energy released when an atom gains an electron. Electronegativity is not measured in energy units, but is rather a relative scale. All elements are compared to one another, with the most electronegative element, fluorine, being assigned an electronegativity value of 3.98. Fluorine attracts electrons better than any other element. **Figure** 6.26 shows electronegativity values for all elements.

Since metals have few valence electrons, they tend to increase their stability by losing electrons to become cations. Consequently, the electronegativities of metals are generally low. Nonmetals have more valence electrons and increase theirs stability by gaining electrons to become anions. The electronegativities of nonmetals are generally high.

Trends

Electronegativities generally increase from left to right across a period. This is due to an increase in nuclear charge. Alkali metals have the lowest electronegativities, while halogens have the highest. Because most noble gases do not form compounds, they do not have electronegativities. Note that there is little variation among the transition metals. Electronegativities generally decrease from top to bottom within a group due to the larger atomic size.



PAULING ELECTRONEGATIVITY VALUES

1 H 2.20																
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
55 Cs 0.79	56 Ba 0.89	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.20	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Tl 1.62	82 Pb 2.33	83 Bi 2.02	84 Po 2.0	85 At 2.2
87 Fr 0.7	88 Ra 0.9															

FIGURE 6.26

The electronegativity scale was developed by Nobel Prize winning American chemist Linus Pauling. The largest electronegativity (3.98) is assigned to fluorine and all other electronegativities measurements are on a relative scale.

Metallic and Nonmetallic Character

Metallic character refers to the level of reactivity of a metal. Metals tend to lose electrons in chemical reactions, as indicated by their low ionization energies. Within a compound, metal atoms have relatively low attraction for electrons, as indicated by their low ionization energies. By following the trend summary in **Figure** 6.27, you can see that the most reactive metals would reside in the lower left portion of the periodic table. The most reactive metal is cesium, which is not found in nature as a free element. It reacts explosively with water and will ignite spontaneously in air. Francium is below cesium in the alkali metal group, but is so rare that most of its properties have never been observed.

Nonmetals tend to gain electrons in chemical reactions and have a high attraction for electrons within a compound. The most reactive nonmetals reside in the upper right portion of the periodic table. Since the noble gases are a special group because of their lack of reactivity, the element fluorine is the most reactive nonmetal. It is not found in nature as a free element. Fluorine gas reacts explosively with many other elements and compounds and is considered to be one of the most dangerous known substances.

Look at the reactivity of metals in the form of Sumo Wrestlers at http://freezeray.com/flashFiles/ReactivitySumo.htm !

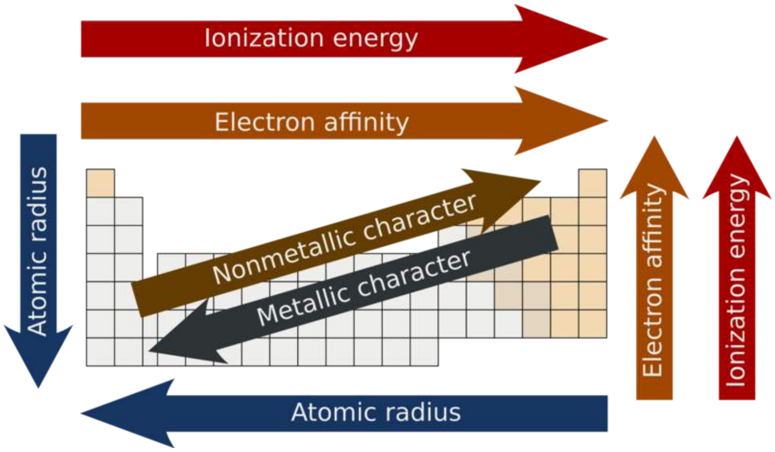


FIGURE 6.27
Summary of periodic trends within periods and groups.

Lesson Summary

- Atomic radius generally decreases from left to right across a period and increases from top to bottom within a group.
- The number of valence electrons varies from one to eight among the groups of representative elements.
- Metals tend to lose electrons easily to form positively-charged cations, while nonmetals tend to gain electrons to form negatively-charged anions.
- Ionization energy is the energy required to remove an electron. It generally increases from left to right across a period and decreases from top to bottom within a group.
- Electron affinity is the energy released when an atom gains an electron and is highest for the halogen group.
- Cations are smaller than their parent atom, while anions are larger than their parent atom.
- Electronegativity is a measure of the attraction for electrons. It generally increases from left to right across a period and decreases from top to bottom within a group.
- Metallic reactivity is greatest for elements in the lower left area of the periodic table, while nonmetallic reactivity is greatest for elements in the upper right area.

Lesson Review Questions

Reviewing Concepts

1. Answer the following:
    a. How does atomic radius change from left to right across a period? Explain.
    b. How does atomic radius change from top to bottom within a group? Explain.
2. Answer the following:
    a. How does ionization energy change from left to right across a period? Explain.
    b. How does ionization energy change from top to bottom within a group? Explain.
3. Why is the second ionization energy of an element always larger than the first?
4. Why are most electron affinities negative numbers?
5. Answer the following:
    a. Which group of elements has the highest electronegativities?

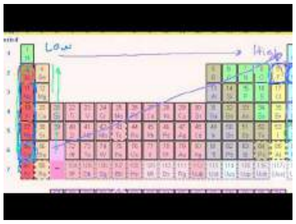
   b. Which group of elements has the lowest electronegativities?

Problems

6. Which element in each pair below has the larger atomic radius?
   a. K, Na
   b. S, Cl
   c. F, Br
   d. Ba, Cs
7. Which equation below shows the second ionization of an alkaline earth metal?
   a. $\text{Sr} \rightarrow \text{Sr}^+ + e^-$
   b. $\text{Sr}^+ \rightarrow \text{Sr}^{2+} + e^-$
   c. $\text{Cs} \rightarrow \text{Cs}^+ + e^-$
   d. $\text{Cs}^+ \rightarrow \text{Cs}^{2+} + e^-$
8. Which element in each pair has the higher first ionization energy?
   a. Mg, P
   b. Se, O
   c. Li, Rb
   d. Ne, N
9. Why does the third ionization energy (IE_3) of calcium show an unusually large jump in value compared to the second ionization energy?
10. Which atom/ion of each pair is larger?
    a. Na, Na^+
    b. Br, Br^-
    c. Be^{2+}, Ca^{2+}
    d. F^-, O^{2-}
11. Which element in each pair has a higher electronegativity value?
    a. F, Cl
    b. P, S
    c. Sr, Be
    d. Al, Na
12. Which equation below represents the electron affinity of a halogen atom?
    a. $\text{Cl} \rightarrow \text{Cl}^+ + e^-$
    b. $\text{S} + e^- \rightarrow \text{S}^-$
    c. $\text{Br} + e^- \rightarrow \text{Br}^-$
    d. $\text{Na} \rightarrow \text{Na}^+ + e^-$
13. Arrange the elements in order of increasing metallic character: K, Al, Cs, Na.
14. Arrange the elements in order of increasing nonmetallic character: O, P, F, N.
15. Explain why the Na^{2+} ion is very unlikely to form.

Further Reading / Supplemental Links

- Watch a video lecture on Other Periodic Table Trends at http://www.youtube.com/watch?v=XMLd-O6PgVs (15:39).



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CHAPTER 7 The Process of Radioactive Decay

Chapter Outline

NC Standards addressed in this chapter

Chm.1.1.4 Explain the process of radioactive decay using nuclear equations and half-life.

Chm.1.1.4 • Use the symbols for and distinguish between alpha (24He), and beta (-10e) nuclear particles, and gamma (γ) radiation include relative mass).

• Use shorthand notation of particles involved in nuclear equations to balance and solve for unknowns.

• Compare the penetrating ability of alpha, beta, and gamma radiation.

• Conceptually describe nuclear decay, including:

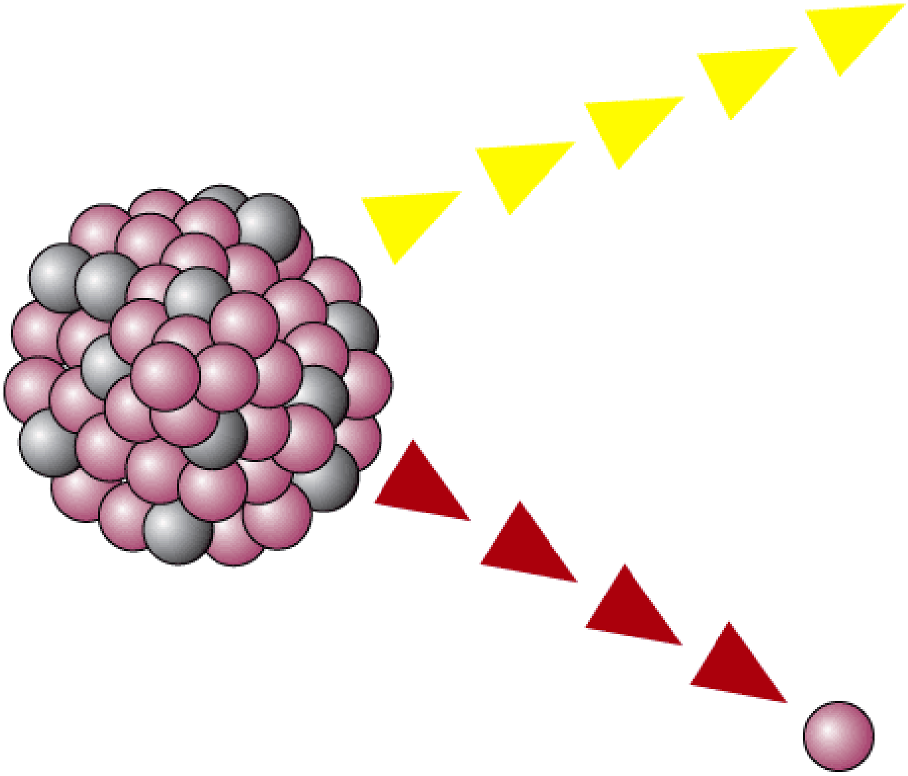
1. Decay as a random event, independent of other energy influences

2. Using symbols to represent simple balanced decay equations

3. Half-life (including simple calculations) • Compare radioactive decay with fission and fusion.

7.1 Radioactive Decay Introduction

- Define radioactive decay.
- Describe three types of radioactive decay.
- Explain how radioactive decay changes one element to another.
- Outline the dangers of radioactive decay.



This diagram could almost illustrate a pinball game. It actually illustrates the process of radioactive decay. This is a process that occurs all around you. It may even occur inside your body.

Q: Based on the diagram, what happens during radioactive decay?

A: The nucleus of a radioactive atom gives off radiation in the form of energy and a particle of matter.

Introducing Radioactive Decay

Radioactive decay is the process in which the nuclei of radioactive atoms emit charged particles and energy, which are called by the general term *radiation*. Radioactive atoms have unstable nuclei, and when the nuclei emit radiation, they become more stable. Radioactive decay is a nuclear—rather than chemical—reaction because it involves only the nuclei of atoms. In a nuclear reaction, one element may change into another. For a fun introduction to radioactive decay, watch the short video at this URL: http://www.youtube.com/watch?v=o-9yt7OAYmE

Types of Radioactive Decay

There are several types of radioactive decay, including alpha, beta, and gamma decay. In all three types, nuclei emit radiation, but the nature of the radiation differs. The **Table** 7.1 shows the radiation emitted in each type of decay. You can watch a video about the three types at this URL: http://www.youtube.com/watch?v=3koOwozY4oc .

TABLE 7.1: **Types of Radioactive Decay**

Type	Radiation Emitted
Alpha decay	alpha particle (2 protons and 2 neutrons) + energy
Beta decay	beta particle (1 electron or 1 positron) + energy
Gamma decay	energy (gamma ray)

It's Elemental

Both alpha and beta decay change the number of protons in an atom's nucleus, thereby changing the atom to a different element. In alpha decay, the nucleus loses two protons. In beta decay, the nucleus either loses a proton or gains a proton. In gamma decay, no change in proton number occurs, so the atom does not become a different element.

Q: If the radioactive element polonium (Po) undergoes alpha decay, what element does it become?

A: From the periodic table, the atomic number of polonium is 84, so it has 84 protons. If it loses two protons through alpha decay, it will have 82 protons. Atoms with 82 protons are the element lead (Pb).

Dangers of Radioactive Decay

The charged particles and energy emitted during radioactive decay can harm living things, but the three types of radioactive decay aren't equally dangerous. That's because they differ in how far they can travel and what they can penetrate. You can see this in the **Figure** 7.1 and in the animation at the following URL: http://teachnuclear.ca/contents/cna_radiation/gamma_rays/

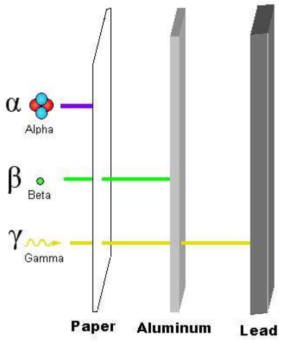


FIGURE 7.1

Alpha particles can travel only a few centimeters through air. They can burn the skin but not penetrate it. Beta particles can travel up to a meter through air. They can penetrate and damage skin. Gamma rays can travel thousands of meters through air. They can penetrate and damage cells deep inside the body.

Summary

- Radioactive decay is the process in which unstable nuclei of radioactive atoms become stable by emitting charged particles and energy.

- There are three types of radioactive decay: alpha decay, beta decay, and gamma decay. Alpha and beta decay change one element into another. Gamma decay does not.
- Radioactive decay can damage living things. Alpha decay is the least damaging, and gamma decay is the most damaging.

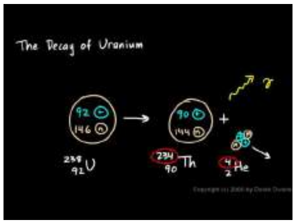
Vocabulary

- **radioactive decay**: Process in which the unstable nucleus of a radioisotope becomes stable by emitting charged particles and energy and changing to another element.

Practice

Watch the video about the decay of uranium at the following URL, and then answer the questions below.

http://www.youtube.com/watch?v=F8sAkMQmOlU



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1. In the first step of the decay of uranium-238, what are the products?
2. How does the atomic number of uranium-238 change in the first step of the decay?
3. What happens to the new element that results from the first step of uranium-238 decay?
4. What type of decay is the decay of uranium-238?
5. What is the half-life of a radioactive element? What is the half-life of uranium-238?
6. Uranium-238 decays very slowly, so why is a lump of uranium-238 dangerous?

Review

1. What is radioactive decay?
2. Compare and contrast alpha, beta, and gamma decay.

7.2 Radioactive Decay

Lesson Objectives

- Identify three types of radioactive decay.
- Define the half-life of a radioisotope.
- Explain radioactive dating.

Lesson Vocabulary

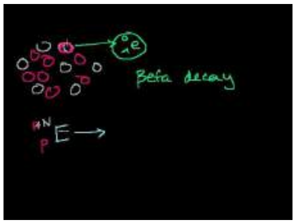
- half-life
- radioactive dating
- radioactive decay

Introduction

Unstable nuclei of radioisotopes become stable by emitting charged particles and energy. This process is called **radioactive decay**. It is a nuclear, as opposed to chemical, reaction, in which one element changes into another.

Types of Radioactive Decay

There are three types of radioactive decay: alpha, beta, and gamma decay. In all three types, nuclei emit radiation, but the nature of that radiation differs from one type of decay to another. You can watch a video about the three types at this URL: http://www.youtube.com/watch?v=3koOwozY4oc (17:02).



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Alpha Decay

Alpha decay occurs when an unstable nucleus emits an alpha particle and energy. The diagram in **Figure** 7.2 represents alpha decay. An alpha particle contains two protons and two neutrons, giving it a charge of +2. A helium nucleus has two protons and two neutrons, so an alpha particle is represented in nuclear equations by the symbol

$^{4}_{2}\text{He}$. The superscript 4 is the mass number (2 protons + 2 neutrons). The subscript 2 is the charge of the particle as well as the number of protons.

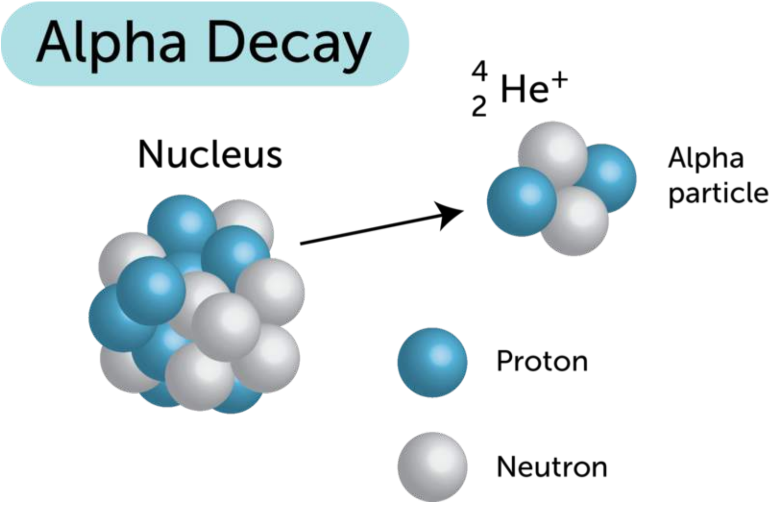


FIGURE 7.2

Alpha decay results in the loss of two protons and two neutrons from a nucleus.

An example of alpha decay is the decay of uranium-238 to thorium-234. In this reaction, uranium loses two protons and two neutrons to become the element thorium. The reaction can be represented by this equation:

$$^{238}_{92}\text{U} \rightarrow ^{234}_{90}\text{Th} + ^{4}_{2}\text{He} + \text{Energy}$$

If you count the number of protons and neutrons on each side of this equation, you'll see that the numbers are the same on both sides of the arrow. This means that the equation is balanced. The thorium-234 produced in this reaction is unstable, so it will undergo radioactive decay as well. The alpha particle ($^{4}_{2}\text{He}$) produced in the reaction can pick up two electrons to form the element helium. This is how most of Earth's helium formed.

Problem Solving

Problem: Fill in the missing subscript and superscript to balance this nuclear equation: $^{208}_{84}\text{Po} \rightarrow ^{?}_{?}\text{Pb} + ^{4}_{2}\text{He} + \text{Energy}$

Solution: The subscript is 82, and the superscript is 204.

You Try It!

Problem: Fill in the missing subscript and superscript to balance this nuclear equation: $^{222}_{?}\text{Ra} \rightarrow ^{?}_{86}\text{Rn} + ^{4}_{2}\text{He} + \text{Energy}$

Beta Decay

Beta decay occurs when an unstable nucleus emits a beta particle and energy. A beta particle is an electron. It has a charge of -1. In nuclear equations, a beta particle is represented by the symbol $^{0}_{-1}e$. The subscript -1 represents the particle's charge, and the superscript 0 shows that the particle has virtually no mass. Nuclei contain only protons and neutrons, so how can a nucleus emit an electron? A neutron first breaks down into a proton and an electron (see **Figure** 7.3). Then the electron is emitted from the nucleus, while the proton stays inside the nucleus. The proton increases the atomic number by one, thus changing one element into another.

An example of beta decay is the decay of thorium-234 to protactinium-234. In this reaction, thorium loses a neutron and gains a proton to become protactinium. The reaction can be represented by this equation:

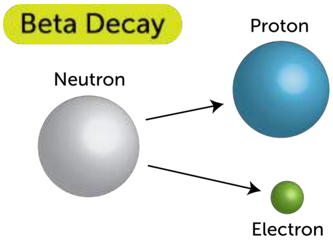


FIGURE 7.3
In beta decay, an electron and a proton form from a neutron (another unusual particle, called an antineutrino, is also produced). Only the electron is emitted from the nucleus. How does this change the atomic number and atomic mass of the atom?

$$^{234}_{90}\mathrm{Th} \rightarrow ^{234}_{91}\mathrm{Pa} + ^{0}_{-1}e + \mathrm{Energy}$$

The protactinium-234 produced in this reaction is radioactive and decays to another element. The electron produced in the reaction (plus another electron) can combine with an alpha particle to form helium.

Problem Solving

Problem: Fill in the missing subscript and superscript in this nuclear equation: $^{131}_{53}\mathrm{I} \rightarrow ^{?}_{?}\mathrm{Xe} + ^{0}_{-1}e + \mathrm{Energy}$

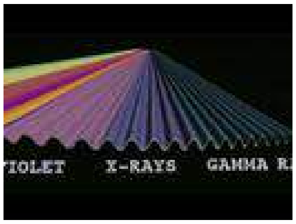
Solution: The subscript is 54, and the superscript is 131.

You Try It!

Problem: Fill in the missing subscript and superscript in this nuclear equation: $^{14}_{?}\mathrm{C} \rightarrow ^{?}_{7}\mathrm{N} + ^{0}_{-1}e + \mathrm{Energy}$

Gamma Decay

In alpha and beta decay, both particles and energy are emitted. In gamma decay, only energy is emitted. Gamma decay occurs when an unstable nucleus gives off gamma rays. Gamma rays, like rays of visible light and X-rays, are waves of energy that travel through space at the speed of light. Gamma rays have the greatest amount of energy of all such waves. By itself, gamma decay doesn't cause one element to change into another, but it is released in nuclear reactions that do. Some of the energy released in alpha and beta decay is in the form of gamma rays. You can learn more about gamma radiation at this URL: http://www.youtube.com/watch?v=okyynBaSOtA (2:45).



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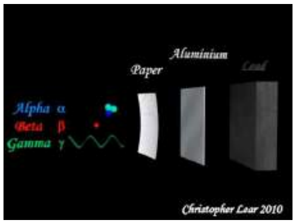
Comparing Types of Radiation

The different types of radiation vary in how far they are able to travel and what they can penetrate (see **Figure** 7.4 and the URL below).

http://www.youtube.com/watch?v=ec8iomUS34U (0:30)



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- Alpha particles can travel only a few centimeters through air. They cannot pass through a sheet of paper or thin layer of clothing. They may burn the skin but cannot penetrate tissues beneath the skin.
- Beta particles can travel up to a meter through air. They can pass through paper and cloth but not through a sheet of aluminum. They can penetrate and damage tissues beneath the skin.
- Gamma rays can travel thousands of meters through air. They can pass through a sheet of aluminum as well as paper and cloth. They can be stopped only by several centimeters of lead or several meters of concrete. They can penetrate and damage organs deep inside the body.

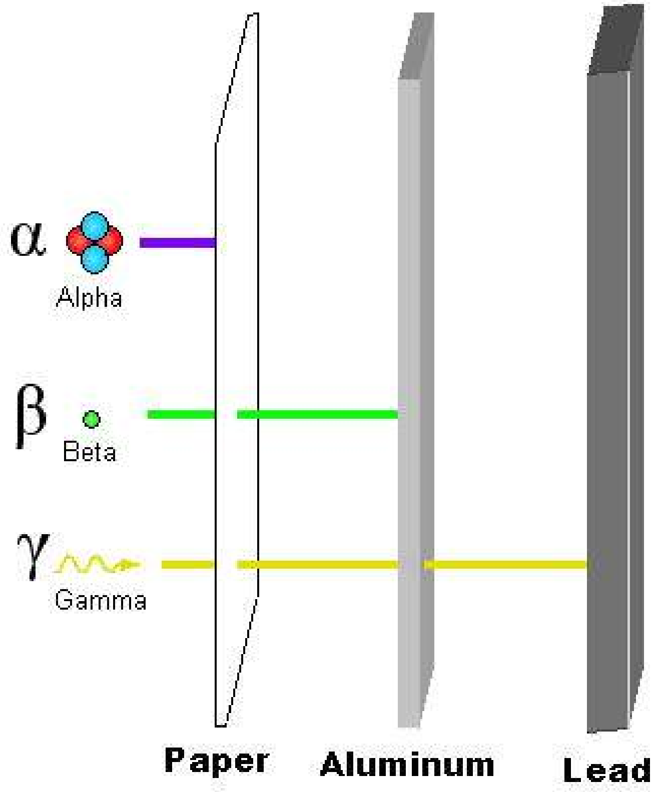


FIGURE 7.4

It's easy to stop alpha particles and even beta particles. However, it's very difficult to stop gamma rays.

Rate of Radioactive Decay

A radioactive isotope decays at a certain constant rate. The rate is measured in a unit called the **half-life**. This is the length of time it takes for half of a given amount of the isotope to decay. The concept of half-life is illustrated in **Figure** 7.5 for the beta decay of phosphorus-32 to sulfur-32. The half-life of this radioisotope is 14 days. After

14 days, half of the original amount of phosphorus-32 has decayed. After another 14 days, half of the remaining amount (or one-quarter of the original amount) has decayed, and so on.

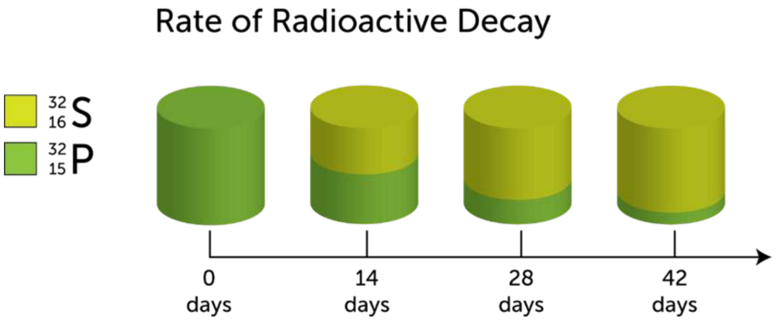


FIGURE 7.5

This diagram models the rate of decay of phosphorus-32 to sulfur-32.

Different radioactive isotopes vary greatly in their rate of decay. As you can see from the examples in **Table** 7.2, the half-life of a radioisotope can be as short as a split second or as long as several billion years. You can simulate radioactive decay of radioisotopes with different half-lives at the URL below.

http://www.colorado.edu/physics/2000/isotopes/radioactive_decay3.html

Some radioisotopes decay much more quickly than others.

TABLE 7.2: **Half-Life of Some Radioisotopes**

Isotope	Half-life
Uranium-238	4.47 billion years
Potassium-40	1.28 billion years
Carbon-14	5,730 years
Hydrogen-3	12.3 years
Radon-222	3.82 days
Polonium-214	0.00016 seconds

Problem Solving

Problem: If you had a gram of carbon-14, how many years would it take for radioactive decay to reduce it to one-quarter of a gram?

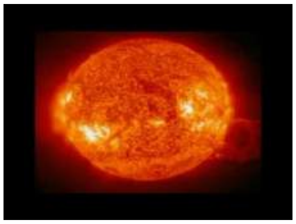
Solution: One gram would decay to one-quarter of a gram in 2 half-lives $\left(\frac{1}{2}\times\frac{1}{2}=\frac{1}{4}\right)$, or 2 × 5,730 years = 11,460 years.

You Try It!

Problem: What fraction of a given amount of hydrogen-3 would be left after 36.9 years of decay?

Radioactive Dating

Radioactive isotopes can be used to estimate the ages of fossils and rocks. The method is called **radioactive dating**. Carbon-14 dating is an example of radioactive dating. It is illustrated in the video at this URL: http://www.youtube.com/watch?v=udkQwW6aLik (11:00).



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How Carbon-14 Dating Works

Carbon-14 forms naturally in Earth's atmosphere when cosmic rays strike atoms of nitrogen-14. Living things take in and use carbon-14, just as they do carbon-12. The carbon-14 in living things gradually decays to nitrogen-14. However, it is constantly replaced because living things keep taking in carbon-14. As a result, there is a fixed ratio of carbon-14 to carbon-12 in organisms as long as they are alive. This is illustrated in the top part of **Figure** 7.6.

After organisms die, the carbon-14 they already contain continues to decay, but it is no longer replaced (see bottom part of **Figure** 7.6). Therefore, the carbon-14 in a dead organism constantly declines at a fixed rate equal to the half-life of carbon-14. Half of the remaining carbon-14 decays every 5,730 years. If you measure how much carbon-14 is left in a fossil, you can determine how many half-lives (and how many years) have passed since the organism died.

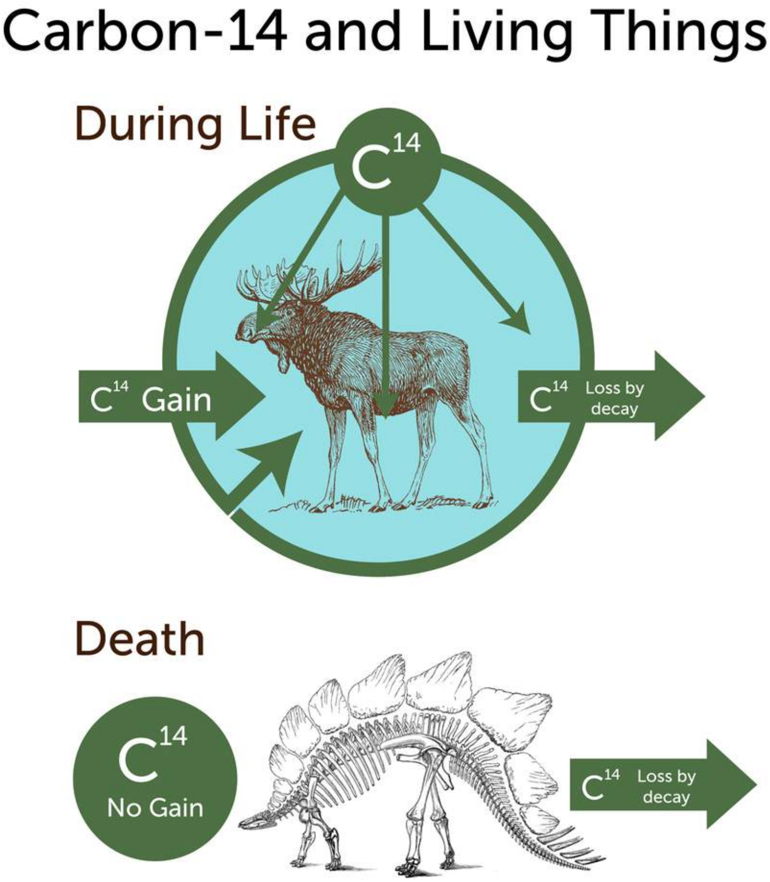


FIGURE 7.6

After organisms die, the carbon-14 they contain is lost at a constant rate.

Limit on Carbon-14 Dating

Carbon-14 has a relatively short half-life (see **Table** 7.2). After about 50,000 years, too little carbon-14 is left in a fossil to be measured. Therefore, carbon-14 dating can only be used to date fossils that are less than 50,000 years old. Radioisotopes with a longer half-life, such as potassium-40, must be used to date older fossils and rocks.

Lesson Summary

- Radioactive decay is the process in which unstable nuclei become stable by emitting particles and energy and changing to different elements. There are three types of radioactive decay: alpha, beta, and gamma decay. Each differs in what is emitted, how far it can travel, and what it can penetrate.
- The rate of radioactive decay varies from one radioisotope to another. The rate is measured by the half-life. This is the length of time it takes for half of a given amount of a radioisotope to decay.
- Radioactive dating is the use of radioactive decay to estimate the ages of fossils and rocks. An example is carbon-14 dating.

Lesson Review Questions

Recall

1. What is radioactive decay?
2. Define half-life.
3. Which type of radioactive decay is represented by each of the following nuclear equations?
   a. ${}^{60}_{27}\text{Co} \rightarrow {}^{60}_{28}\text{Ni} + {}^{0}_{-1}e + \text{Energy}$
   b. ${}^{263}_{106}\text{Sg} \rightarrow {}^{259}_{104}\text{Rf} + {}^{4}_{2}\text{He} + \text{Energy}$

Apply Concepts

4. Complete the following nuclear equation. First add the missing subscript. Then use the periodic table in **Figure** above to identify the unknown element (X). ${}^{211}_{85}\text{At} \rightarrow {}^{207}_{?}X + {}^{4}_{2}\text{He} + \text{Energy}$

5. Create a model to demonstrate the concept of the half-life of a radioisotope.

6. Assume that a fossil contains one-eighth of the carbon-14 that was present in the organism when it was alive. How long has it been since the organism died?

Think Critically

7. Compare and contrast alpha, beta, and gamma decay.

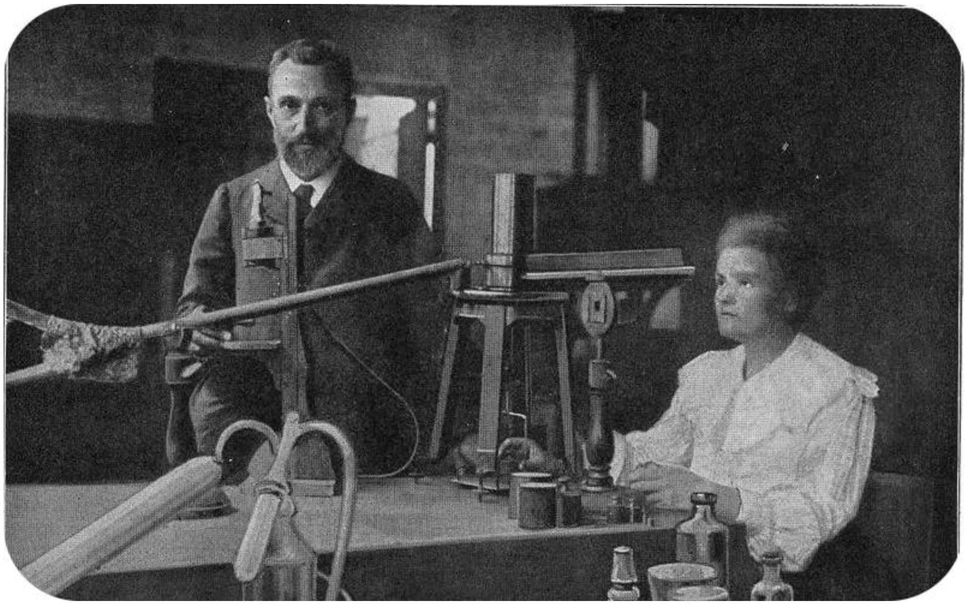
Points to Consider

Radioactive decay releases energy. Other types of nuclear reactions also release energy.

- What other types of reactions might the nucleus undergo?
- What might be the pros and cons of using nuclear reactions for energy?

7.3 Radioactive Decay as a Measure of Age

- Radioactive decay gives a way to determine the age of some types of rocks.



Why did this couple win the Nobel Prize?

Pierre and Marie Curie, a husband and wife team of physicists, discovered the spontaneous emission of particles from certain elements. They called this phenomenon "radioactivity." Together they won three Nobel prizes, and the element curium was named in their honor.

Radioactive Decay

Radioactivity is the tendency of certain atoms to decay into lighter atoms, a process that emits energy. Radioactivity also provides a way to find the absolute age of a rock. First, we need to know about radioactive decay.

Radioactive Isotopes

Some isotopes are radioactive; **radioactive isotopes** are unstable and spontaneously change by gaining or losing particles. Two types of radioactive decay are relevant to dating Earth materials (**Table** 7.3):

TABLE 7.3: **Types of Radioactive Decay**

Particle	Composition	Effect on Nucleus
Alpha	2 protons, 2 neutrons	The nucleus contains two fewer protons and two fewer neutrons.
Beta	1 electron	One neutron decays to form a proton and an electron. The electron is emitted.

The radioactive decay of a **parent isotope** (the original element) leads to the formation of stable **daughter product**, also known as daughter isotope. As time passes, the number of parent isotopes decreases and the number of daughter isotopes increases (**Figure** 7.7).

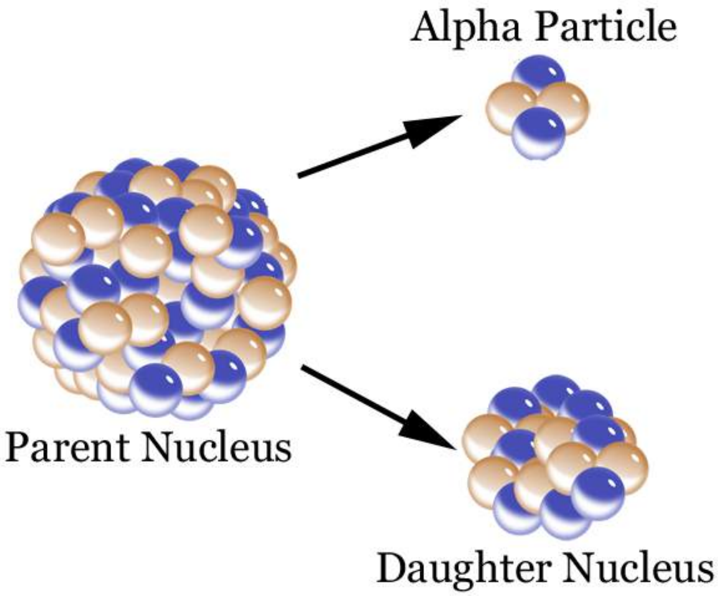


FIGURE 7.7
A parent emits an alpha particle to create a daughter.

An animation of radioactive decay: http://lectureonline.cl.msu.edu/~mmp/applist/decay/decay.htm .

Half-Lives

Radioactive materials decay at known rates, measured as a unit called **half-life**. The half-life of a radioactive substance is the amount of time it takes for half of the parent atoms to decay. This is how the material decays over time (see **Table** 7.4).

TABLE 7.4: **Radioactive Decay**

No. of half lives passed	Percent parent remaining	Percent daughter produced
0	100	0
1	50	50
2	25	75
3	12.5	87.5
4	6.25	93.75
5	3.125	96.875
6	1.563	98.437
7	0.781	99.219
8	0.391	99.609

Pretend you find a rock with 3.125% parent atoms and 96.875% daughter atoms. How many half lives have passed? If the half-life of the parent isotope is 1 year, then how old is the rock? The decay of radioactive materials can be shown with a graph (**Figure** 7.8).

An animation of half-life: http://einstein.byu.edu/~masong/htmstuff/Radioactive2.html .

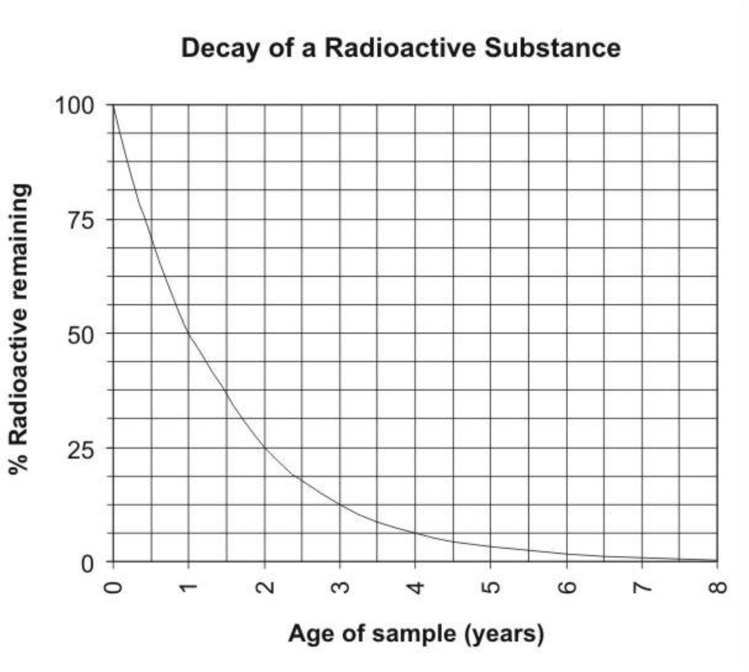
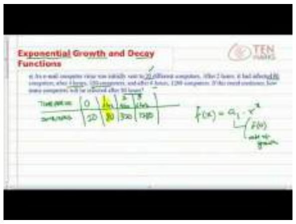


FIGURE 7.8

Decay of an imaginary radioactive substance with a half-life of one year.

Notice how it doesn't take too many half lives before there is very little parent remaining and most of the isotopes are daughter isotopes. This limits how many half lives can pass before a radioactive element is no longer useful for dating materials. Fortunately, different isotopes have very different half lives.

Radiometric decay is exponential. Learn how exponential growth and decay can be described mathematically in this video: http://www.youtube.com/watch?v=UbwMW7Q6F3E (4:46).



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Summary

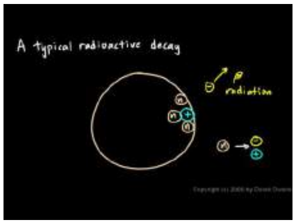
- A half life is the time it takes for half of the parent isotopes of an element to change to daughter isotopes.
- With alpha decay, the nucleus loses two protons and two neutrons; with beta decay only one electron is lost.
- Radiometric decay is exponential.

Practice

Use this resource to answer the questions that follow.

A Typical Radioactive Decay

http://www.youtube.com/watch?v=r-t01FsUXao



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1. Why are protons necessary near neutrons?
2. What does a neutron decay into?
3. What does the electron ejection cause?
4. What happens to the element when it gains a proton?
5. What is an alpha particle?
6. What is transmutation?
7. What is produced from transmutation?

Review

1. Describe the two types of radioactive decay that are relevant to dating earth materials.
2. For how many half lives is a set of parent and daughter isotopes useful as a system of dating?
3. What does it mean that radioactive decay is exponential?

7.4 References

1. User:Penubag/Wikimedia Commons. Radiation has different levels of penetration. Public Domain
2. Christopher Auyeung. CK-12 Foundation . CC BY-NC 3.0
3. Christopher Auyeung. CK-12 Foundation . CC BY-NC 3.0
4. User:Penubag/Wikimedia Commons. http://commons.wikimedia.org/wiki/File:RadiationPenetration2-pn.png . Public Domain
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CHAPTER 8 Ionic and Metallic Bonding

Chapter Outline

The red crystal above is called rhodochrosite. Most of the best specimens of this striking gemstone, including the one in the photo, have been found in central Colorado. What are the properties of crystals? How are the atoms or ions that make up crystals arranged? What is responsible for the beautiful colors of minerals and gems? This chapter will reveal the answers to these questions, as you learn about the composition and properties of ionic compounds and of pure metals. In an ionic crystalline solid, positive and negative ions are arranged in an orderly, repeating fashion. The colors are contributed by various amounts of transition metal ions in the crystals. The red color of rhodochrosite comes from manganese(II) carbonate, $MnCO_3$. Metals also have a crystalline structure, which contribute to the familiar properties of metals like malleability and electrical conductivity. We begin with a detailed look at how ions form.

Opening image courtesy of Sam Wise. CC−BY−NC−SA 2.0.

8.1 Ions

Lesson Objectives

- Be able to determine the number of valence electrons for any element and show with an electron dot diagram.
- Use the octet rule to predict the charges of the most common ion of representative elements
- Write electron configurations for ions.
- Identify what atoms or other ions are isoelectronic with a particular ion.
- Understand the particular stability that comes from transition metal ions that have either half-filled or completely filled *d* sublevels.

Lesson Vocabulary

- electron dot diagrams
- isoelectronic
- octet rule

Check Your Understanding

Recalling Prior Knowledge

- What is the Aufbau principle?
- What energy is associated with the removal of an electron from a neutral atom?
- What happens to this energy when a second or third electron is removed from a given atom?

As learned in previous chapters, ions are formed when atoms lose or gain electrons. Metal atoms have relatively few valence electrons, so when metals undergo chemical reactions, they tend to lose those valence electrons. Nonmetal atoms have more valence electrons than metals, so when nonmetals undergo reactions with metals, they tend to gain electrons. The chapter *Chemical Nomenclature* focused on the systems of naming and writing chemical formulas. In this lesson, we take a closer look at ions so that we can understand the physical and chemical properties of ionic compounds.

Electron Dot Diagrams

Recall that the valence electrons of an atom are the electrons that are in the highest occupied principal energy level. Valence electrons are primarily responsible for the chemical properties of elements. The number of valence electrons can be easily determined from the electron configuration. Several examples from the second period elements are shown in **Table** 8.1.

TABLE 8.1: **Electron Configuration**

Element	Electron Configuration	Number of Valence Electrons
lithium	$1s^2 2s^1$	1
beryllium	$1s^2 2s^2$	2
nitrogen	$1s^2 2s^2 2p^3$	5
neon	$1s^2 2s^2 2p^6$	8

In each case, valence electrons are those in the second principal energy level. As one proceeds left to right across a period, the number of valence electrons increases by one. In the *s* block, Group 1 elements have one valence electron, while Group 2 elements have two valence electrons. In the *p* block, the number of valence electrons is equal to the group number minus ten. Group 13 has three valence electrons, Group 14 has four, up through Group 18 with eight. The eight valence electrons, a full outer *s* and *p* sublevel, give the noble gases their special stability.

When examining chemical bonding, it is necessary to keep track of the valence electrons of each atom. **Electron dot diagrams** *are diagrams in which the valence electrons of an atom are shown as dots distributed around the element's symbol*. A beryllium atom, with two valence electrons, would have the electron dot diagram below.

Since electrons repel each other, the dots for a given atom are distributed evenly around the symbol before they are paired. **Table** 8.2 shows the electron dot diagrams for the entire second period.

TABLE 8.2: **Electron Dot Diagrams for the Second Period Elements**

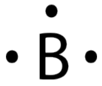
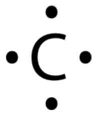
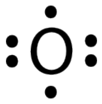
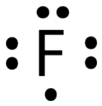
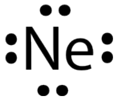
Group Number	Electron Dot Diagram
1	
2	
13	
14	
15	

TABLE 8.2: (continued)

Group Number	Electron Dot Diagram
16	
17	
18	

Electron dot diagrams would be the same for each element in the representative element groups. Most transition elements have two valence electrons, though some that have unusual electron configurations have only one.

Go to http://hyperphysics.phy-astr.gsu.edu/hbase/pertab/perlewis.html to answer the following question:

> From this periodic table, explain how valence electrons are added to the symbol from one column to the next.

The Octet Rule

The noble gases are unreactive because of their electron configurations. American chemist Gilbert Lewis (1875-1946) used this observation to explain the types of ions and molecules that are formed by other elements. He called his explanation the octet rule. The **octet rule** *states that atoms tend to form compounds in ways that give them eight valence electrons and thus the electron configuration of a noble gas*. An exception to an octet of electrons is in the case of the first noble gas, helium, which only has two valence electrons. This primarily affects the element hydrogen, which forms stable compounds by achieving two valence electrons.

There are two ways in which atoms can satisfy the octet rule. One way is by sharing their valence electrons with other atoms, which will be covered in the next chapter. The second way is by transferring valence electrons from one atom to another. Atoms of metals tend to lose all of their valence electrons, which leaves them with an octet from the next lowest principal energy level. Atoms of nonmetals tend to gain electrons in order to fill their outermost principal energy level with an octet.

Cations

As you have seen before, cations are the positive ions formed by the loss of one or more electrons. The most commonly formed cations of the representative elements are those that involve the loss of all of the valence electrons. Consider the alkali metal sodium (Na). It has one valence electron in the third principal energy level. Upon losing that electron, the sodium ion now has an octet of electrons from the second principal energy level. The equation below illustrates this process.

$$\begin{array}{lcl} \text{Na} & \rightarrow & \text{Na}^+ + e^- \\ 1s^22s^22p^63s^1 & & 1s^22s^22p^6 \text{ (octet)} \end{array}$$

The electron configuration of the sodium ion is now the same as that of the noble gas neon. The term **isoelectronic** *refers to an atom and an ion of a different atom, or two different ions, that have the same electron configuration*. The sodium ion is isoelectronic with the neon atom. Consider a similar process with magnesium and with aluminum:

$$\begin{array}{lcl} \text{Mg} & \rightarrow & \text{Mg}^{2+} + 2e^- \\ 1s^22s^22p^63s^2 & & 1s^22s^22p^6 \text{ (octet)} \end{array}$$

$$\begin{array}{lcl} \text{Al} & \rightarrow & \text{Al}^{3+} + 3e^- \\ 1s^22s^22p^63s^23p^1 & & 1s^22s^22p^6 \text{ (octet)} \end{array}$$

In this case, the magnesium atom loses its two valence electrons in order to achieve the same noble-gas configuration. The aluminum atom loses its three valence electrons. The Mg^{2+} ion, the Al^{3+} ion, the Na^+ ion, and the Ne atom are all isoelectronic. For representative elements under typical conditions, three electrons is that maximum number that will be lost.

We can also show the loss of valence electron(s) with an electron dot diagram.

$$\text{Na}\cdot \rightarrow \text{Na}^+ + e^-$$

Anions

Anions are the negative ions formed from the gain of one or more electrons. When nonmetal atoms gain electrons, they often do so until their outermost principal energy level achieves an octet. This process is illustrated below for the elements fluorine, oxygen, and nitrogen.

$$\begin{array}{lcl} \text{F} + e^- & \rightarrow & \text{F}^- \\ 1s^22s^22^p5 & & 1s^22s^22p^6 \text{ (octet)} \end{array}$$

$$\begin{array}{lcl} \text{O} + 2e^- & \rightarrow & \text{O}^{2-} \\ 1s^22s^22p^4 & & 1s^22s^22p^6 \text{ (octet)} \end{array}$$

$$\begin{array}{lcl} \text{N} + 3e^- & \rightarrow & \text{N}^{3-} \\ 1s^22s^22p^3 & & 1s^22s^22p^6 \text{ (octet)} \end{array}$$

All of these anions are isoelectronic with each other and with neon. They are also isoelectronic with the three cations from the previous section. Under typical conditions, three electrons is the maximum that will be gained in the formation of anions.

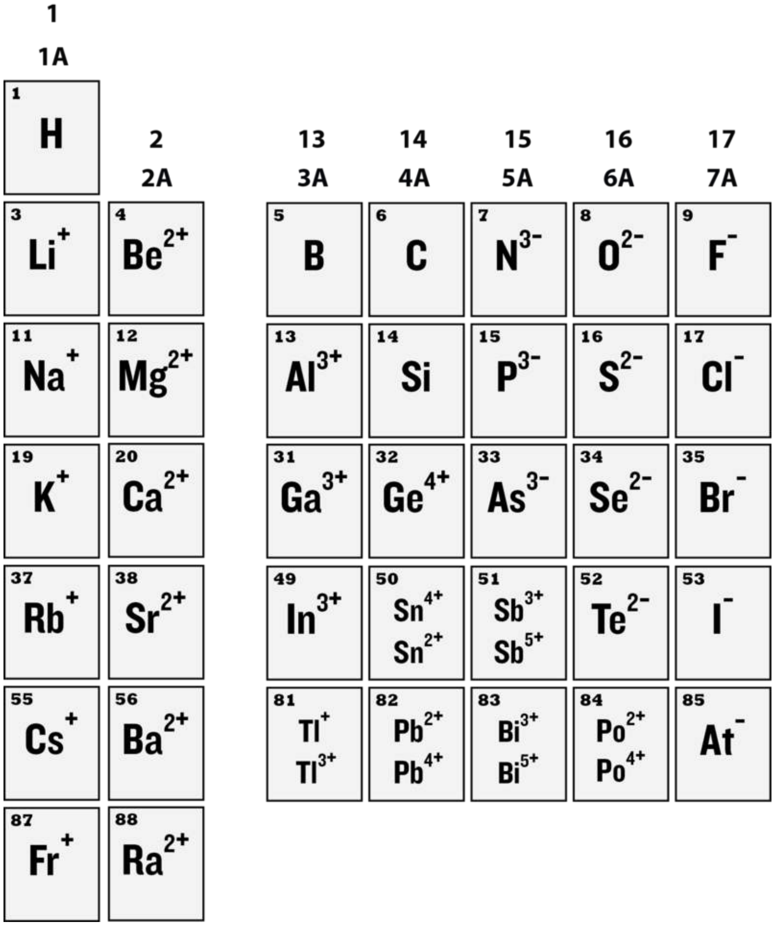


FIGURE 8.1

Outer electron configurations are constant within a group, so this pattern of ion formation repeats itself for Periods 3, 4, etc. (**Figure** 8.1).

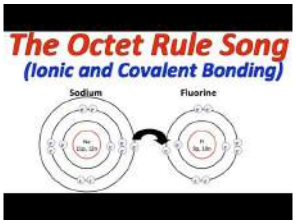
It is important not to misinterpret the concept of being isoelectronic. A sodium ion is very different from a neon atom (**Figure** 8.2) because the nuclei of the two contain different numbers of protons. One is an essential ion that is a part of table salt, while the other is an unreactive gas that is a very small part of the atmosphere. Likewise, sodium ions are very different than magnesium ions, fluoride ions, and all the other members of this isoelectronic series (N^{3-}, O^{2-}, F^{-}, Ne, Na^{+}, Mg^{2+}, Al^{3+}).

FIGURE 8.2

Neon atoms and sodium ions are isoelectronic. Neon is a colorless and unreactive gas that glows a distinctive red-orange color in a gas discharge tube. Sodium ions are most commonly found in crystals of sodium chloride, ordinary table salt.

You can go to http://web.jjay.cuny.edu/~acarpi/NSC/3-atoms.htm to see animations of atoms and ions.

Learning the octet rule can be fun! Watch this music video about the octet rule: http://www.youtube.com/watch?v=WzWk-mx_14E (6:30)



MEDIA

Click image to the left for more content.

1. How does this song compare an outer energy level with 8 electrons to emotions? 2. What are the two exceptions to the octet rule in this song?

Transition Metal Ions

Transition metals belong to the *d* block, meaning that the *d* sublevel of electrons is in the process of being filled with up to ten electrons. Many transition metals cannot lose enough electrons to attain a noble-gas electron configuration. In addition, you have learned that the majority of transition metals are capable of adopting ions with different charges. Iron, which forms either the Fe^{2+} or Fe^{3+} ions, loses electrons as shown below.

$$\begin{array}{ccc} \mathrm{Fe} & \rightarrow & \mathrm{Fe}^{2+} + 2e^- \\ [\mathrm{Ar}]3d^6 4s^2 & & [\mathrm{Ar}]3d^6 \end{array}$$

$$\begin{array}{ccc} \mathrm{Fe} & \rightarrow & \mathrm{Fe}^{3+} + 3e^- \\ [\mathrm{Ar}]3d^6 4s^2 & & [\mathrm{Ar}]3d^5 \end{array}$$

According to the Aufbau process, the electrons fill the 4*s* sublevel before beginning to fill the 3*d* sublevel. However, the outermost *s* electrons are always the first to be removed in the process of forming transition metal cations. Because most transition metals have two valence electrons, the charge of 2+ is a very common one for their ions. This is the case for iron above. A half-filled *d* sublevel (d^5) is particularly stable, which is the result of an iron atom losing a third electron.

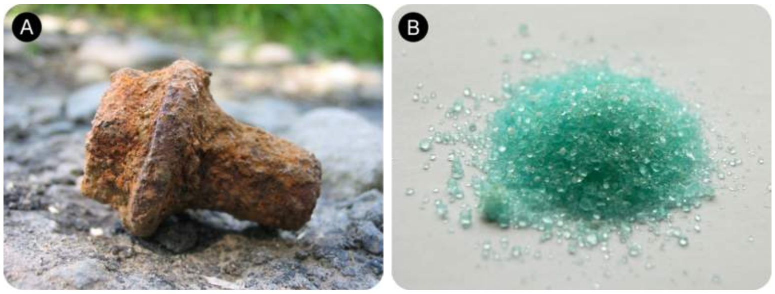


FIGURE 8.3

(A) Rust is a complex combination of oxides of iron(III), among them iron(III) oxide, Fe_2O_3. (B) Iron(II) sulfate, $FeSO_4$, has been known since ancient times as green vitriol and was used for centuries in the manufacture of inks.

Some transition metals that have relatively few *d* electrons may attain a noble-gas electron configuration. Scandium is an example.

$$\begin{array}{lcl} \mathrm{Sc} & \rightarrow & \mathrm{Sc}^{3+} + 3e^{-} \\ [\mathrm{Ar}]3d^{1}4s^{2} & & [\mathrm{Ar}] \end{array}$$

Others may attain configurations with a full *d* sublevel, such as zinc and copper.

$$\begin{array}{lcl} \mathrm{Zn} & \rightarrow & \mathrm{Zn}^{2+} + 2e^{-} \\ [\mathrm{Ar}]3d^{10}4s^{2} & & [\mathrm{Ar}]3d^{10} \end{array}$$

$$\begin{array}{lcl} \mathrm{Cu} & \rightarrow & \mathrm{Cu}^{+} + e^{-} \\ [\mathrm{Ar}]3d^{10}4s^{1} & & [\mathrm{Ar}]3d^{10} \end{array}$$

The resulting configuration above, with 18 electrons in the outermost principal energy level, is referred to as a pseudo noble-gas electron configuration. It gives particular stability to the Zn^{2+} and Cu^{+} ions.

Lesson Summary

- An electron dot diagram shows the chemical symbol of an element with dots that represent valence electrons evenly distributed around the symbol.
- The octet rule states that elements form chemical compounds so that each atom will acquire the electron configuration of a noble gas. In most instances, that is eight electrons except for helium which has only two electrons.
- Representative metals lose their valence electrons when forming ions, leaving them with a complete octet of the next-lowest energy level. Most nonmetals gain electrons when forming ions until their outer energy level as acquired an octet.
- Atoms and ions that have the same electron configuration are called isoelectronic. Common ions of representative elements are isoelectronic with a noble gas.
- In forming ions, transition metals lose their valence *s*-sublevel electrons before they lose their *d*-sublevel electrons. Half-filled or completely filled *d* sublevels give transition metal ions greater stability.

Lesson Review Questions

Reviewing Concepts

1. What is the maximum number of valence electrons that an atom can have?
2. State the number of protons and electrons in each of the following ions.
   a. K^{+}
   b. F^{-}
   c. P^{3-}
   d. Ti^{4+}
   e. Cd^{2+}

   f. Cr^{3+}

3. What is wrong with this statement? "When a chlorine atom gains an electron, it becomes an argon atom."
4. Why can the majority of transition metals form 2+ ions?
5. What is a pseudo noble-gas electron configuration?

Problems

6. How many electrons must each of the atoms below lose to achieve a noble gas electron configuration?
   a. Li
   b. Sr
   c. Al
   d. Ba
7. For the four elements in number 6, write the symbol of the most common ion and state which noble gas the ion is isoelectronic with.
8. How many electrons must each of the atoms below gain to achieve a noble gas electron configuration?
   a. Br
   b. S
   c. N
   d. I
9. For the four elements in number 8, write the symbol of the most common ion and state which noble gas the ion is isoelectronic with.
10. Write electron configurations for each of the following atoms. Then write the symbol for the most common ion each would form and the electron configuration of that ion.
    a. Be
    b. Cl
    c. Se
    d. Rb
11. Write electron configurations for the following ions.
    a. Cs^{+}
    b. Y^{3+}
    c. Ni^{2+}
    d. As^{3-}
    e. Te^{2-}
    f. Ag^{+}
    g. Pb^{4+}
    h. Mn^{2+}
12. For each ion in number 11, state whether each has (1) a noble-gas configuration, (2) a pseudo noble-gas configuration, or (3) neither.
13. Split the following ions into isoelectronic groups by noble gas: O^{2-}, Sr^{2+}, Ca^{2+}, H^{-}, V^{5+}, I^{-}, Ba^{2+}, Na^{+}, S^{2-}, Al^{3+}, La^{3+}, Li^{+}, As^{3-}.

Further Reading / Supplemental Links

- Ionic (Electrovalent) Bonding (http://www.chemguide.co.uk/atoms/bonding/ionic.html
- Bonding (http://www.chemteam.info/Bonding/Bonding.html
- Claude H. Yoder, *Ionic Compounds: Applications of Chemistry to Mineralogy*. Wiley-Interscience, 2006.

Points to Consider

Ionic compounds adopt the structure of an extended, three-dimensional lattice of alternating positive and negative ions held together by electrostatic attractive forces.

- How strong is an ionic crystal?
- Is an ionic crystal malleable or brittle? Why?
- Will ionic compounds conduct an electric current?

8.2 Ionic Bonds and Ionic Compounds

Lesson Objectives

- Explain how an ionic bond results from the transfer of one or more electrons from one atom to another and the resulting electrostatic attraction between the ions. Draw diagrams showing this process.
- Describe the structural arrangements of ions in a crystal, including coordination number and its relationship to a given compound's formula unit.
- Explain how various physical properties result from the ionic crystal lattice: strength, hardness, high melting points, brittleness, and electrical conductivity.

Lesson Vocabulary

- coordination number
- formula unit
- ionic bond
- ionic compound

Check Your Understanding

Recalling Prior Knowledge

- What types of charged particles attract each other and what types repel each other?
- How is an empirical formula different from a molecular formula?
- What is a crystal lattice?

Most of the rocks and minerals that make up the Earth's crust are composed of positive and negative ions held together by ionic bonding. An **ionic compound** *is an electrically neutral compound consisting of positive and negative ions*. You are very familiar with some ionic compounds such as sodium chloride (NaCl). A sodium chloride crystal consists of equal numbers of positive sodium ions (Na^+) and negative chloride ions (Cl^-).

Ionic Bonds

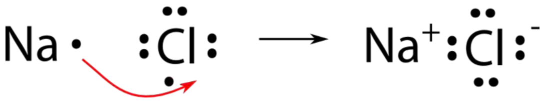
Oppositely charged particles attract each other. This attractive force is often referred to as an electrostatic force. An **ionic bond** *is the electrostatic force that holds ions together in an ionic compound*. The strength of the ionic bond is directly dependent upon the quantity of the charges and inversely dependent on the distance between the charged particles. A cation with a 2+ charge will make a stronger ionic bond than a cation with a 1+ charge. A larger ion makes a weaker ionic bond because of the greater distance between its electrons and the nucleus of the oppositely charged ion.

Watch an animation of ionic bonding at http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/IonicBonding.html .

1. How does this animation represent the transfer of electrons?
2. How do the sodium chloride units join together?

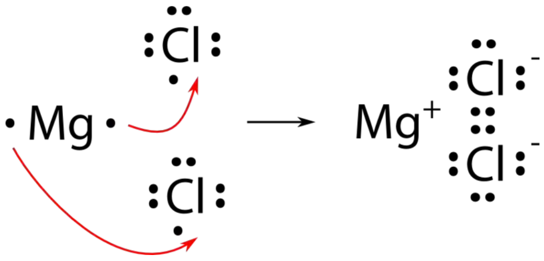
Electron Dot Diagrams

We will use sodium chloride as an example to demonstrate the nature of the ionic bond and how it forms. As you know, sodium is a metal and loses its one valence electron to become a cation. Chlorine is a nonmetal and gains one electron in becoming an anion. Both achieve a noble-gas electron configuration. However, electrons cannot be simply "lost" to nowhere in particular. A more accurate way to describe what is happening is that a single electron is transferred from the sodium atom to the chlorine atom as shown below.



The ionic bond is the attraction of the Na^+ ion for the Cl^- ion. It is conventional to show the cation without dots around the symbol to emphasize that the original energy level that contained the valence electron is now empty. The anion is now shown with a complete octet of electrons.

For a compound such as magnesium chloride, it is not quite as simple. Because magnesium has two valence electrons, it needs to lose both to achieve the noble-gas configuration. Therefore, two chlorine atoms will be needed.



The final formula for magnesium chloride is $MgCl_2$.

Formula Units

The **formula unit** *is the lowest whole number ratio of ions represented in an ionic compound*. The formula unit of sodium chloride is NaCl, while the formula unit of magnesium chloride is $MgCl_2$. The formula unit of an ionic compound is always an empirical formula. In a previous chapter, *Chemical Nomenclature*, you learned how to write correct formula units for ionic compounds by employing the crisscross method.

Ionic Compounds

The electron dot diagrams show the nature of the electron transfer that takes place between metal and nonmetal atoms. However, ionic compounds do not exist as discrete molecules, as the dot diagrams may suggest. In order to minimize the potential energy of the system, as nature prefers, ionic compounds take on the form of an extended three-dimensional array of alternating cations and anions. This maximizes the attractive forces between the oppositely charges ions. **Figure** 8.4 shows two different ways of representing the ionic crystal lattice. A ball and stick model makes it easier to see how individual ions are oriented with respect to one another. A space filling diagram is a more accurate representation of how the ions pack together in the crystal.

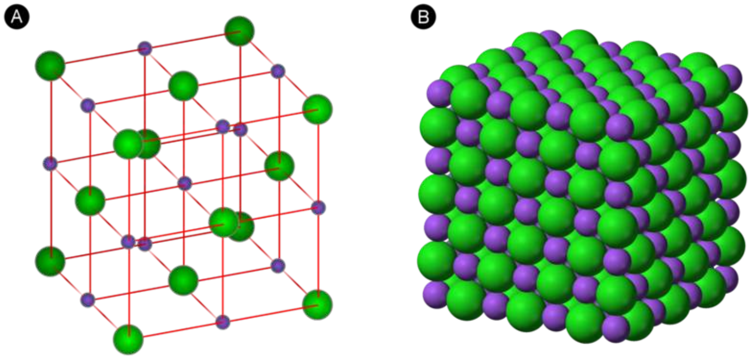


FIGURE 8.4

Two models of a sodium chloride crystal are shown. The purple spheres represent the Na^+ ions, while the green spheres represent the Cl^- ions. (A) In an expanded view, the distances between ions are exaggerated, more easily showing the coordination numbers of each ion. (B) In a space filling model, the electron clouds of the ions are in contact with each other.

Coordination Number

The **coordination number** *is the number of ions that immediately surround an ion of the opposite charge within a crystal lattice*. If you examine **Figure** 8.4 (A), you will see that there are six chloride ions immediately surrounding a single sodium ion. The coordination number of sodium is 6. Likewise, six sodium ions immediately surround each chloride ions, making the coordination number of chloride also equal to 6. Because the formula unit of sodium chloride displays a 1:1 ratio between the ions, the coordination numbers must be the same.

The formula unit for cesium chloride is CsCl, also a 1:1 ratio. However, as shown in **Figure** 8.5, the coordination numbers are not 6 as in NaCl. The center ion is the Cs^+ ion and is surrounded by the eight Cl^- ions at the corners of the cube. Each Cl^- ion is also surrounded by eight Cs^+ ions. The coordination numbers in this type of crystal are both 8. CsCl and NaCl do not adopt identical crystal packing arrangements because the Cs^+ ion is considerably larger than the Na^+ ion.

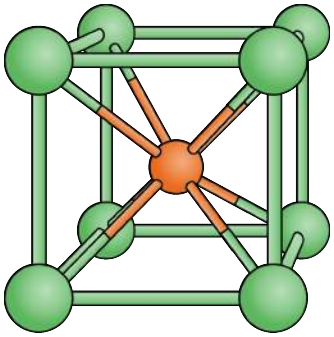


FIGURE 8.5

In a cesium chloride crystal, the cesium ion (orange) occupies the center, while the chloride ions (green) occupy each corner of the cube. The coordination number for both ions is 8.

Another type of crystal is illustrated by titanium(IV) oxide, TiO_2, which is commonly known as rutile. The rutile crystal is shown in **Figure** 8.6.

The gray Ti^{4+} ions are surrounded by six red O^{2-} ions. The O^{2-} ions are surrounded by three Ti^{4+} ions. The

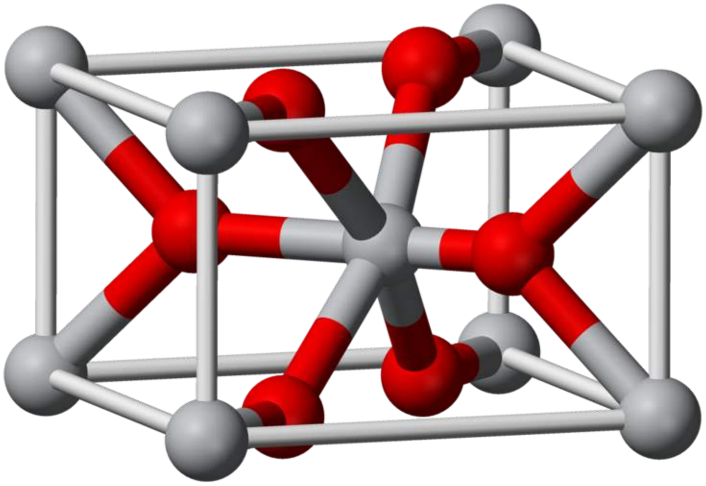


FIGURE 8.6

Titanium(IV) oxide forms tetragonal crystals. The coordination number of the Ti^{4+} ions (gray) is 6, while the coordination number of the O^{2-} ions (red) is 3.

coordination of the titanium(IV) cation is 6, which is twice the coordination number of the oxide anion, which is 3. This fits with the formula unit of TiO_2, since there are twice as many O^{2-} ions as Ti^{4+} ions.

The crystal structure of all ionic compounds must reflect the formula unit. In a crystal of iron(III) chloride, $FeCl_3$, there are three times as many chloride ions as iron(III) ions.

View an animated example at http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/IonicBonding.html

1. How does this animation represent the transfer of electrons?

2. How do the sodium chloride units join together?

To better understand the structure and properties of ionic compounds go to <http://www.chemguide.co.uk/atoms/structures/ionicst
http://www.chemguide.co.uk/atoms/structures/ionicstruct.html>

Physical Properties of Ionic Compounds

Figure 8.7 shows just a few examples of the color and brilliance of naturally occurring ionic crystals.

The regular and orderly arrangement of ions in the crystal lattice is responsible for the various shapes of these crystals, while transition metal ions give rise to the colors.

Because of the many simultaneous attractions between cations and anions that occur, ionic crystal lattices are very strong. The process of melting an ionic compound requires the addition of large amounts of energy in order to break all of the ionic bonds in the crystal. For example, sodium chloride has a melting temperature of about 800°C.

Ionic compounds are generally hard, but brittle. Why? It takes a large amount of mechanical force, such as striking a crystal with a hammer, to force one layer of ions to shift relative to its neighbor. However, when that happens, it brings ions of the same charge next to each other (**Figure** 8.8). The repulsive forces between like-charges ions causes the crystal to shatter. When an ionic crystal breaks, it tends to do so along smooth planes because of the regular arrangement of the ions.

Another characteristic property of ionic compounds is their electrical conductivity. **Figure** 8.9 shows three experiments in which two electrodes that are connected to a light bulb are placed in beakers containing three different substances.

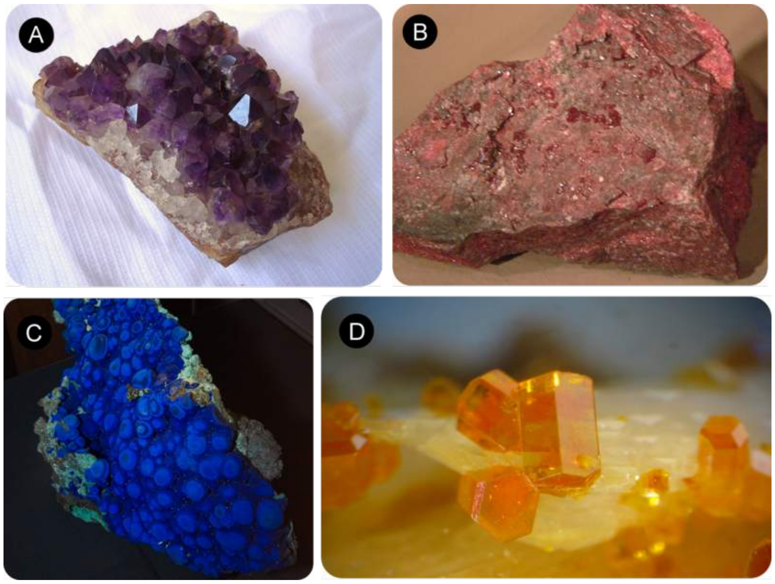


FIGURE 8.7

In nature, the ordered arrangement of ionic solids give rise to beautiful crystals. (A) Amethyst –a form of quartz, SiO_2, whose purple color comes from iron ions. (B) Cinnabar –the primary ore of mercury is mercury(II) sulfide, HgS. (C) Azurite –a copper mineral, $Cu_3(CO_3)_2(OH)_2$. (D) Vanadinite –the primary ore of vanadium, $Pb_5(VO_4)_3Cl$.

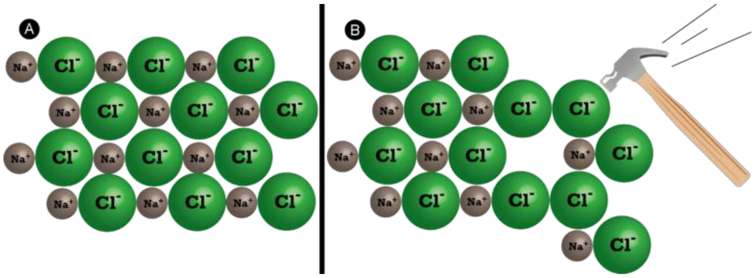
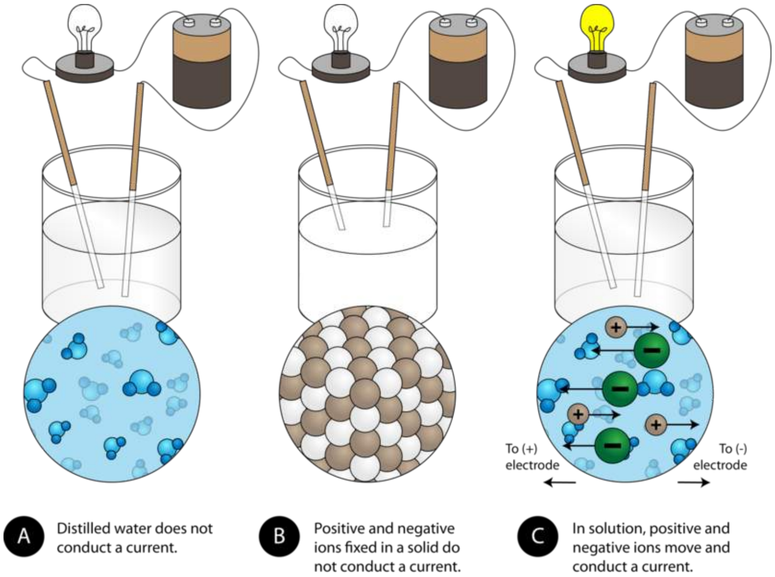


FIGURE 8.8

(A) The sodium chloride crystal is shown in two dimensions. (B) When struck by a hammer, the negatively-charged chloride ions are forced near each other and the repulsive force causes the crystal to shatter.



Distilled water does not conduct a current.

Positive and negative ions fixed in a solid do not conduct a current.

In solution, positive and negative ions move and conduct a current.

FIGURE 8.9

(A) Distilled water does not conduct electricity. (B) A solid ionic compound also does not conduct. (C) A water solution of an ionic compound conducts electricity well.

In the first beaker, distilled water does not conduct a current because water is a molecular compound. In the second beaker, solid sodium chloride also does not conduct a current. Despite being ionic and thus composed of charges particles, the solid crystal lattice does not allow the ions to move between the electrodes. Mobile charged particles are required for the circuit to be complete and the light bulb to light up. In the third beaker, the NaCl has been dissolved into the distilled water. Now the crystal lattice has been broken apart and the individual positive and negative ions can move. Cations move to one electrode, while anions move to the other, allowing electricity to flow (**Figure** 8.10). Melting an ionic compound also frees the ions to conduct a current. *Ionic compounds conduct an electric current when melted or dissolved in water.*

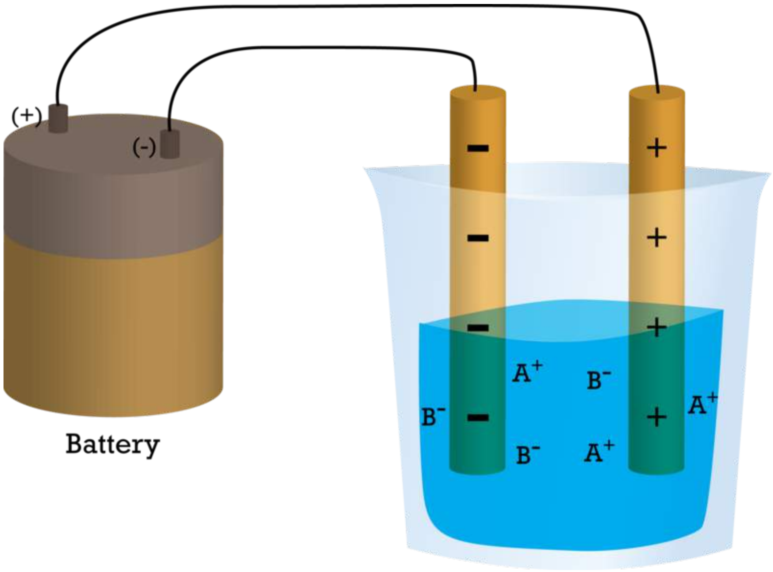


FIGURE 8.10

In an ionic solution, the A^+ ions migrate toward the negative electrode, while the B^- ions migrate toward the positive electrode.

Lesson Summary

- One or more electrons are transferred from a metal atom to a nonmetal atom to form ions. Ionic bonds are the electrostatic attractions between positive and negative ions.
- An ionic compound is a three-dimensional network of alternating cations and anions mutually attracted to one another. The coordination number of an ion is the number of nearest neighbors that it has within the crystal lattice.
- Ionic compounds are hard and have high melting points. They are difficult to break, but are very brittle. They conduct electricity only when melted or dissolved in water to form a solution.

Lesson Review Questions

Reviewing Concepts

1. Draw an electron dot diagram for the following atoms and ions.
    a. S
    b. S^{2-}
    c. Cs

   d. Cs^+
   e. P
   f. P^{3-}
   g. Sn
   h. Sn^{2+}

2. Which of the following pairs of atoms could be expected to combine chemically to form an ionic compound? Explain.
   a. Li and O
   b. N and H
   c. Al and S
   d. Cl and F
   e. Sr and Br
   f. Zn and I
3. Explain why most ionic compounds are strong and hard, yet brittle.
4. Explain why potassium fluoride does not conduct an electric current as a solid, but does conduct after being dissolved in water.

Problems

5. Use electron dot diagrams to demonstrate the formation of ionic compounds involving the following elements. Use arrows to show the transfer of electron(s) from one atom to another.
   a. K and O
   b. Ca and N
   c. Ba and S
6. Write the formula units for each of the ionic compounds from question number 5.
7. Answer the following:
   a. What is a coordination number?
   b. If the coordination numbers for each of the two ions in a crystal lattice are identical, what must be true about the formula unit of the compound?
   c. An ionic compound forms between metal A and nonmetal B. The coordination number of the cation of element A is 4 and the coordination number of the anion of element B is 8. Write the chemical formula of the compound.
8. A general ionic bond forms between a cation X^+ and an anion Y^-. How will the strength of the ionic bond change if the following changes are made? In other words, will the resulting bond be stronger or weaker?
   a. The charge of the cation is doubled.
   b. The size of the anion is increased.

Further Reading / Supplemental Links

- Ionic Bond (http://en.wikipedia.org/wiki/Ionic_bond
- Bonding (http://www.chemteam.info/Bonding/Bonding.html
- Claude H. Yoder, *Ionic Compounds: Applications of Chemistry to Mineralogy*. Wiley-Interscience, 2006.

Points to Consider

Metals are also crystalline materials. For a pure metal, each lattice point is an atom of the metal, rather than cations and anions as in an ionic crystal lattice. The familiar properties of metals result from its crystalline structure.

- What are the physical properties of metals?
- What is an alloy?

8.3 Metallic Bonds

Lesson Objectives

- Describe the electron-sea model of metallic bonding.
- Explain how metallic bonding is responsible for the conductivity and luster of metals.
- Explain why metals are malleable and ductile, while ionic-crystalline compounds are not.
- Describe how metal atoms are arranged, including the three most common packing systems.
- Identify some common alloys and explain their importance.

lesson Vocabulary

- alloy
- closest packing
- metallic bond

Check Your Understanding

Recalling Prior Knowledge

- How are the cations and anions in an ionic crystal arranged?
- What happens to an ionic crystal when it is put under a large stress?

The bonding that occurs in a metal is responsible for its distinctive properties: luster, malleability, ductility, and excellent conductivity.

The Metallic Bond

Pure metals are crystalline solids, but unlike ionic compounds, every point in the crystal lattice is occupied by an identical atom. The electrons in the outer energy levels of a metal are mobile and capable of drifting from one metal atom to another. This means that the metal is more properly viewed as an array of positive ions surrounded by a "sea of mobile valence electrons." Electrons which are capable of moving freely throughout the empty orbitals of the metallic crystal are called delocalized electrons (**Figure** 8.11). A **metallic bond** *is the attraction of the stationary metal cations to the surrounding mobile electrons*.

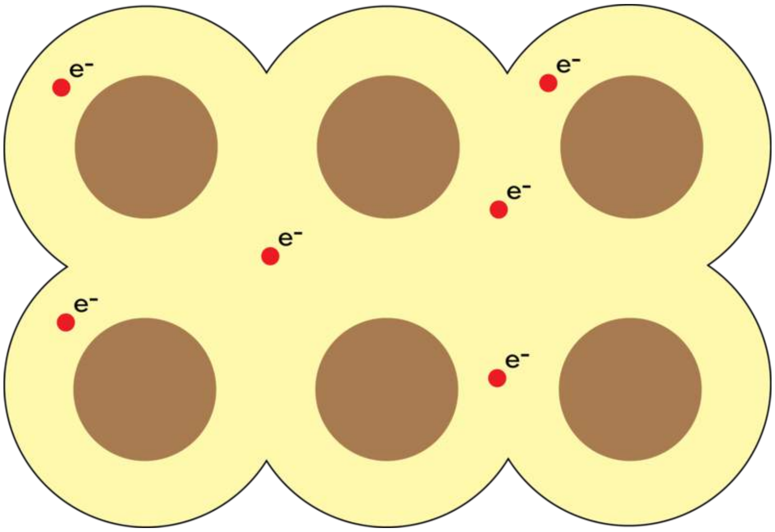


FIGURE 8.11

In a metal, the stationary metal cations are surrounded by a sea of mobile valence electrons that are not associated with any one cation.

Properties of Metals

The metallic bonding model explains the physical properties of metals. Metals conduct electricity and heat very well because of their free-flowing electrons. As electrons enter one end of a piece of metal, an equal number of electrons flow outward from the other end. When light is shone on to the surface of a metal, its electrons absorb small amounts of energy and become excited into one of its many empty orbitals. The electrons immediately fall back down to lower energy levels and emit light. This process is responsible for the high luster of metals (**Figure** 8.12).

FIGURE 8.12

The American Platinum Eagle is the official platinum bullion coin of the United States and was first minted in 1997. The luster of a metal is due to its metallic bonds.

Recall that ionic compounds are very brittle. Application of a force results in like-charged ions in the crystal coming too close to one another, causing the crystal to shatter. When a force is applied to a metal, the free-flowing electrons can slip in between the stationary cations and prevent them from coming in contact. Imagine ball bearings that have been coated with oil sliding past one another. As a result, metals are very malleable and ductile. They can be hammered into shapes, rolled into thin sheets, or pulled into thin wires.

Crystal Structures of Metals

If you wanted to make a stack of identical spheres, you may come up with an arrangement like that shown in **Figure** 8.13.

FIGURE 8.13

When identical spheres are stacked, each successive layer fits into the small spaces where different spheres come together. Closest packing minimizes the amount of empty space and is how metal atoms are arranged in a crystal.

This orderly and regular arrangement of the metal balls minimizes the empty space between them. **Closest packing** *is the most efficient arrangement of spheres.* Atoms of a metal crystal are arranged in similar patterns, called close-packed structures. Pure metals adopt one of several related close-packed structures as shown in **Figure** 8.14.

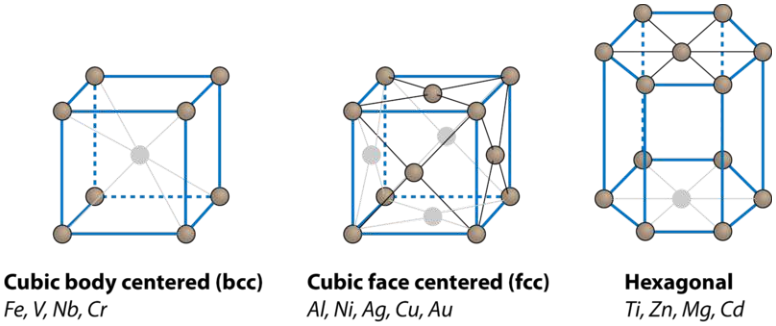


FIGURE 8.14

Most pure metals naturally adopt one of these three closest packing arrangements.

On the far left is the body-centered cubic (bcc) structure. In that crystal, metal atoms occupy the eight corners of a cube along with one atom in the very center. The coordination number of each atom in the body-centered cubic structure is 8. In the face-centered cubic (fcc) structure, there are eight atoms at each corner of the cube and six atoms in the center of each face. The coordination number of each atom in the face-centered cubic structure is 12. The hexagonal close-packed (hcp) structure also has a coordination number of 12, but crystals of this type are hexagonally shaped rather than cubic.

Alloys

An **alloy** *is a mixture composed of two or more elements, at least one of which is a metal*. You are probably familiar with some alloys such as brass and bronze (**Figure** 8.15). Brass is an alloy of copper and zinc. Bronze is an alloy of copper and tin. Alloys are commonly used in manufactured items because the properties of these metal mixtures are often superior to a pure metal. Bronze is harder than copper and more easily cast. Brass is very malleable and its acoustic properties make it useful for musical instruments.

FIGURE 8.15

Bronze, an alloy of copper and tin, has been in use since ancient times. The Bronze Age saw the increased use of metals rather than stone for weapons, tools, and decorative objects. Brass, an alloy of copper and zinc, is widely used in musical instruments like the trumpet and trombone.

Steels are a very important class of alloys. The many types of steels are primarily composed of iron, with various amounts of the elements carbon, chromium, manganese, nickel, molybdenum, and boron. Steels are widely used in building construction because of their strength, hardness, and resistance to corrosion. Most large modern structures like skyscrapers and stadiums are supported by a steel skeleton (**Figure** 8.16).

Alloys can be one of two general types. In one type, called a substitutional alloy, the various atoms simply replace each other in the crystal structure. In another type, called an interstitial alloy, the smaller atoms such as carbon fit in between the larger atoms in the crystal packing arrangement.

Lesson Summary

- A metallic bond is the attraction of stationary metal cations to the surrounding sea of mobile valence electrons. Metallic bonding gives metals their properties of high electrical and thermal conductivity, malleability, conductivity, and luster.
- Metal atoms are arranged in regular, orderly patterns called closest packing.
- Alloys are mixtures of metals that are widely used because the properties of alloys are often superior to those of pure metals.

FIGURE 8.16

The Willis Tower (formerly called the Sears Tower) in Chicago was once the tallest building in the world and is still the tallest in the Western Hemisphere. The use of steel columns makes it possible to build taller, stronger, and lighter buildings.

Lesson Review Questions

Reviewing Concepts

1. Which of the following elements displays metallic bonding?
    a. Se
    b. Ti
    c. Ru
    d. Te
2. Explain the behavior of electrons in a metal.
3. How does a metallic bond contribute to the electrical conductivity of metals?
4. Name the three common closest packing arrangements of metal atoms in a crystal. Give an example of a metallic element that crystallizes in each.

Problems

5. Why are ionic crystals brittle, while most metals are malleable?

6. What are some advantages of using steel rather than iron?
7. Sodium has one valence electron per atom, while magnesium has two. Predict whether sodium or magnesium has stronger metallic bonds. How do you think that you could test your hypothesis?
8. What are the coordination numbers of a metal atom in a body-centered cubic structure and in a face-centered cubic structure? In which type are the atoms more closely packed? Iridium is one of the densest known elements at 22.6 g/cm^3. In which crystal structure do you think iridium is more likely to crystallize?

Further Reading / Supplemental Links

- Metallic Bonding (http://www.chemguide.co.uk/atoms/bonding/metallic.html
- V.K. Grigorovich, *The Metallic Bond and the Structure of Metals*. Nova Science Pub. Inc. (1989).

Points to Consider

Molecular compounds are a class of substances that take the form of individual molecules.

- Which types of elements make up molecular compounds?
- How is the chemical bonding within a molecular compound different from the bonding that occurs in ionic compounds or metals?
- How are the physical properties of molecular compounds different from ionic compounds or metals?

NC Standards Addressed

Chm.1.2 Understand the bonding that occurs in simple compounds in terms of bond type, strength, and properties.

Chm.1.2.1

- Describe how ions are formed and which arrangements are stable (filled d-level, or half-filled d-level).
- Appropriately use the term cation as a positively charged ion and anion as negatively charged ion.
- Predict ionic charges for representative elements based on valence electrons.

Chm.1.2.2

- Determine that a bond is predominately ionic by the location of the atoms on the Periodic Table (metals combined with nonmetals) or when ΔEN >1.7.

Chm 1.2.5

- Explain how ionic bonding in compounds determines their characteristics: high MP, high BP, brittle, and high electrical conductivity
- either in molten state or in aqueous solution.

8.4 References

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2. (A) Andy Wright (Flickr: rightee); (B) Jeffrey Scism (Flickr: scismgenie). (A) http://www.flickr.com/photos/rightee/4356950/; (B) http://www.flickr.com/photos/jgscism/6960776486/ . (A) CC-BY 2.0; (B) CC-BY-NC-SA 2.0
3. (A) Paulnasca; (B) Ben Mills (Wikimedia: Benjah-bmm27). (A) http://commons.wikimedia.org/wiki/File:Rust_screw.jpg; (B) http://commons.wikimedia.org/wiki/File:Iron%28II%29-sulfate-heptahydrate-sample.jpg . (A) CC-BY 2.0; (B) Public Domain
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CHAPTER 9

Chemical Nomenclature

Chapter Outline

You are familiar with the names of many chemical compounds. Recall that a compound is two or more elements that have been chemically combined. Water is a chemical compound and you probably know that its chemical formula is H_2O. That formula tells you that water is composed of the elements hydrogen and oxygen. However, the physical and chemical properties of liquid water are nothing like the properties of hydrogen and oxygen, which are both gases. Water is an example of a common name that is given to a compound because it is something that everybody is accustomed to seeing and using every day. Yet there are millions of known chemical compounds out there. To give each and every one of them a common name would be a hopeless and confusing task. The figure above is a famous painting of French chemist Antoine Lavoisier (1743-1794) and his wife. Lavoisier is generally considered to be the founder of modern chemistry. Among his many accomplishments was the recognition that a systematic method was needed for naming the ever-increasing number of chemical compounds that were being created. Such a naming system is called nomenclature. This chapter will teach you the rules of nomenclature and allow you to name and write formulas for many simple chemical compounds.

Opening image by Jacques–Louis David. Public Domain.

9.1 Ionic Compounds

Lesson Objectives

- Distinguish between inorganic and organic chemistry.
- Interpret a molecular formula.
- Explain why an ionic compound is represented by an empirical formula.
- Be able to determine the charge of monatomic ions from the position of the element on the periodic table.
- Use the Stock system to identify the charge of transition metal ions.
- Name an ionic compound given its formula.
- Write the correct formula for an ionic compound given its name.

Lesson Vocabulary

- binary ionic compound
- empirical formula
- inorganic chemistry
- molecular formula
- monatomic ion
- organic chemistry
- polyatomic ion
- ternary ionic compound

Check Your Understanding

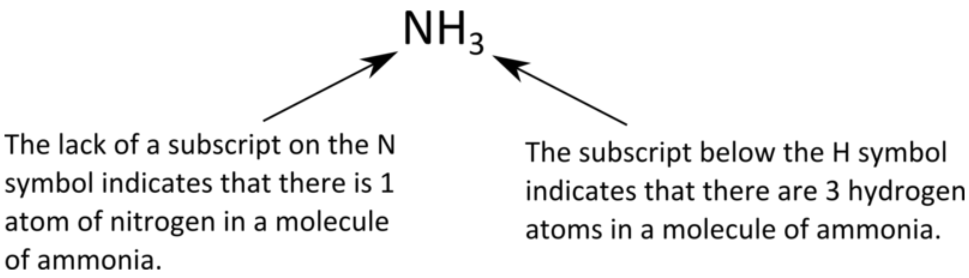
Recalling Prior Knowledge

- Which subatomic particles are lost or gained in order to form ions?
- The elements of which group on the periodic table are particularly stable and unreactive?
- Which groups contain the representative elements and which groups contain the transition elements?

Chemistry can be broadly divided into two main classes based on the identity of the elements present in the chemical compounds. **Organic chemistry** *is the branch of chemistry that deals with compounds containing carbon*. Organic chemistry is very complex and will be covered in a later chapter. This chapter is about inorganic chemistry. **Inorganic chemistry** *is the branch of chemistry that deals with compounds that do not contain carbon*. There are a few exceptions of small molecules that contain carbon but are not subject to the rules for naming organic compounds. We will begin our study with an examination of two types of formulas. The remainder of this lesson will be about the naming and formula writing of ionic compounds.

Types of Formulas

Recall that a molecule is two or more atoms that have been chemically combined. A **molecular formula** *is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound*. Ammonia is a compound of nitrogen and hydrogen as shown below:



An **empirical formula** *is a formula that shows the elements in a compound in their lowest whole-number ratio*. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O. When chemists analyze an unknown compound, often the first step is to determine its empirical formula, as we will see in a later chapter. There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.

Ionic compounds are quite different from molecular compounds such as water. Water and other molecular compounds exist as individual molecules (**Figure** 9.1).

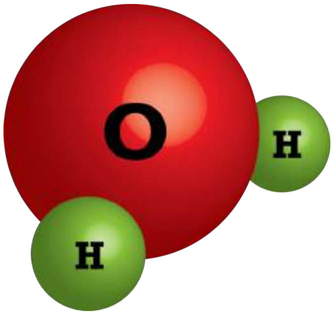
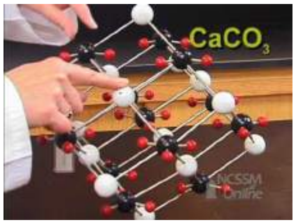


FIGURE 9.1

A water molecule consists of one atom of oxygen bonded to two atoms of hydrogen.

Ionic compounds do not exist as discrete molecular units. Instead an ionic compound such as sodium chloride (NaCl) consists of a large three-dimensional array of alternating Na^+ and Cl^- ions (**Figure** 9.2).

Watch a video about The Structure of Ionic Solids: http://www.youtube.com/watch?v=TLPY9Z6z4Mg (1:35) .



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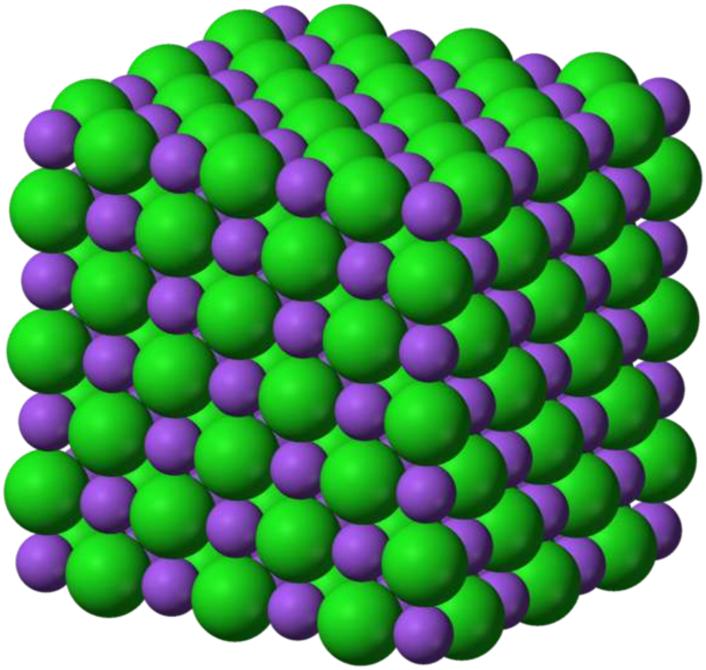


FIGURE 9.2

A crystal of table salt, sodium chloride, is a large array of alternating positive and negative ions. The purple spheres represent the Na^+ ions, while the green spheres represent the Cl^- ions.

Because of this structure, the chemical formulas of ionic compounds always represent the lowest whole-number ratio between the two ions. In the case of NaCl, there are equal numbers of sodium ions and chloride ions in the salt crystal. In the case of magnesium chloride, with a formula of $MgCl_2$, there are twice as many chloride ions as magnesium ions in each crystal. The formulas of ionic compounds are always empirical formulas.

Monatomic Ions

An ionic compound consists of a positive metal ion and a negative nonmetal ion combined in a proportion so that the overall compound is electrically neutral. In other words, the total positive charge is exactly canceled out by the total negative charge. So in order to accurately write formulas for all ionic compounds, it is necessary to know the charges of the ions involved. We will begin our study with **monatomic ions**, *which are ions that consist of a single atom with either a positive or negative charge*.

Cations

Recall that cations are positive ions that are formed when a metal atom loses one or more electrons. For the representative elements, all of the valence electrons are removed from the atom. Since the valence electrons are constant within a particular group, all we need to know is the group number of the given element and we will be able to know its charge when it becomes a cation. Group 1 elements form ions with a 1+ charge, Group 2 with a 2+ charge, and Group 13 with a 3+ charge. The heavier *p*-block metals such as tin and lead are a special case and will be discussed with the transition metal ions. Naming monatomic ions is done simply by using the element's name. For example, the Na^+ ion is the sodium ion, while the Al^{3+} ion is the aluminum ion.

Anions

Anions are negative ions that are formed when a nonmetal atom gains one or more electrons. Atoms typically gain electrons so that they will have the electron configuration of a noble gas. All the elements in Group 17 have seven valence electrons due to the outer ns^2np^5 configuration. Therefore, each of these elements would gain one electron and become an anion with a 1− charge. Likewise, Group 16 elements form ions with a 2− charge, and the Group 15 nonmetals form ions with a 3− charge. Naming anions is slightly different than naming cations. The ending of the element's name is dropped and replaced with the *–ide* suffix. For example, F^- is the fluoride ion, while O^{2-} is the oxide ion. **Table** 9.1 shows the common monatomic ions for the representative elements.

TABLE 9.1: **Common Monatomic Ions**

1+	2+	3+	3-	2-	1-
lithium, Li^+	beryllium, Be^{2+}	aluminum, Al^{3+}	nitride, N^{3-}	oxide, O^{2-}	fluoride, F^-
sodium, Na^+	magnesium, Mg^{2+}	gallium, Ga^{3+}	phosphide, P^{3-}	sulfide, S^{2-}	chloride, Cl^-
potassium, K^+	calcium, Ca^{2+}		arsenide, As^{3-}	selenide, Se^{2-}	bromide, Br^-
rubidium, Rb^+	strontium, Sr^{2+}			telluride, Te^{2-}	iodide, I^-
cesium, Cs^+	barium, Ba^{2+}				

Transition Metal Ions

Most transition metals differ from the metals of Groups 1, 2, and 13 in that they are capable of forming more than one cation with different ionic charges. As an example, iron commonly forms two different ions. It can sometimes lose two electrons to form the Fe^{2+} ion, while at other times it loses three electrons to form the Fe^{3+} ion. Tin and lead, though members of the *p* block rather than the *d* block, also are capable of forming multiple ions. The names of these types of ions needs to indicate the particular charge that ion carries. The Stock system of naming these ions uses a Roman numeral in parentheses after the name of the ion. The iron ions previously mentioned are named the iron(II) ion and iron(III) ion. In speaking, the name of iron(II) is said, "iron two ion." **Table** 9.2 lists the names and formulas of some of the common transition metal ions.

TABLE 9.2: **Common Transition Metal Ions**

1+	2+	3+	4+
copper(I), Cu^+	cadmium, Cd^{2+}	chromium(III), Cr^{3+}	lead(IV), Pb^{4+}
gold(I), Au^+	chromium(II), Cr^{2+}	cobalt(III), Co^{3+}	tin(IV), Sn^{4+}
mercury(I), $Hg_2{}^{2+}$	cobalt(II), Co^{2+}	gold(III), Au^{3+}	
silver, Ag^+	copper(II), Cu^{2+}	iron(III), Fe^{3+}	
	iron(II), Fe^{2+}		
	lead(II), Pb^{2+}		
	manganese(II), Mn^{2+}		
	mercury(II), Hg^{2+}		
	nickel(II), Ni^{2+}		
	platinum(II), Pt^{2+}		
	tin(II), Sn^{2+}		
	zinc, Zn^{2+}		

Notice in **Table** 9.2 that there are three cation names that lack the Roman numeral. Silver, cadmium, and zinc only form ions with one charge, so by convention the Stock system is not used with these ions. The mercury(I) ion is a special case in that consists of a pair of mercury atoms bonded together and a total charge of 2+. So each mercury ion can be thought of as carrying a 1+ charge. In some instances, charges greater than 4+ are seen, but these are not the most common charges and so are not included in the table.

There is an older system for naming some of these cations that is still occasionally used. The Latin root of the metal name is written with one of two suffixes: (1) *–ic* for the ion with a higher charge, and (2) *–ous* for the ion with a lower charge. Again, iron will serve as an example. The Latin name for iron is ferrum, so the Fe^{3+} ion is called the ferric ion, while the Fe^{2+} ion is called the ferrous ion. The primary disadvantage of this system is that the suffixes do not tell you exactly what the charge is for a given ion. For copper, the Cu^{2+} is the highest charge, so it is called the cupric ion, while the Cu^{+} ion is the cuprous ion. The Stock system is a much more specific system and will be used as the primary method of naming transition metal compounds throughout this book.

Binary Ionic Compounds

A **binary ionic compound** *is a compound composed of a monatomic metal cation and a monatomic nonmetal anion.*

Naming Binary Ionic Compounds

When examining the formula of a compound in order to name it, you must first decide what kind of compound it is. For a binary ionic compound, a metal will always be the first element in the formula, while a nonmetal will always be the second. The metal cation is named first, followed by the nonmetal cation. Subscripts in the formula do not affect the name. **Table** 9.3 shows three examples.

TABLE 9.3: **Examples of Binary Ionic Compounds**

Formula	Name
KF	potassium fluoride
Na_3N	sodium nitride
Ca_3P_2	calcium phosphide

Notice that in each of the formulas above, the overall charge of the compound is zero. Potassium ion is K^{+}, while fluoride ion is F^{-}. Since the magnitude of the charges is equal, the formula contains one of each ion. This would also be the case for a compound such as MgS, in which the ions are Mg^{2+} and S^{2-}. For sodium nitride, the sodium ion is Na^{+}, while the nitride ion is N^{3-}. In order to make a neutral compound, three of the 1+ sodium ions are required in order to balance out the single 3− nitride ion. So the Na is given a subscript of 3. For calcium phosphide, the calcium ion is Ca^{2+}, while the phosphide ion is P^{3-}. The least common multiple of 2 and 3 is 6. To make the compound neutral, three calcium ions have a total charge of 6+, while two phosphide ions have a total charge of 6−. The Ca is given a subscript of 3, while the P is given a subscript of 2.

Practice naming binary ionic compounds with this short crossword puzzle: https://sites.google.com/site/pattihowellsciencespot/uploads/binary%20ionic%20compounds.html?attredirects=0&d=1 .

Naming Compounds Using the Stock System

Naming compounds that involve transition metal cations necessitates the use of the Stock system. Consider the binary ionic compound $FeCl_3$. To simply name this compound iron chloride would be incomplete because iron is capable of forming two ions with different charges. The name of any iron-containing compound must reflect which iron ion is in the compound. In this case, the subscript in the formula indicates that there are three chloride ions, each with a 1− charge. Therefore, the charge of the single iron ion must be 3+. The correct name of $FeCl_3$ is iron(III) chloride, with the cation charge written as the Roman numeral. **Table** 9.4 some several other examples.

TABLE 9.4: **Examples of Naming Using the Stock System**

Formula	Name
Cu_2O	copper(I) oxide
CuO	copper(II) oxide
SnO_2	tin(IV) oxide

The first two are both oxides of copper (shown in **Figure** 9.3). The ratio of copper ions to oxide ions determines the name. Since the oxide ion is O_{2-}, the charges of the copper ion must be 1+ in the first formula and 2+ in the second formula. In the third formula, there is one tin ion for every two oxide ions. This means that the tin must carry a 4+ charge, making the name tin(IV) oxide.

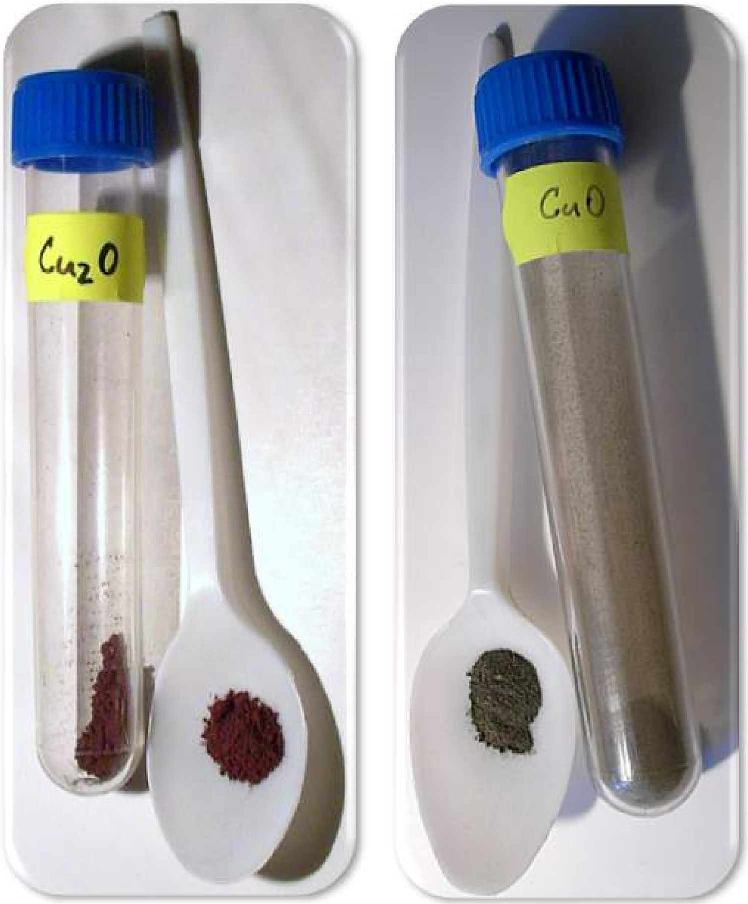


FIGURE 9.3

Copper(I) oxide, a red solid, and copper(II) oxide, a black solid, are different compounds because of the charge of the copper ion.

You can practice naming multivalent metal compounds by completing this online crossword puzzle: https://sites.google.com/site/pattihowellsciencespot/uploads/multivalent%20metals%20compounds.html?attredirects=0&d=1 .

Writing Formulas for Binary Ionic Compounds

If you know the name of a binary ionic compound, you can write its formula. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charge to cancel each other out. Consider the compound aluminum nitride. The ions are:

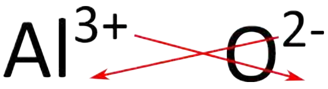
Al^{3+} N^{3-}

Since the ions have charges that are equal in magnitude, one of each will be the lowest ratio of ions in the formula. The formula of aluminum nitride is AlN. The ions for the compound lithium oxide are:

Li^{+} O^{2-}

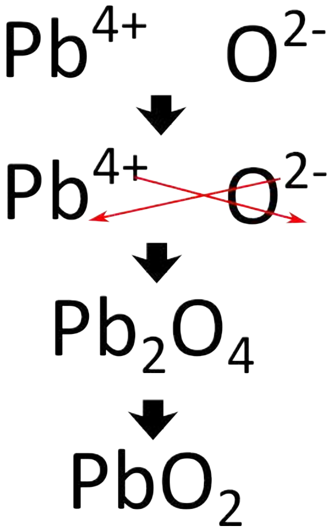
In this case, two lithium ions are required to balance out the charge of one oxide ion. The formula of lithium oxide is Li_2O.

An alternative way to writing a correct formula for an ionic compound is to use the crisscross method. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped. Shown below is the crisscross method for aluminum oxide.



The red arrows indicate that the 3 from the 3+ charge will cross over to become the subscript of the O. The 2 from the 2− charge will cross over to become the subscript of the Al. The formula for aluminum oxide is Al_2O_3.

Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions. In the case of aluminum nitride, the crisscross method would yield a formula of Al_3N_3, which is not correct. It must be reduced to AlN. Following the crisscross method to write the formula for lead(IV) oxide would involve the following steps:



The crisscross first yields Pb_2O_4 for the formula, but that must be reduced to the lower ratio and PbO_2 is the correct formula.

Click here to watch an animation of ionic bonding: http://www.dlt.ncssm.edu/core/Chapter9-Bonding_and_Geometry/Chapter9-Animations/IonicBonding.html .

Sample Problem 7.1: Binary Ionic Compound Formulas

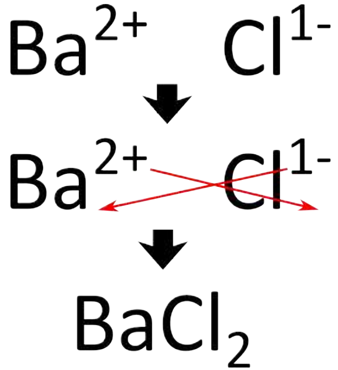
Write the correct formulas for the following ionic compounds:

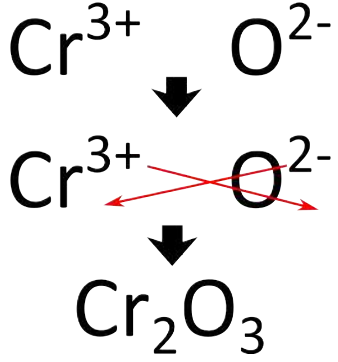
1. barium chloride
2. chromium(III) oxide

Step 1: Plan the problem.

In each case, write the metal cation, followed by the nonmetal anion. Crisscross the ion charges in order to make the ionic compound neutral. Reduce to the lowest ratio if necessary.

Step 2: Solutions

1. 

2. 

Step 3: Think about your result.

The formula for barium chloride is $BaCl_2$. Notice that the charge of the chloride ion is written as 1− rather than just –in order to do the crisscross. However, the number 1 is not used as a subscript. The formula for chromium(III) oxide is Cr_2O_3.

Practice Problems

1. Write formulas for the binary ionic compounds formed between the following pairs of elements:
    a. cesium and fluorine
    b. calcium and sulfur
    c. aluminum and chlorine
    d. zinc and nitrogen
2. Write the formula and give the name for the compound formed by the following ions:
    a. Fe^{3+} and O^{2-}
    b. Ni^{2+} and S^{2-}
    c. Au^{+} and Cl^{-}
    d. Sn^{4+} and I^{-}
3. Give names for the following compounds:
    a. $Ag_2{}^S$
    b. PdO
    c. $PtCl_4$
    d. V_2O_5

Ternary Ionic Compounds

Not all ionic compounds are composed of only monatomic ions. A **ternary ionic compound** *is an ionic compound composed of three or more elements*. In a typical ternary ionic compound, there is still one type of cation and one type of anion involved. The cation, the anion, or both, is a polyatomic ion.

Polyatomic Ions

A **polyatomic ion** *is an ion composed of more than one atom*. The ammonium ion consists of one nitrogen atom and four oxygen atoms. Together, they comprise a single ion with a 1+ charge and a formula of $NH_3{}^+$. The carbonate ion consists of one carbon atom and three oxygen atoms and carries an overall charge of 2−. The formula of the carbonate ion is $CO_3{}^{2-}$. The atoms of a polyatomic ion are tightly bonded together and so the entire ion behaves as a single unit. **Figure** 9.4 shows several models and **Table** 9.5 lists many of the most common polyatomic ions.

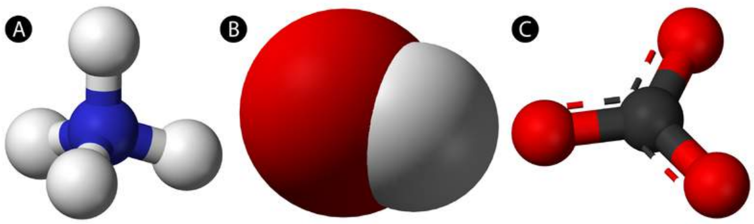


FIGURE 9.4

(A) The ammonium ion ($NH_4{}^+$) is a nitrogen atom (blue) bonded to four hydrogen atoms (white). (B) The hydroxide ion (OH^-) is an oxygen atom (red) bonded to a hydrogen atom. (C) The carbonate ion ($CO_3{}^{2-}$) is a carbon atom (black) bonded to three oxygen atoms.

TABLE 9.5: Common Polyatomic Ions

1-	2-	3-	1+	2+
acetate, CH_3COO^-	carbonate, $CO_3{}^{2-}$	arsenate, $AsO_3{}^{3-}$	ammonium, $NH_4{}^+$	dimercury, $Hg_2{}^{2+}$
bromate, $BrO_3{}^-$	chromate, $CrO_4{}^{2-}$	phosphite, $PO_3{}^{3-}$		
chlorate, $ClO_3{}^-$	dichromate, $Cr_2O_7{}^{2-}$	phosphate, $PO_4{}^{3-}$		
chlorite, $ClO_2{}^-$	hydrogen phosphate, $HPO_4{}^{2-}$			
cyanide, CN^-	oxalate, $C_2O_4{}^{2-}$			
dihydrogen phosphate, $H_2PO_4{}^-$	peroxide, $O_2{}^{2-}$			
hydrogen carbonate, $HCO_3{}^-$	silicate, $SiO_3{}^{2-}$			
hydrogen sulfate, $HSO_4{}^-$	sulfate, $SO_4{}^{2-}$			
hydrogen sulfide, HS^-	sulfite, $SO_3{}^{2-}$			
hydroxide, OH^-				
hypochlorite, ClO^-				
nitrate, $NO_3{}^-$				
nitrite, $NO_2{}^-$				
perchlorate, $ClO_4{}^-$				
permanganate, $MnO_4{}^-$				

The vast majority of polyatomic ions are anions, many of which end in *–ate* or *–ite*. Notice that in some cases such as nitrate ($NO_3{}^-$) and nitrite ($NO_2{}^-$), there are multiple anions that consist of the same two elements. In these cases, the difference between the ions is in the number of oxygen atoms present, while the overall charge is the same. As a class, these are called oxoanions. When there are two oxoanions for a particular element, the one with the greater number of oxygen atoms gets the *–ate* suffix, while the one with the fewer number of oxygen atoms gets the *–ite* suffix. The four oxoanions of chlorine are shown below.

- ClO^-, hypochlorite
- $ClO_2{}^-$, chlorite
- $ClO_3{}^-$, chlorate
- $ClO_4{}^-$, perchlorate

In cases such as this, the ion with one more oxygen atom than the *–ate* anion is given a *per-* prefix. The ion with one fewer oxygen atom than the *–ite* anion is given a *hypo-* prefix.

Naming Ternary Ionic Compounds

The process of naming ternary ionic compounds is the same as naming binary ionic compounds. The cation is named first, followed by the anion. Some examples are shown in **Table** 9.6:

TABLE 9.6: Examples of Ternary Ionic Compounds

Formula	Name
$NaNO_3$	sodium nitrate

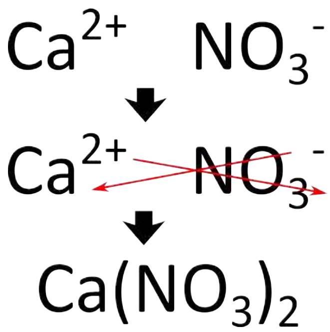
TABLE 9.6: (continued)

Formula	Name
NH_4Cl	ammmonium chloride
$Fe(OH)_3$	iron(III) hydroxide

When more than one polyatomic ion is present in a compound, the formula of the ion is placed in parentheses with a subscript outside of the parentheses that indicates how many of those ions are in the compound. In the last example above, there is one Fe^{3+} cation and three OH^- anions.

Writing Formulas for Ternary Ionic Compounds

Writing a formula for a ternary ionic compound also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion. Use the crisscross method to ensure that the final formula is neutral. Calcium nitrate is composed of a calcium cation and a nitrate anion.



The charge is balanced by the presence of two nitrate ions and one calcium ion. Parentheses are used around the nitrate ion because more than one of the polyatomic ion is needed. If only one polyatomic ion is in a formula, parentheses are not used. As an example, the formula for calcium carbonate is $CaCO_3$. The carbonate ion carries a 2− charge and so exactly balances the 2+ charge of the calcium ion.

Sample Problem 7.2: Ternary Ionic Compound Formulas

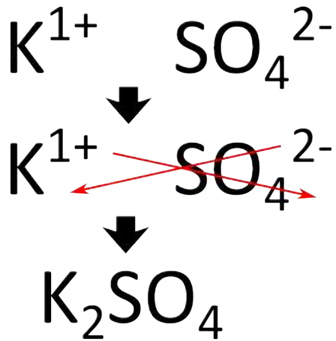
Write the correct formulas for the following ionic compounds:

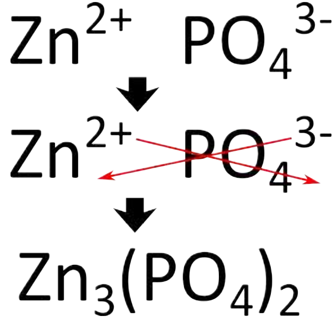
1. potassium sulfate
2. zinc phosphate

Step 1: Plan the problem.

In each case, write the metal cation, followed by the nonmetal anion. Crisscross the ion charges in order to make the ionic compound neutral. Use parentheses around the polyatomic ion if more than one is present in the final formula. Reduce to the lowest ratio if necessary.

Step 2: Solutions

1. 

2. 

Step 3: Think about your result.

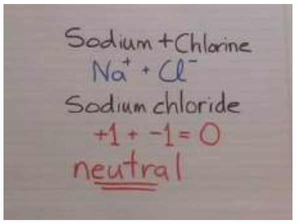
The formula for potassium sulfate is K_2SO_4. Two potassium cations with 1+ charge balance out the 2− charge of the sulfate ion. The formula for zinc phosphate is $Zn_3(PO_4)_2$. Three zinc cations with a 2+ charge balance out two phosphate anions with a 3− charge.

Practice Problems

4. Name the following compounds:
    - (a) NH_4NO_3
    - (b) $Na_2Cr_2O_7$
    - (c) $PbCO_3$
    - (d) $Mg(CH_3COO)_2$
5. Write formulas for the following compounds:
    - (a) potassium hydrogen sulfate
    - (b) iron(III) oxalate
    - (c) sodium peroxide
    - (d) tin(IV) chromate

There are two polyatomic ions that produce unusual formulas. The $Hg_2{}^{2+}$ ion is called either the dimercury ion or, preferably, the mercury(I) ion. When bonded with an anion with a 1− charge, such as chloride, the formula is Hg_2Cl_2. Because the cation consists of two Hg atoms bonded together, this formula is not reduced to HgCl. Likewise, the peroxide ion, $O_2{}^{2-}$, is also a unit that must stay together in its formulas. For example, the formula for potassium peroxide is K_2O_2.

Watch a humorous video lecture about Naming Ionic Compounds from BossChemDude: http://www.youtube.com/watch?v=q2s8hQ5NIpE (8:30)



MEDIA

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Lesson Summary

- Inorganic chemistry is the study of chemical compounds that do not contain carbon.
- Molecular formulas show the type and number of atoms that occur in a molecule. Empirical formulas show the atoms or ions in their lowest whole-number ratio in a compound. Because ionic compounds have an extended three-dimensional structure, all formulas of ionic compounds are empirical formulas.
- Atoms of representative elements form monatomic ions by losing or gaining electrons in order to attain the electron configuration of a noble gas.
- Most transition metals are capable of forming multiple cations with different charges. The Stock system is used to show the charge as a Roman numeral in the name.
- The nomenclature of ionic compounds is always to name the cation first, followed by the name of the anion.
- Ionic compounds must have no net electrical charge, so the total amount of positive charge must balance the total amount of negative charge. The crisscross method can be used to write correct formulas for ionic compounds.
- Some ionic compounds contain polyatomic ions. A set of parentheses is used to designate more than one polyatomic ion in a formula.

Lesson Review Questions

Reviewing Concepts

1. What element is present in all organic compounds?
2. Write the molecular formula of a compound whose molecules contain one atom of nitrogen and three atoms of fluorine.
3. For which three transition elements is Roman numeral not used in the names of its compounds?
4. Give an example of each of the following: (1) a monatomic cation, (2) a monatomic anion, (3) a polyatomic cation, and (4) a polyatomic anion.
5. What is the overall charge of an ionic compound?
6. When are parentheses used in the formula of an ionic compound?

Problems

7. Write the empirical formula for the following compounds:
    a. C_2H_6
    b. Hg_2Cl_2
    c. Sb_2O_5
    d. $C_8H_{16}O_2$
8. Using only the periodic table, write the symbol of the common ion formed by each of the following elements:

   a. Sr
   b. I
   c. Se
   d. Ba
   e. P
   f. Rb
   g. Al
   h. Br

9. Name the following monatomic ions:
   a. O^{2-}
   b. Li^{+}
   c. W^{3+}
   d. Cu^{2+}
   e. Ga^{3+}
   f. F^{-}
10. Write formulas for binary ionic compounds formed from each of the following pairs of elements:
   a. potassium and sulfur
   b. silver and chlorine
   c. calcium and oxygen
   d. aluminum and iodine
   e. barium and nitrogen
   f. sodium and selenium
11. Name the following polyatomic ions:
   a. ${SO_3}^{2-}$
   b. ${MnO_4}^{-}$
   c. ${CO_3}^{2-}$
   d. ClO^{-}
   e. CH_3COO^{-}
   f. ${H_2PO_4}^{-}$
12. Write the symbol and charge for the following polyatomic ions:
   a. hydrogen sulfide ion
   b. nitrate ion
   c. perchlorate ion
   d. chromate ion
   e. phosphate ion
   f. oxalate ion
13. Name the following ionic compounds:
   a. $KClO_3$
   b. $Cd(NO_3)_2$
   c. $CuCl$
   d. $Ca_3(PO_4)_2$
   e. $NaCN$
   f. $(NH_4)_2SO_3$
   g. Li_2O_2
   h. PbS_2
   i. Rb_2SO_4
   j. Mn_3P_2
   k. $NiCO_3$

    l. $Co(OH)_3$
    m. Hg_2Br_2
    n. $Zn(NO_2)_2$

14. Write correct formulas for the following ionic compounds:
    a. copper(II) bromide
    b. aluminum hydroxide
    c. silver sulfide
    d. barium acetate
    e. mercury(II) nitrate
    f. lead(II) chromate
    g. potassium permanganate
    h. sodium hydrogen carbonate
    i. calcium silicate
    j. tin(IV) sulfate
    k. ammonium phosphate
    l. gold(III) fluoride
    m. magnesium bromate
    n. chromium(VI) oxide

Further Reading / Supplemental Links

- Antoine Lavoisier, *Elements of Chemistry*. Dover Publications, 1984.
- An Introduction to Chemistry: Shockwave Tutorials (http://preparatorychemistry.com/Bishop_ionic_nom_Flash1.htm
- List of inorganic compounds (http://en.wikipedia.org/List_of_inorganic_compounds

Points to Consider

The large class of inorganic chemical compounds can be sub-classified into ionic compounds and molecular compounds.

- How is the structures of a molecular compound different from the structure of an ionic compound?
- Relate molecular compounds to the Law of Multiple Proportions.
- What is the nomenclature of molecular compounds?

9.2 Molecular Compounds

Lesson Objectives

- Describe the difference between an ionic compound and a molecular compound.
- Name a molecular compound given its formula.
- Write the correct formula for a molecular compound given its name.

Lesson Vocabulary

- binary molecular compound

Check Your Understanding

Recalling Prior Knowledge

- Describe the relationships between the following: atom, element, molecule, and compound.

Inorganic chemical compounds can be broadly classified into two groups: ionic compounds and molecular compounds. In the last lesson, we learned that the structure of all ionic compounds is an extended three-dimensional array of alternating positive and negative ions. Since ionic compounds do not take the form of individual molecules, they are represented by empirical formulas. Now we will begin to examine the formulas and nomenclature of molecular compounds.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete molecules. Examples include such familiar substance as water (H_2O) and carbon dioxide (CO_2) (see **Figure** 9.5). These compounds are very different from ionic compounds like sodium chloride (NaCl). Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other. You will learn more about ionic chemical bonding in the chapter *Ionic and Metallic Bonding*.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their valence electrons in such a way that a bond forms between pairs of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

Larger molecules can have many, many bonds that serve to keep the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

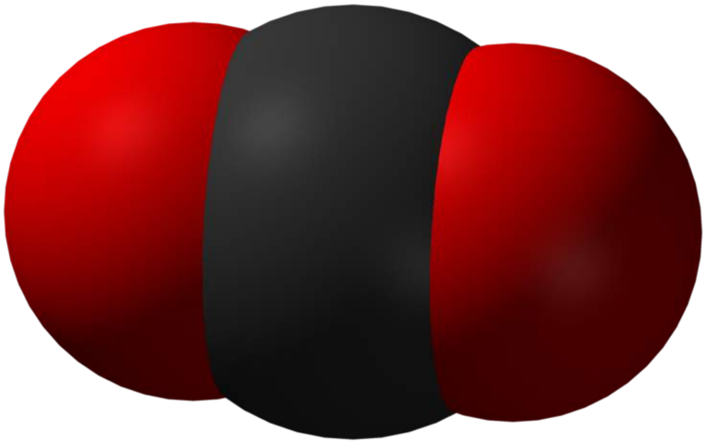


FIGURE 9.5

Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18}.

A **binary molecular compound** *is a molecular compound that is composed of two elements*. The elements that combine to form binary molecular compounds are both nonmetal atoms. This contrasts with ionic compounds, which were formed from a metal ion and a nonmetal ion. Therefore, binary molecular compounds are different because ionic charges cannot be used to name them or to write their formulas. Another difference is that two nonmetal atoms will frequently combine with one another in a variety of ratios. This discovery was discussed in the *Atomic Structure* chapter and led to the law of multiple proportions. Consider the elements nitrogen and oxygen. They combine to make several compounds including NO, NO_2, and N_2O. They all can't be called nitrogen oxide! How would someone know which one you were talking about? Each of the three compounds has very different properties and reactivity. A system to distinguish between compounds such as these is necessary.

Prefixes are used in the names of binary molecular compounds to identify the number of atoms of each element. **Table** 9.7 show the prefixes up to ten.

TABLE 9.7: **Numerical Prefixes**

Number of Atoms	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

FIGURE 9.6

Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

The rules for using the prefix system of nomenclature of binary molecular compounds can be summarized as follows.

1. Generally, the less-electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
2. When naming, the appropriate prefix is used only if there are more than one atom of that element in the formula.
3. The second element is named after the first, but with the ending of the element's name changed to *–ide*. The

appropriate prefix is always used for the second element.

4. The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms is tetroxide instead of tetraoxide.

Some examples of molecular compounds are listed in **Table** 9.8.

TABLE 9.8: Examples of Molecular Compounds

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Notice that the mono- prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The S_2Cl_2 emphasizes that the formulas for molecular compounds are not reduced to their lowest ratios. The *o* of mono- and the *a* of hepta- are dropped from the name when paired with oxide.

Writing Formulas for Binary Molecular Compounds

When you know the name of a molecular compound, the prefixes directly tell you which subscript to place with that element in the formula. If there is no prefix, only one atom of that element is present and no subscript is used. For example, if given the name diboron hexahydride, you would realize that the molecule must contain two atoms of boron and six atoms of hydrogen. Its formula is B_2H_6. Notice that metalloids such as boron generally form molecular rather than ionic compounds.

Lesson Summary

- Molecular compounds can be described as individual molecules.
- The formula of a binary molecular compound shows the number of each atom in the molecule with the less electronegative element written first.
- Prefixes are used in the names of molecular compounds to designate the number of each atom in the molecule.

Lesson Review Questions

Reviewing Concepts

1. What are the elements that make up a binary molecular compound?
2. What determines the order of the elements in the formula of a molecular compound?
3. Based on your answer to number 2, which of the formulas in each pair is correct?
    a. BrCl or ClBr
    b. ON_2 or N_2O
4. Give the prefix for each of the following numbers of atoms of an element in the formula of a molecular compound.
    a. 6

- b. 3
- c. 9
- d. 5
- e. 8

Problems

5. Name these molecular compounds.
    - a. CCl_4
    - b. ClF_3
    - c. N_2O_3
    - d. As_2O_5
    - e. B_4H_{10}
    - f. IF_7
6. Write correct formulas for the following molecular compounds.
    - a. xenon hexafluoride
    - b. tetraphosphorus decasulfide
    - c. oxygen difluoride
    - d. tellurium trioxide
    - e. germanium monoxide
    - f. trisilicon tetranitride
7. Name the compounds below. Some are ionic and some are molecular.
    - a. N_2O
    - b. Na_2O
    - c. $K_2Cr_2O_7$
    - d. P_2O_5
    - e. $CuSO_4$
    - f. SnO
    - g. BF_3
    - h. W_2O_3
8. Write formulas for the following compounds. Some are ionic and some are molecular.
    - a. sodium hydrogen sulfide
    - b. carbon disulfide
    - c. tetraselenium tetranitride
    - d. iron(II) dihydrogen phosphate
    - e. scandium(III) nitrate
    - f. ammonium hydroxide
    - g. bromine pentafluoride
    - h. aluminum selenide

Further Reading / Supplemental Links

- Molecular Compounds (http://www.buzzle.com/articles/molecular-compounds.html
- Click here for a crossword puzzle to practice naming common molecules: http://education.jlab.org/sciencecrossword/molecules_01.html .
- Click here for a practice quiz on inorganic nomenclature: http://www.funtrivia.com/playquiz/quiz115018d2cfb8.html .

Points to Consider

Acids and bases are a special and important class of inorganic compounds. When an acid or a base dissolves in water, it breaks apart into ions.

- What ion must be present in its water solution for a compound to be considered an acid? A base?
- How are acids and bases named?

9.3 Acids and Bases

Lesson Objectives

- Define an acid and a base in terms of ions that are present in each compound.
- Use the three rules to name acids given the formula.
- Write correct formulas for acids.
- Name and write formulas for bases.

Lesson Vocabulary

- acid
- base
- binary acid
- oxoacid

Check Your Understanding

Recalling Prior Knowledge

- What are cations and anions and how are they formed?
- How do ionic compounds and molecular compounds compare structurally?
- How do you make sure that formulas are electrically neutral?

Acids and bases are very two very important classes of chemicals that you probably use every day. Many foods and beverages such as citrus fruits and juices, soda, and vinegar, contain acids. Bases are very prevalent in cleaners such as ammonia and the active ingredients in drain cleaner. Acid-base chemistry will be covered in great detail in a later chapter. In this lesson, we will focus on the nomenclature of these two important compounds.

Acids

An acid can be defined in several ways. For the purposes of this chapter, the most straightforward is that an **acid** *is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions* (H^+) *when dissolved in water*.

This is a different type of compound than the others we have seen in this chapter. Acids are molecular, which means that in their pure state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the

FIGURE 9.7

(A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$HCl \rightarrow H^+ + Cl^-$$

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will primarily be discussed in a later chapter. A **binary acid** *is an acid that consists of hydrogen and one other element*. The most common binary acids contain a halogen. An **oxoacid** *is an acid that consists of hydrogen, oxygen, and a third element*. The third element is usually a nonmetal.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. Recall from earlier in the chapter that anions can either be monatomic or polyatomic. The name of all monatomic ions ends in *–ide*. The majority of polyatomic ions end in either *–ate* or *–ite*, though there are a few exceptions such as the cyanide ion (CN^-). It is this suffix of the anion that determines how the acid is named as displayed in **Table** 16.1 and the rules below.

TABLE 9.9: **Naming System for Acids**

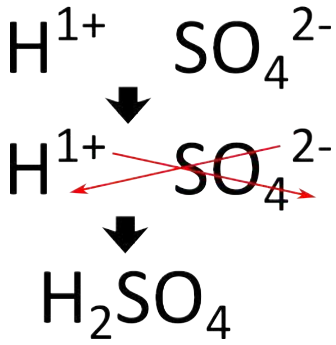
Anion Suffix	Example	Name of acid	Example
-ide	chloride (Cl^-)	hydro_____ic acid	hydrochloric acid (HCl)
-ate	sulfate ($SO_4{}^{2-}$)	_____ic acid	sulfuric acid H_2SO_4)
-ite	nitrite ($NO_2{}^-$)	_____ous acid	nitrous acid (HNO_2)

The three different suffixes that are possible for the anions lead to the three rules below.

1. When the anion ends in *–ide*, the acid name begins with the prefix *hydro-*. The root of the anion name goes in the blank (*chlor-* for chloride), followed by the suffix *–ic*. HCl is hydrochloric acid because Cl^- is the chloride ion. HCN is hydrocyanic acid because CN^- is the cyanide ion.
2. When the anion ends in *–ate*, the name of the acid is the root of the anion followed by the suffix *–ic*. There is no prefix. H_2SO_4 is sulfuric acid (not sulfic) because $SO_4{}^{2-}$ is the sulfate ion.
3. When the anion ends in *–ite*, the name of the acid is the root of the anion followed by the suffix *–ous*. Again, there is no prefix. HNO_2 is nitrous acid because $NO_2{}^-$ is the nitrite ion.

Note how the root for a sulfur-containing oxoacid is *sulfur-* instead of just *sulf-*. The same is true for a phosphorus-containing oxoacid. The root is *phosphor-* instead of simply *phosph-*.

Writing Formulas for Acids



Bases

The simplest way to define a **base** *is an ionic compound that produces hydroxide ions when dissolved in water*. One of the most commonly used bases is sodium hydroxide (**Figure** 16.2).

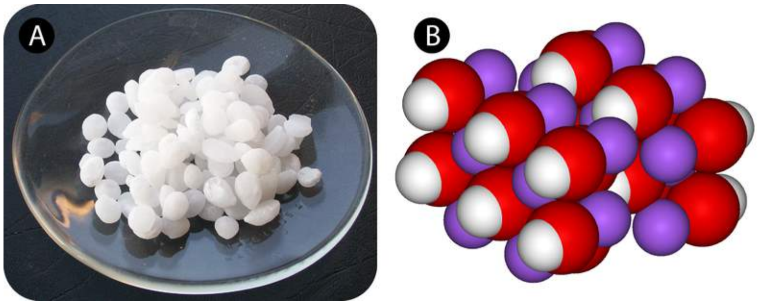


FIGURE 9.8

(A) Sodium hydroxide, a base, is a solid that is typically produced as small white pellets. (B) The structure of sodium hydroxide is an extended three-dimensional network. The purple spheres are the sodium ions (Na^+). The red and white spheres are oxygen and hydrogen atoms respectively, which are bonded together to form hydroxide ions (OH^-).

Names and Formulas of Bases

There is no special system for naming bases. Since they all contain the OH^- anion, names of bases end in hydroxide. The cation is simply named first. Some examples of names and formulas for bases are shown in **Figure** 16.2.

TABLE 9.10: **Examples of Bases**

Formula	Name
NaOH	sodium hydroxide
$Ca(OH)_2$	calcium hydroxide
NH_4OH	ammonium hydroxide

Notice that because bases are ionic compounds, the number of hydroxides in the formula does not affect the name. The compound must be neutral, so the charges of the ions are balanced just as for other ionic compounds. Sodium ion (Na^+) requires one OH^- ion to balance the charge, so the formula is NaOH. Calcium ion (Ca^{2+}) requires two OH^- ions to balance the charge, so the formula is $Ca(OH)_2$. Hydroxide ion is a polyatomic ion and must be put in parentheses when there are more than on in a formula.

Lesson Summary

- Acids are molecular compounds that dissolve in water to produce hydrogen ions and an anion. Bases are ionic compounds consisting of hydroxide ions combined with a cation.
- The three naming rules for acids are based on the suffix of the anion. Formulas for acids are written by balancing out the charge of the anion with the appropriate number of hydrogen ions.
- Naming and formula writing for bases follows the same guidelines as for other ionic compounds.

Lesson Review Questions

Reviewing Concepts

1. What ion must be produced when an acid dissolves in water?
2. What ion must be present for a compound to be considered a base?
3. Identify each of the following compounds as either an acid, a base, or neither.
   a. RbOH
   b. HBr
   c. MgO
   d. CH_4
   e. H_3PO_3
   f. $Sr(OH)_2$

Problems

4. Name the following acids.
   a. HF
   b. $HClO_2$
   c. H_2CrO_4
   d. H_2SO_3
   e. H_3PO_4
   f. $HClO_4$
   g. H_2S
   h. HNO_3
5. Write correct formulas for the following acids.
   a. carbonic acid
   b. hydroiodic acid
   c. chloric acid
   d. phosphorous acid
   e. oxalic acid
   f. hypochlorous acid
   g. hydrobromic acid
   h. permanganic acid
6. Write names or formulas for these compounds.
   a. LiOH
   b. $Mg(OH)_2$

   c. $Fe(OH)_3$
   d. nickel(II) hydroxide
   e. aluminum hydroxide
   f. silver hydroxide

7. Identify each compound below as either a (1) ionic compound, (2) molecular compound, (3) acid, or (4) base. Then, name it properly.
   a. HCN
   b. KNO_3
   c. N_2O_4
   d. $Sr(OH)_2$
   e. $HBrO_3$
   f. SO_3
   g. HCH_3COO
   h. CsOH
   i. $Pb_3(PO_4)_2$
   j. AsH_3

Further Reading / Supplemental Links

- Kristi Lew, *Acids and Bases* (*Essential Chemistry*). Chelsea House, 2008.
- Chris Oxlade, *Acids and Bases*. Heinemann-Raintree, 2007.
- An Introduction to Chemistry (http://preparatorychemistry.com/Bishop_acid_nomenclature_help.htm

Points to Consider

The extended three-dimensional structure of ionic compounds is commonly called a crystal lattice. The interactions of the ions within a crystal lattice lead directly to many important physical and chemical properties of ionic compounds.

- How is a crystal lattice formed?
- Why is a crystal lattice the most stable structural form for an ionic compound?
- What physical properties do ionic compounds have in common as a result of their structure?

Chm.1.2.4

- Write binary compounds of two nonmetals: use Greek prefixes (di-, tri-, tetra-, ...)
- Write binary compounds of metal/nonmetal*
-
- Write ternary compounds (polyatomic ions)*
- Write, with charges, these polyatomic ions: nitrate, sulfate, carbonate, acetate, and ammonium.
- Know names and formulas for these common laboratory acids: HCl, HNO3, H2SO4, HC2H3O2, (CH3COOH) *The Stock system is the correct IUPAC convention for inorganic nomenclature

Chm 1.2.2

- Determine that a bond is predominately covalent by the location of the atoms on the Periodic Table (nonmetals combined with
  nonmetals) or when $\Delta EN < 1.7$.

- Predict chemical formulas of compounds using Lewis structures.

Chm.1.2.4 Interpret the name and formula of compounds using IUPAC convention.

9.4 References

1. CK-12 Foundation - Christopher Auyeung. . CC-BY-NC-SA 3.0
2. Ben Mills (Wikimedia: Benjah-bmm27). http://commons.wikimedia.org/wiki/File:Sodium-chloride-3D-ionic.png . Public Domain
3. Ondřej Mangl. Copper(I) oxide: http://commons.wikimedia.org/wiki/File:Cu2O.png; Copper(II) oxide: http://commons.wikimedia.org/wiki/File:Oxid_m%C4%9B%C4%8Fnat%C3%BD.PNG . Public Domain
4. Ben Mills (Wikimedia: Benjah-bmm27). (A) http://commons.wikimedia.org/wiki/File:Ammonium-3D-balls.png; (B) http://commons.wikimedia.org/wiki/File:Hydroxide-3D-vdW.png; (C) http://commons.wikimedia.org/wiki/File:Carbonate-3D-balls.png . Public Domain
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8. (A) Martin Walker (Wikimedia: Walkerma); (B) Ben Mills (Wikimedia: Benjah-bmm27). (A) http://commons.wikimedia.org/wiki/File:SodiumHydroxide.jpg; (B) http://commons.wikimedia.org/wiki/File:Sodium-hydroxide-crystal-3D-vdW.png . Public Domain

CHAPTER 10 Covalent Bonding

Chapter Outline

Chm 1.2.2

- Determine that a bond is predominately covalent by the location of the atoms on the Periodic Table (nonmetals combined with nonmetals) or when ΔEN <1.7.
- Predict chemical formulas of compounds using Lewis structures.

Chm 1.2.3

- Explain why intermolecular forces are weaker than ionic, covalent or metallic bonds
- Explain why hydrogen bonds are stronger than dipole-dipole forces which are stronger than dispersion forces
- Apply the relationship between bond energy and length of single, double, and triple bonds (conceptual, no numbers).
- Describe intermolecular forces for molecular compounds.
- H-bond as attraction between molecules when H is bonded to O, N, or F. Dipole-dipole attractions between polar molecules. London dispersion forces (electrons of one molecule attracted to nucleus of another molecule) –i.e. liquefied inert gases. Relative strengths (H>dipole>London/van der Waals).

Chm 1.2.5

- Explain how covalent bonding in compounds determines their characteristics: low MP, low BP, poor electrical conductivity, polar nature, etc.
  - Apply Valence Shell Electron Pair Repulsion Theory (VSEPR) for these electron pair geometries and molecular geometries, and bond angles - Electron pair - Molecular (bond angle); Linear framework –linear; Trigonal planar framework–trigonal planar, bent; Tetrahedral framework–tetrahedral, trigonal pyramidal, bent; Bond angles (include distorting effect of lone pair electrons –no specific angles, conceptually only)
  - Describe bond polarity. Polar/nonpolar molecules (relate to symmetry) ; relate polarity to solubility—"like dissolves like"

• Describe macromolecules and network solids: water (ice), graphite/diamond, polymers (PVC, nylon), proteins (hair, DNA) intermolecular structure as a class of molecules with unique properties.

10.1 Molecular Compounds

- Define covalent compound.
- Explain how covalent compounds are named.
- Describe how covalent compounds differ from ionic compounds.

The burner on a gas stove burns with a pretty blue flame like the one pictured in the opening image. The fuel burned by most gas stoves is natural gas, which consists mainly of methane. Methane is a compound that contains only carbon and hydrogen. Like many other compounds that consist of just these two elements, methane is used for fuel because it burns very easily. Methane is an example of a covalent compound.

What Are Covalent Compounds?

Compounds that form from two or more nonmetallic elements, such as carbon and hydrogen, are called **covalent compounds**. In a covalent compound, atoms of the different elements are held together in molecules by covalent bonds. These are chemical bonds in which atoms share valence electrons. The force of attraction between the shared electrons and the positive nuclei of both atoms holds the atoms together in the molecule. A molecule is the smallest particle of a covalent compound that still has the properties of the compound.

The largest, most complex covalent molecules have thousands of atoms. Examples include proteins and carbohydrates, which are compounds in living things. The smallest, simplest covalent compounds have molecules with just two atoms. An example is hydrogen chloride (HCl). It consists of one hydrogen atom and one chlorine atom, as you can see in the **Figure** 10.1. You can watch an animation of hydrogen chloride forming at the following URL.

http://www.bbc.co.uk/schools/gcsebitesize/science/add_aqa_pre_2011/atomic/covalentrev1.shtml

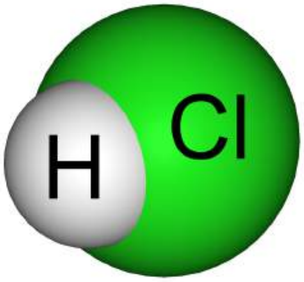


FIGURE 10.1

Naming and Writing Formulas for Covalent Compounds

To name simple covalent compounds, follow these rules:

- Start with the name of the element closer to the left side of the periodic table.
- Follow this with the name of element closer to the right of the periodic table. Give this second name the suffix *–ide*.
- Use prefixes to represent the numbers of the different atoms in each molecule of the compound. The most commonly used prefixes are shown in the **Table** 10.1.

TABLE 10.1: **Prefixes for Naming Covalent Compounds**

Number	Prefix
1	mono- (or none)
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

Q: What is the name of the compound that contains three oxygen atoms and two nitrogen atoms?

A: The compound is named dinitrogen trioxide. Nitrogen is named first because it is farther to the left in the periodic table than oxygen. Oxygen is given the *-ide* suffix because it is the second element named in the compound. The prefix *di-* is added to nitrogen to show that there are two atoms of nitrogen in each molecule of the compound. The prefix *tri-* is added to oxygen to show that there are three atoms of oxygen in each molecule.

In the chemical formula for a covalent compound, the numbers of the different atoms in a molecule are represented by subscripts. For example, the formula for the compound named carbon dioxide is CO_2.

Q: What is the chemical formula for dinitrogen trioxide?

A: The chemical formula is N_2O_3.

Properties of Covalent Compounds

The covalent bonds of covalent compounds are responsible for many of the properties of the compounds. Because valence electrons are shared in covalent compounds, rather than transferred between atoms as they are in ionic compounds, covalent compounds have very different properties than ionic compounds.

- Many covalent compounds, especially those containing carbon and hydrogen, burn easily. In contrast, many ionic compounds do not burn.
- Many covalent compounds do not dissolve in water, whereas most ionic compounds dissolve well in water.
- Unlike ionic compounds, covalent compounds do not have freely moving electrons, so they cannot conduct

TABLE 10.2: **Boiling Points of Some Common Compounds**

Name of Compound(Chemical Formula)	Type of Compound	Boiling Point (°C)
Methane (CH_4)	covalent	-164
Nitrogen oxide (NO)	covalent	-152
Sodium chloride (NaCl)	ionic	1413
Lithium fluoride (LiF)	ionic	1676

Q: The two covalent compounds in the table are gases at room temperature, which is 20°C. For a compound to be a liquid at room temperature, what does its boiling point have to be?

A: To be a liquid at room temperature, a covalent compound has to have a boiling point higher than 20°C. Water is an example of a covalent compound that is a liquid at room temperature. The boiling point of water is 100°C.

Summary

- Covalent compounds contain two or more nonmetallic elements held together by covalent bonds, in which atoms share pairs of valence electrons. A molecule is the smallest particle of a covalent compound that still has the properties of the compound.
- A few rules can be applied to name simple covalent compounds: name the left-most element in the periodic table first, add *-ide* to the second named element, and use prefixes for the numbers of atoms.
- Covalent bonds are responsible for many of the properties of covalent compounds. Covalent compounds have relatively low boiling points, cannot conduct electricity, and may not dissolve in water.

Vocabulary

- **covalent compound**: Compound that forms when atoms of nonmetals form molecules that are held together by covalent bonds.

Practice

Practice naming and writing chemical formulas for covalent compounds by completing the worksheet at the following URL. http://misterguch.brinkster.net/covalentname.pdf

Review

1. What are covalent compounds? Give two examples.
2. What is the name of the covalent compound that consists of two atoms of silicon and six atoms of bromine? What is its chemical formula.
3. Compare and contrast the properties of ionic and covalent compounds.

10.2 Covalent Bonding

- Define covalent bond.
- Describe covalent compounds and diatomic elements.
- Explain why covalent bonds form.

In a tennis match, two players keep hitting the ball back and forth. The ball bounces from one player to the other, over and over again. The ball keeps the players moving together on the court. What if the two players represented the nuclei of two atoms and the ball represented valence electrons? What would the back and forth movement of the ball represent? The answer is a covalent bond.

Sharing Electrons

A **covalent bond** is the force of attraction that holds together two atoms that share a pair of valence electrons. The shared electrons are attracted to the nuclei of both atoms. This forms a molecule consisting of two or more atoms. Covalent bonds form only between atoms of nonmetals.

Covalent Compounds and Diatomic Elements

The two atoms that are held together by a covalent bond may be atoms of the same element or different elements. When atoms of different elements form covalent bonds, a new substance, called a covalent compound, results. Water is an example of a covalent compound. A water molecule is modeled below. A molecule is the smallest particle of a covalent compound that still has the properties of the compound.

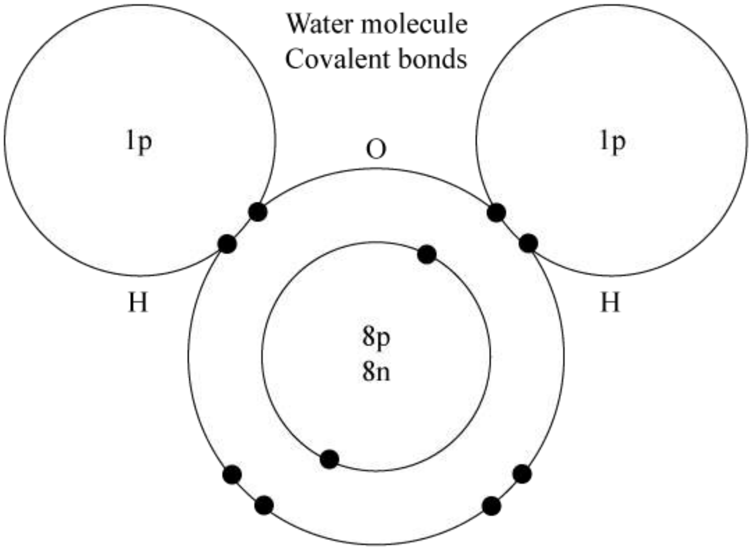


FIGURE 10.2

Q: How many valence electrons does the oxygen atom (O) share with each hydrogen atom (H)? How many covalent bonds hold the water molecule together?

A: The oxygen atom shares one pair of valence electrons with each hydrogen atom. Each pair of shared electrons represents one covalent bond, so two covalent bonds hold the water molecule together.

The diagram below shows an example of covalent bonds between two atoms of the same element, in this case two atoms of oxygen. The diagram represents an oxygen molecule, so it's not a new compound. Oxygen normally occurs in diatomic ("two-atom") molecules. Several other elements also occur as diatomic molecules: hydrogen, nitrogen, and all but one of the halogens (fluorine, chlorine, bromine, and iodine). At the URL below, you can see an animation of electron sharing in a molecule of chlorine.

http://visservices.sdsc.edu/projects/discovery/ClCl_SD.mov

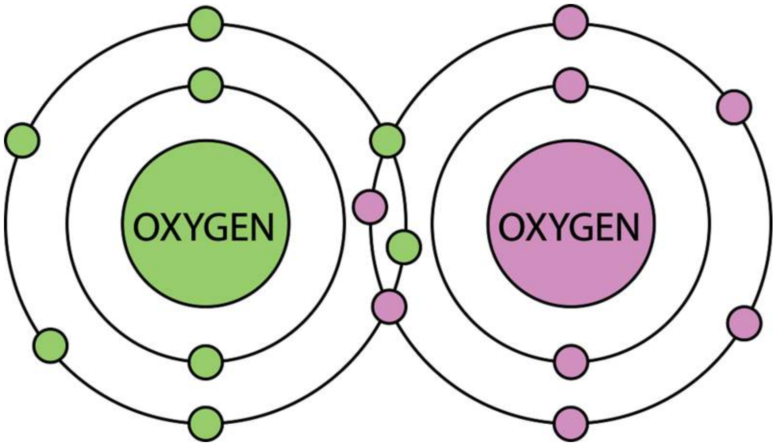


FIGURE 10.3

Q: How many electrons do these two oxygen atoms share? How many covalent bonds hold the oxygen molecule together?

A: The two oxygen atoms share two pairs of electrons, so two covalent bonds hold the oxygen molecule together.

Why Covalent Bonds Form

Covalent bonds form because they give atoms a more stable arrangement of electrons. Look at the oxygen atoms above. Alone, each oxygen atom has six valence electrons. By sharing two pairs of valence electrons, each oxygen atom has a total of eight valence electrons. This fills its outer energy level, giving it the most stable arrangement of electrons. The shared electrons are attracted to both oxygen nuclei, and this force of attraction holds the two atoms together in the oxygen molecule.

Summary

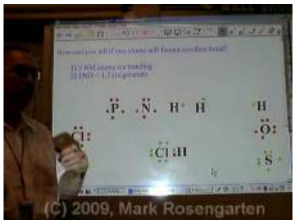
- A covalent bond is the force of attraction that holds together two atoms that share a pair of valence electrons. Covalent bonds form only between atoms of nonmetals.
- The two atoms that are held together in a covalent bond may be atoms of the same element or different elements. When atoms of different elements bond together, it forms a covalent compound.
- Covalent bonds form because the shared electrons fill each atom's outer energy level and this is the most stable arrangement of electrons.

Vocabulary

- **covalent bond**: Force of attraction that holds together two atoms that share a pair of electrons.

Practice

Watch the video about covalent bonding at the following URL, and then answer the questions below. http://www.youtube.com/watch?v=-Eh_0Dseg3E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MEDIA

Click image to the left for more content.

1. How can you tell the number covalent bonds the atoms of an element can form?
2. How many covalent bonds can nitrogen form? How many covalent bonds can chlorine form?
3. Why are covalent bonds stronger than ionic bonds?
4. Why can't molecules of a covalent compound conduct electricity?

Review

1. What is a covalent bond?
2. Nitrogen is a diatomic element with five valence electrons. Create a model of a molecule of nitrogen.
3. Which of the following represent molecules? Which represent compounds? Which represents a covalent compound?
    a. O_2
    b. CO_2
    c. Cl_2
    d. NaCl

10.3 The Covalent Bond and Lewis Structures

Lesson Objectives

The student will:

- explain what covalent bonds are.
- explain why covalent bonds are formed.
- compare covalent bonds with ionic bonds in terms of how their definitions and how they are formed.
- draw Lewis structures for simple covalent molecules.
- use Lewis structures to show the formation of single, double, and triple covalent bonds.
- identify pairs of atoms that will form covalent bonds.
- define coordinate covalent bond.
- explain the equivalent bond strengths in resonance structures.

Vocabulary

bond energy
the energy required to break a given chemical bond

bond length
the distance between the nuclei of the two atoms connected by a bond

coordinate covalent bonds
a type of covalent bond when the two shared electrons of a covalent bond are both donated by the same atom

covalent bond
a type of chemical bond where two atoms are connected to each other by the sharing of two or more electrons in overlapped orbitals

double bond
a bond in which two pairs of electrons are shared by two nuclei

pi bond
a covalent bond in which **p** orbitals share an electron pair occupying the space above and below the line joining the atoms' nuclei

resonance
a condition occurring when more than one valid Lewis structure must be written for a particular molecule; the actual electronic structure is not represented by any one of the Lewis structures, but rather by the average of all of them

sigma bond
a covalent bond in which the electron pair is shared in an area centered along a line running between the atoms' nuclei

triple bond
a bond in which three pairs of electrons are shared by two nuclei

Introduction

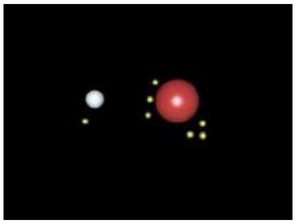
As we saw in the chapter "Ionic Bonds and Formulas," metallic atoms can transfer one or more electrons to nonmetallic atoms, producing positively charged cations and negatively charged anions. The attractive force between these oppositely charged ions is called an ionic bond. However, chemical bonding does not always require the complete transfer of electrons from one atom to another. When a bond forms between two nonmetallic atoms, neither has a low enough ionization energy to completely give up an electron to its partner. Instead, the atoms overlap their orbitals, and the electrons in these shared orbitals can be considered to be in "belong" to both atoms at the same time. These atoms are now in a **covalent bond**, held together by the attraction of both nuclei to the shared electrons.

Ionic versus Covalent Bonding

The way that atoms bind together is due to a combination of factors: the electrical attraction and repulsion between atoms, the arrangement of electrons in atoms, and the natural tendency for matter to achieve the lowest potential energy possible. In most cases, these factors favor atoms that have obtained a complete octet of valence electrons. In ionic bonding, the atoms acquired this octet by gaining or losing electrons, while in covalent bonding, the atoms acquire the noble gas electron configuration by sharing electrons.

As you may recall from the discussion of ionic bonds in the chapter "Ionic Bonds and Formulas," ionic bonds form between metals and nonmetals. Nonmetals take electrons away from metals and the resulting oppositely charged metal and nonmetal ions are then attracted to each other. In covalent bonds, electrons are shared, meaning that metals will form few, if any, covalent bonds. Metals do not hold on to electrons with enough strength to participate in covalent bonding. For a covalent bond to form, we need two atoms that both attract electrons strongly, or two atoms with high electronegativity. Hence, the great majority of covalent bonds will form between two nonmetals. When both atoms in a bond are from the upper right side of the periodic table, the bond is likely to be covalent.

An animation showing ionic and covalent bonding **(2a)** is available at http://www.youtube.com/watch?v=QqjcCvzWwww (1:57).



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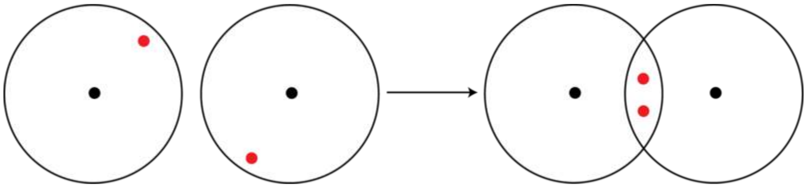
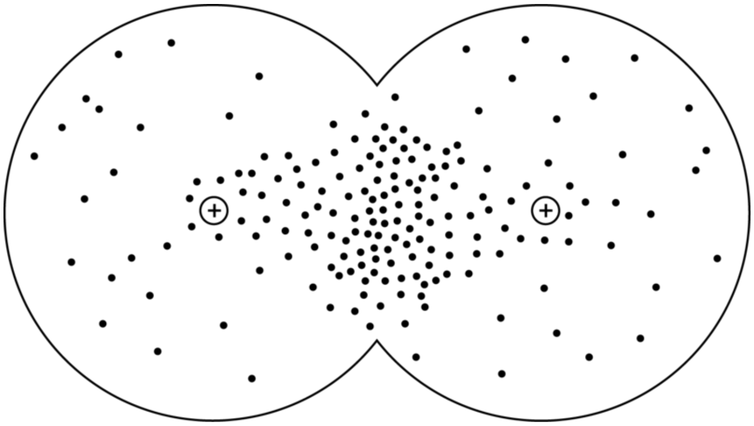
Single Bond

In covalent bonding, the atoms acquire a stable octet of electrons by sharing electrons. The covalent bonding produces molecular substances, as opposed to the crystal lattice structures produced by ionic bonding. The covalent bond, in general, is much stronger than the ionic bond, and there are far more covalently bonded substances than ionically bonded ones.

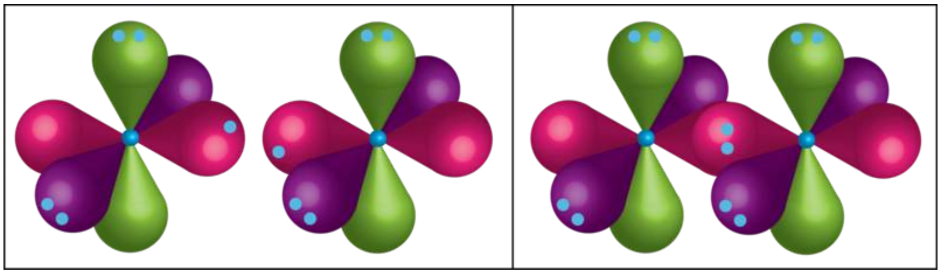
The simplest covalent bond is made between two atoms of hydrogen. Each hydrogen atom has one electron in its $1s$ orbital. When two hydrogen atoms approach one another and their orbitals overlap, it creates a common space between the two nuclei that both electrons occupy (as seen in the figure below). Since these electrons are shared, both hydrogen atoms now have a full valence shell.



Since electrons are constantly in motion, they will not always be directly in the center of a covalent bond. However, the simulated probability pattern below shows that they do spend the majority of their time in the area directly between the two nuclei. The extra time spent between the two nuclei is the source of the attraction that holds the atoms together in a covalent bond.

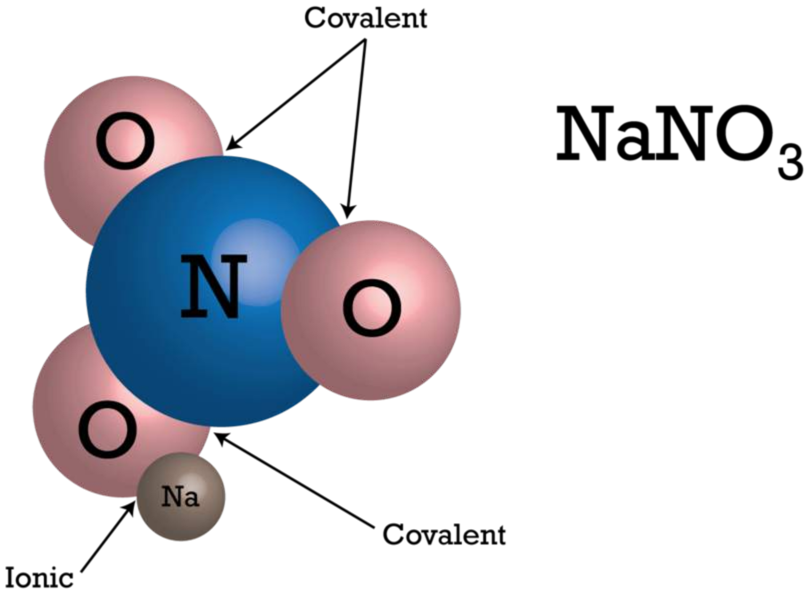


Another example of a covalent bond is found in the diatomic F_2 molecule (seen below). Recall that fluorine atoms have 7 valence electrons. Two are placed in the 2*s* orbital, and five are placed in the three 2*p* orbitals. Since each orbital can hold two electrons, this means that one of the 2*p* orbitals is only half full, so it is available for bonding. When two fluorine atoms interact, their half-filled orbitals overlap, creating a covalent bond that gives both atoms a complete octet of valence electrons.



Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction to polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds, and the nitrate ion is attached to the sodium ion by an ionic bond (see sketch below).



Bond Strength

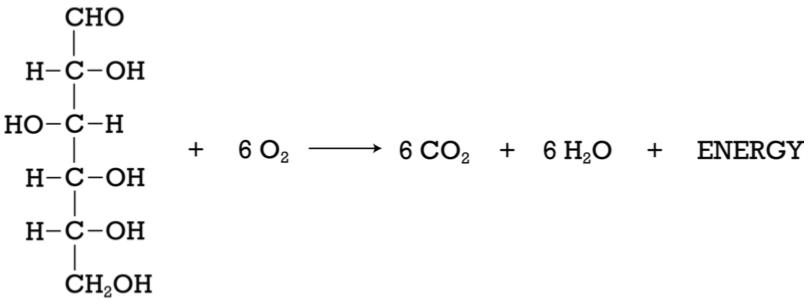
When atoms that attract each other move closer together, the potential energy of the system (the two atoms) decreases. When a covalent bond is formed, the atoms move so close together that their electron clouds overlap. As the atoms continue to move closer yet, the potential energy of the system continues to decrease –to a point. If you continue to move atoms closer and closer together, eventually the two nuclei will begin to repel each other. If you push the nuclei closer together beyond this point, the repulsion causes the potential energy to increase. Each pair of covalently bonding atoms will have a distance between their nuclei that is the lowest potential energy distance. This position has the atoms close enough for the attraction between the nucleus of one atom and the electrons of the other is maximum, but not so close that the nuclei have not begun to repel each other strongly. This distance is called the bond length. The more potential energy released as this bond formed, the stronger the bond will be. In order to break this bond, you must input an equivalent amount of energy.

The strength of a diatomic covalent bond can be expressed by the amount of energy necessary to break the bond and produce separate atoms. The energy needed to break a covalent bond is called **bond energy** and is measured in kilojoules per mole. Bond energy is used as a measure of bond strength. The bond strength of HBr is 365 kilojoules per mole, meaning that it would take 365 kilojoules to break all the chemical bonds in 1 mole, or 6.02×10^{23} molecules, of HBr and produce separate hydrogen and bromine atoms. In comparison, the bond strength of HCl is 431 kilojoules per mole. Consequently, the bond in HCl is stronger than the bond in HBr.

Molecular Stability

The bond energy can be used to indicate molecular stability. When stable compounds are formed, large amounts of energy are given off. In order to break the molecule apart, all the energy that was given off when the bonds were formed must be put back in to break the bonds. The more energy needed to break a bond, the more stable the compound is. Therefore, compounds with higher bond strengths tend to be more stable.

The molecule of glucose, shown below, can react with six molecules of oxygen to produce six molecules of carbon dioxide and six molecules of water. During the reaction, the atoms of the glucose molecule are rearranged into the structures of carbon dioxide and water. The bonds in glucose are broken, and new bonds are formed. As this occurs, potential energy is released because the new bonds have lower potential energy than the original bonds. Since the bonds in the products are lower energy bonds, the product molecules are more stable than the reactants.

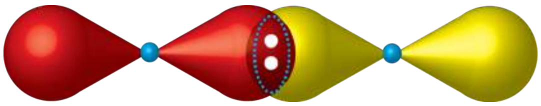


Double and Triple Bonds

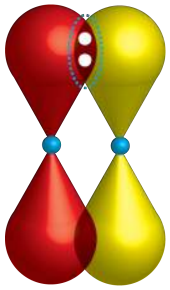
In the previous examples, only one pair of electrons was shared between the two bonded atoms. This type of bond is called a single bond. However, it is possible for atoms to share more than one pair of electrons. When two or three pairs of electrons are shared by two bonded atoms' nuclei, they are referred to as **double bonds** or **triple bonds**, respectively.

A double bond is formed when two pairs of orbitals overlap with each other at the same time. The O_2 molecule provides an example of a double bond. Oxygen has six valence electrons. Two electrons are placed in the 2 *s* orbital, while the remaining four are distributed among the three 2*p* orbitals. Hund's rule tells us that each *p* orbital will receive one electron before a second electron is added, so the oxygen atom will have two half-filled orbitals available for bonding (a filled orbital will not participate in bonding).

Of the two *p* orbitals available for bonding, the first pair will have a head-to-head overlap, placing the shared electrons directly in between the two nuclei. This type of a bond, shown in the figure below, is called a **sigma (σ) bond**.



The second pair of half-filled orbitals is orientated perpendicularly to the first pair, so a similar head-to-head overlap is not possible. However, a bond can still be made by overlapping the two *p* orbitals side-by-side, as shown in the figure below. This type of bond is called a **pi (π) bond**.



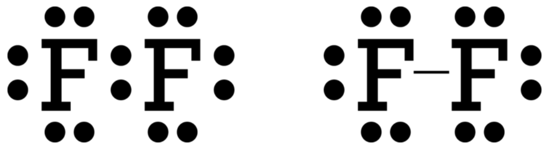
Whenever possible, the first bond will form directly between the two atomic nuclei involved in bonding. This allows for maximum overlap between the two orbitals, helping to minimize the electrostatic repulsion between the two positively charged nuclei. Thus, single bonds consist of one sigma bond, double bonds consist of one sigma and one pi bond, and triple bonds consist of one sigma and two pi bonds.

Note that double bonds are less than twice as strong as a single bond between the same two atoms. Since sigma bonds are "better" (stronger) than pi bonds, a combination of one sigma and one pi is slightly weaker than adding two sigma bonds would be.

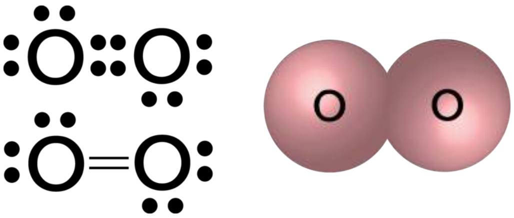
Lewis Dot Structures

It would not be very efficient if we had to draw an orbital picture every time we wanted to describe a molecule. Lewis dot structures, first developed by G.N. Lewis, are a shorthand way of drawing the arrangement of atoms, bonds, and valence electrons in a molecule. In the earlier chapter "The Electron Configuration of Atoms," we introduced Lewis dot diagrams for drawing individual atoms. When we draw molecules, the diagrams are known as Lewis dot diagrams, Lewis structures, or Lewis formulas. The Lewis structures of a molecule show how the valence electrons are arranged among the atoms of the molecule.

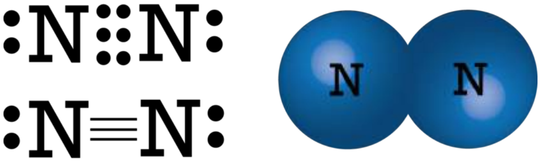
In a Lewis structure, each valence electron is represented by a dot, and bonds are shown by placing electrons in between the symbols for the two bonded atoms. Often, a pair of bonding electrons is further abbreviated by a dash. For example, we can represent the covalent bond in the F_2 molecule by either of the Lewis structures shown below.



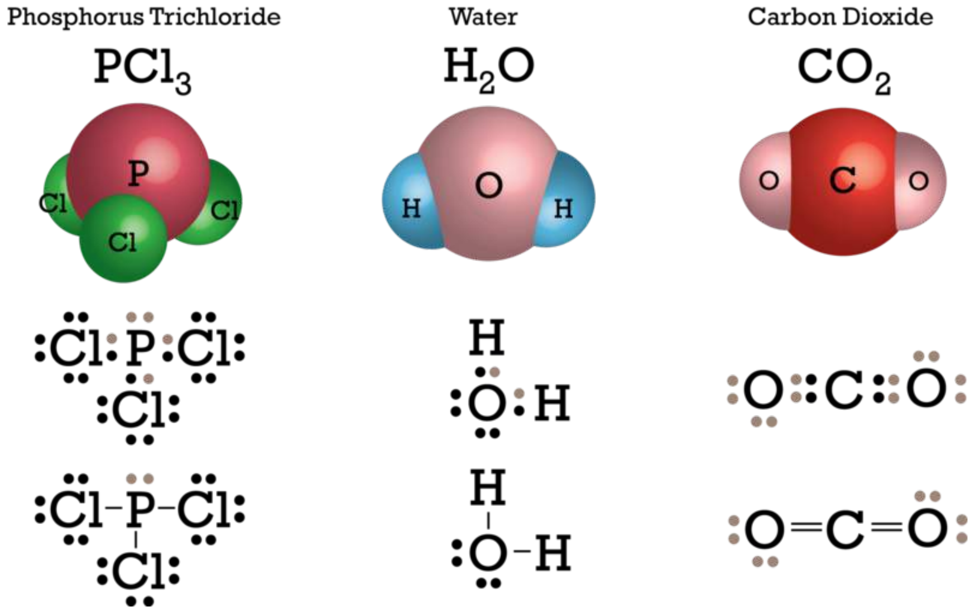
Double bonds (4 electrons shared between two atoms) can be represented either with 4 dots or 2 dashes. The Lewis structure for an oxygen molecule (O_2) is shown below.



Similarly, triple bonds can be written as 6 dots or 3 dashes. An example of a molecule with triple bonds is the nitrogen molecule, N_2. The Lewis structure for a nitrogen molecule can be represented by either of the two ways shown below. It is important to keep in mind that a dash always represents a pair of electrons. .



Several other examples of representing covalent bonds are shown in the figure below.



Drawing Lewis Structures

The rules outlined below for writing Lewis structures are based on observations of thousands of molecules (note that there will be some exceptions to the rules). We learned earlier that atoms are generally most stable with 8 valence electrons (the octet rule). The major exception is hydrogen, which requires only 2 valence electrons to have a complete valence shell (sometimes called the duet rule). Lewis structures allow one to quickly assess whether each atom has a complete octet of electrons(or duet for hydrogen.)

Rules for Writing Lewis Structures:

a. Decide the skeletal arrangement of the atoms.
b. Count up all the valence electrons of all the atoms.
c. Place two electrons between each pair of bonded atoms.
d. Complete all the octets (or duets) of the atoms attached to the central atom.
e. Place any remaining electrons on the central atom.
f. If the central atom does not have an octet, bring in lone pairs to form double or triple bonds.

Example:

Write the Lewis structure for water, H_2O.

Step 1: Decide which atoms are bonded.

Begin by assuming the hydrogen atoms are bounded to the oxygen atom. In other words, assume the oxygen atom is the central atom. Hydrogen is NEVER the central atom.

$$H-O-H$$

Step 2: Count all the valence electrons of all the atoms.

The oxygen atom has 6 valence electrons, and each hydrogen has 1. The total number of valence electrons is 8.

Step 3: Place two electrons between each pair of bounded atoms.

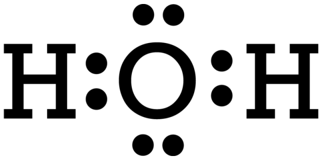
$$H : O : H$$

Step 4: Complete all the octets or duets of the atoms attached to the central atom.

The hydrogen atoms are attached to the central atom and can only have a duet of electrons.

Step 5: Place any remaining electrons on the central atom.

The total number of valence electrons is 8, and we have already used 4 of them. The other 4 will fit around the central oxygen atom as lone pairs.



Is this structure correct?

- Are the total number of valence electrons correct? Yes
- Does each atom have the appropriate duet or octet of electrons? Yes

The structure, then, is correct.

Example:

Write the Lewis structure for carbon dioxide, CO_2.

Step 1: Decide which atoms are bonded. The less electronegative element usually goes in the center.

Begin by assuming the carbon is the central atom and that both oxygen atoms are attached to the carbon.

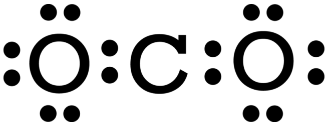
Step 2: Count all the valence electrons of all the atoms.

The oxygen atoms each have 6 valence electrons and the carbon atom has 4. The total number of valence electrons is 16.

Step 3: Place two electrons between each pair of bounded atoms.

O : C : O

Step 4: Complete all the octets or duets of the atoms attached to the central atom.



Step 5: Place any remaining electrons on the central atom.

We have used all 16 of the valence electrons so there are no more to place around the central carbon atom.

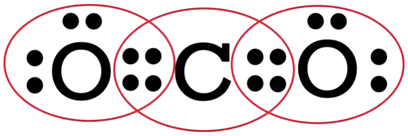
Is this structure correct?

- Is the total number of valence electrons correct? Yes
- Does each atom have the appropriate duet or octet of electrons? NO

Each oxygen has the proper octet of electrons, but the carbon atom only has 4 electrons. Therefore, this structure is *not* correct.

Step 6: If the central atom does not have an octet, move lone pairs in to form double or triple bonds.

Double bonds can be formed between carbon and each oxygen atom.



Notice this time that each atom is surrounded by 8 electrons.

Example:

Write the Lewis structure for ammonia, NH_3.

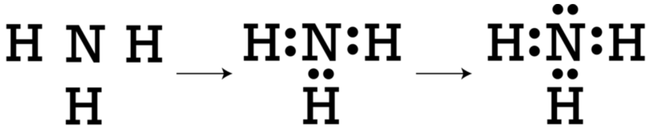
Step 1: Decide which atoms are bonded.

Place nitrogen as the central atom and each hydrogen bonded to the nitrogen.

Step 2: Count all the valence electrons of all the atoms.

The nitrogen atom has five valence electrons, and each hydrogen atom has one. The total number of valence electrons is 8.

Step 3: Place two electrons between each pair of bonded atoms, as seen in the middle drawing of the figure below.



Step 4: Complete all the octets or duets of the atoms attached to the central atom.

In this case, all the non-central atoms are hydrogen, and they can only have a duet of electrons.

Step 5: Place any remaining electrons on the central atom.

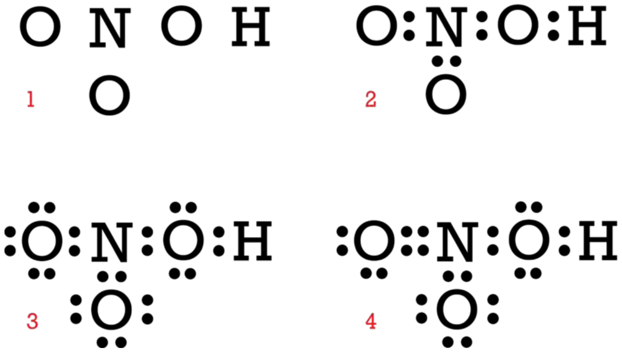
There are still two electrons left, so they would complete the octet for nitrogen. If the central atom, at this point, does not have an octet of electrons, we would look for places to create a double or triple bond, but that is not the case here. The final drawing on the right in the figure above is the Lewis structure for ammonia.

Example:

Write the Lewis structure for nitric acid, HNO_3.

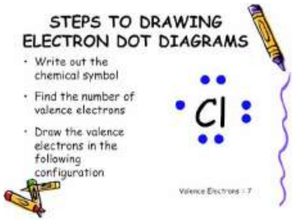
Solution:

The skeleton for nitric acid has the three oxygen atoms bonded to the nitrogen and the hydrogen bonded to one of the oxygen atoms, as seen in diagram 1 shown below. The total number of valence electrons is $5+6+6+6+1=24$.



The next step, shown in diagram 2, is to put in a pair of electrons between each bonded pair. So far, we have accounted for 8 of the 24 valence electrons. The next step is to complete the octet or duet for each of the non-central atoms, as seen in diagram 3. At that point, we have used all of the valence electrons, but the central atom does not have an octet of electrons. The rules tell us to find a place to put a double or triple bond. Based on what we have learned up to this point, any of the three oxygen atoms is just as good as the others for participating in a double bond. For this example, we moved two of the electrons from the oxygen atom on the far left between the oxygen and nitrogen. Now every atom in the molecule has the appropriate octet or duet of electrons. We have a satisfactory Lewis structure for the nitric acid molecule.

For an introduction to drawing Lewis electron dot symbols **(2e)**, see http://www.youtube.com/watch?v=y6QZRBIO0-o (4:19).



STEPS TO DRAWING ELECTRON DOT DIAGRAMS

- Write out the chemical symbol
- Find the number of valence electrons
- Draw the valence electrons in the following configuration

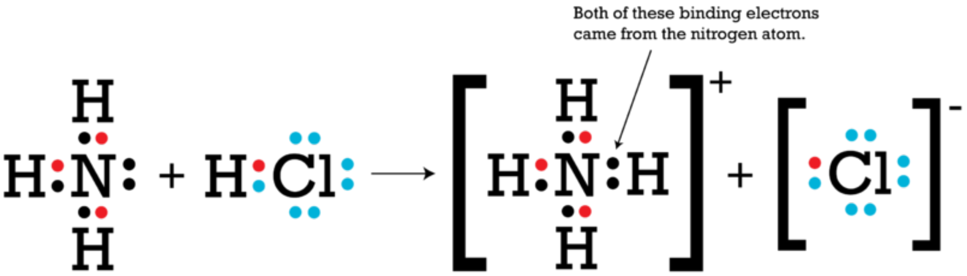
MEDIA

Click image to the left for more content.

Coordinate Covalent Bond

A variation of covalent bonding is coordinate covalent bonding. **Coordinate covalent bonds** form when the two shared electrons of a covalent bond are both donated by the same atom. So far, we have looked at covalent bonds formed by the overlap of two half-filled orbitals. A coordinate covalent bond is different in that it involves the overlap of one full orbital and one empty one. Once formed, a coordinate covalent bond is no different from an ordinary covalent bond. The difference is simply in the source of the electrons forming the bond. You should note that the bond still involves only one pair of electrons and one pair of orbitals, but only one atom provides both of the shared electrons. Many polyatomic ions include coordinate covalent bonds. The ammonium ion is an example of this type of bonding. When hydrogen chloride gas and ammonia gas are brought into contact with each other, they form ammonium chloride, NH_4Cl. When ammonium chloride is dissolved in water, ammonium ions and chloride ions are produced.

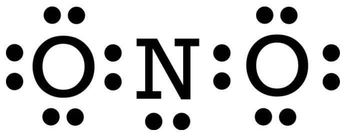
$$HCl + NH_3 \rightarrow NH_4Cl + H_2O \rightarrow NH_{4\ (aq)}^{+} + Cl_{(aq)}^{-}$$



In one of the nitrogen-hydrogen bonds, both electrons came from the nitrogen atom. However, once the ammonium ion has been formed, all four nitrogen-hydrogen bonds are identical, regardless of where the electrons in the bonds came from.

Resonance

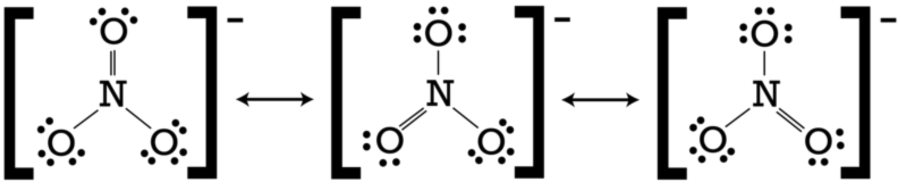
Sometimes, more than one valid Lewis structure is possible for the same molecule. Consider the Lewis structure for the nitrite ion, NO_2^-. The charge on the ion indicates an electron has been gained from an external source. In other words, the ion contains an electron that did not originally belong to the nitrogen or the oxygen atoms in the ion. In this case, the -1 charge indicates that the valence electron count will have one additional electron added in order to account for the electron that came from outside the ion. The total number of valence electrons for this ion is $5+6+6+1=18$. We can then draw the Lewis structure for this ion following the normal rules. Before applying the last rule of creating double or triple bonds, the incomplete Lewis structure will look like that in the figure below.



All 18 of the available valence electrons have been used. The final rule for writing Lewis structures states that if the central atom does not have an octet of electrons, double or triple bonds need to be created. In this case, either one of the nitrogen-oxygen bonds can be made into a double bond, as demonstrated in the figure below.

The two structures in the image above suggest that one of the nitrogen-oxygen bonds should be shorter and stronger than the other one. It has been experimentally shown, however, that the two nitrogen-oxygen bonds are identical. In fact, the nitrogen-oxygen bond lengths are about halfway between the expected values for a single bond and a double bond. Neither of the Lewis structures above matches the experimental evidence, and there is no good way to write one that does. The solution to this problem is the use of a concept called **resonance.** Resonance is the condition where there is more than one valid Lewis structure can and must be written for the molecule or ion. The actual structure of the molecule or ion is actually a composite, or average, of all the valid Lewis structures. In the case of the nitrite ion above, the second pair of electrons in the double bond is actually shared between all the atoms, giving each bond the strength of about 1.5 bonds. Each of the Lewis structures that is drawn is called a resonance structure, with the actual structure being a resonance hybrid of all the contributing structures.

Resonance is also present in the Lewis structures for nitrate ion, NO_3^-.



All of these are valid Lewis structures, but none of them alone accurately portray the structure of the ion. Just like in the case of nitrite, the actual nitrate ion is an average of all three structures. All three nitrogen-oxygen bonds are identical, and they are all slightly shorter and stronger than a normal nitrogen-oxygen single bond. We saw earlier that the second bond in a double bond is a pi bond, created by two *p* orbitals lining up side by side. It is those pi electrons that cannot be described well by a single Lewis structure. Rather than creating a true pi bond between just two atoms, the extra electrons are shared among all four atoms.

When the concept of resonance was first introduced, it was thought that the molecule was rapidly switching, or resonating, between the various resonance forms. Although later evidence showed that this is not the case, the term has survived. Note the double headed arrows ($\longleftrightarrow$) between each structure. Using this type of arrow indicates that the structures shown are resonance structures, implying that the entire figure attempts to describe a single molecule or ion.

Lesson Summary

- Covalent bonds are formed by electrons being shared between two atoms.

- Covalent bonds are formed between atoms with relatively high electronegativity.
- Bond energy is the amount of energy necessary to break the covalent bond and is an indication of the strength and stability of the bond.
- Some atoms are capable of forming double or triple bonds.
- Multiple bonds between atoms require multiple half-filled orbitals.
- End-to-end orbital overlaps are called sigma bonds.
- Side-to-side orbital overlaps are called pi bonds.
- Lewis structures are commonly used to show the valence electron arrangement in covalently bonded molecules.
- Resonance is a condition occurring when more than one valid Lewis structure can be written for a particular molecule. The actual electronic structure is not represented by any one of the Lewis structures but by the average of all of the valid structures.

Further Reading / Supplemental Links

The learner.org website allows users to view the Annenberg series of chemistry videos. You are required to register before you can watch the videos, but there is no charge to register. The video called "Chemical Bonds" explains the differences between ionic and covalent bonds using models and examples from nature.

- http://www.learner.org/vod/vod_window.html?pid=800

This animation explores the differences between ionic and covalent bonding.

- http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/bom1s2_11.swf

This website provides more information about the different types of chemical bonding.

- http://www.visionlearning.com/library/module_viewer.php?mid=55

The website below provides a guide to drawing Lewis structures.

- http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch8/lewis.html#step

Review Questions

1. Describe the characteristics of two atoms that would be expected to form an ionic bond.
2. Describe the characteristics of two atoms that would be expected to form a covalent bond.
3. If a molecule had a very high bond energy, would you expect it to be stable or unstable?
4. When gaseous potassium ions and gaseous fluoride ions join together to form a crystal lattice, the amount of energy released is 821 kJ/mol. When gaseous potassium ions and gaseous chloride ions join together to form a crystal lattice, the amount of energy released is 715 kJ/mol. Which bond is stronger, KF or KCl? If these two compounds were increasingly heated, which compound would break apart at the lower temperature?
5. Of the following compounds listed in **Table** 10.3, which would you expect to be ionically bonded and which would you expect to be covalently bonded?

TABLE 10.3: **Table for Review Question 5**

Compound	Ionic or Covalent
CS_2	
K_2S	
FeF_3	

TABLE 10.3: (continued)

Compound	Ionic or Covalent
PF_3	
BF_3	
AlF_3	
BaS	

1. How many sigma bonds and how many pi bonds are present in a triple bond?
2. Which of the following molecules will have a triple bond?
   a. C_2H_2
   b. CH_2Cl_2
   c. BF_3
   d. CH_3CH_2OH
   e. HF
3. Draw the Lewis structure for CCl_4.
4. Draw the Lewis structure for SO_2.
5. Draw a Lewis structure for CO_3^{2-}. Explain why all three carbon-oxygen bonds have the same length.

10.4 References

1. Ben Mills (Wikimedia: Benjah-bmm27). Structure of HCl. Public Domain
2. Zachary Wilson. Electron sharing in water molecule. CC BY-NC 3.0
3. Laura Guerin. Electron sharing in oxygen molecule. CC BY-NC 3.0

CHAPTER 11 Chemical Reactions

Chapter Outline

Chm.2.2 Analyze chemical reactions in terms of quantities, product formation, and energy.

Chm.2.2.1 Explain the energy content of a chemical reaction. Chm.2.2.2 Analyze the evidence of chemical change. Chm.2.2.3 Analyze the law of conservation of matter and how it applies to various types of chemical equations (synthesis, decomposition, single replacement, double replacement, and combustion).

11.1 Chemical Changes and Chemical Equations

- Describe the general form of a chemical equation.
- Show how to write a simple chemical equation.
- Explain what it means for a chemical equation to be balanced.

Look at this rusty bike wheel. It has been left outside in damp weather too many times, so the iron in the metal parts has rusted. Iron rusts when it combines with oxygen in the air. Iron rusting is an example of a chemical change. In a **chemical change**, substances change into entirely different substances. When a chemical change occurs, the particles or atoms break their bonds and rearrange themselves into a different substance. For example, the iron in the bike and the oxygen in the air have changed into rust.

Q: How could you represent this reaction, besides just describing it in words?

A: Scientists use a standard method to represent a chemical reaction, called a chemical equation.

What Is a Chemical Equation?

A **chemical equation** is a shorthand way to sum up what occurs in a chemical reaction. The general form of a chemical equation is:

Reactants → Products

The **reactants** in a chemical equation are the substances that begin the reaction, and the **products** are the substances that are produced in the reaction. The reactants are always written on the left side of the equation and the products on the right. The arrow pointing from left to right shows that the reactants change into the products during the reaction. This happens when chemical bonds break in the reactants and new bonds form in the products. As a result, the products are different chemical substances than the reactants that started the reaction.

Q: What is the general equation for the reaction in which iron rusts (hint: Think back to the first paragraph.)?

A: Iron combines with oxygen to produce rust, which is the compound named iron oxide. This reaction could be represented by the general chemical equation below. Note that when there is more than one reactant, they are separated by plus signs (+). If more than one product were produced, plus signs would be used between them as well.

Iron + Oxygen → Iron Oxide

Using Chemical Symbols and Formulas

When scientists write chemical equations, they use chemical symbols and chemical formulas instead of names to represent reactants and products. Look at the chemical reaction illustrated below. In this reaction, carbon reacts with oxygen to produce carbon dioxide. Carbon is represented by the chemical symbol C. The chemical symbol for oxygen is O, but pure oxygen exists as diatomic ("two-atom") molecules, represented by the chemical formula O_2. A molecule of the compound carbon dioxide consists of one atom of carbon and two atoms of oxygen, so carbon dioxide is represented by the chemical formula CO_2.

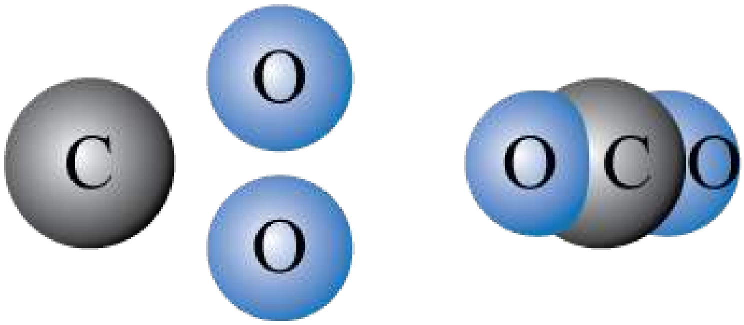


FIGURE 11.1

Q: What is the chemical equation for this reaction?

A: The chemical equation is:

$C + O_2 \rightarrow CO_2$

Q: How have the atoms of the reactants been rearranged in the products of the reaction? What bonds have been broken, and what new bonds have formed?

A: Bonds between the oxygen atoms in the oxygen molecule have been broken, and new bonds have formed between the carbon atom and the two oxygen atoms.

Is It Balanced?

All chemical equations, like equations in math, must balance. This means that there must be the same number of each type of atom on both sides of the arrow. That's because matter is always conserved in a chemical reaction. This is the law of conservation of mass.

Look at the equation above for the reaction between carbon and oxygen in the formation of carbon dioxide. Count the number of atoms of each type. Are the numbers the same on both sides of the arrow? The answer is yes, so the equation is balanced.

Summary

- Scientists use chemical equations to summarize what happens in chemical reactions. Reactants are placed on the left side of the equation and products are placed on the right. An arrow is used to indicate the direction in which the reaction occurs. Plus signs (+) are placed between multiple reactants or products.
- Chemical equations must be balanced. A balanced equation has the same number of each type of atom on both sides of the equation.

Vocabulary

- **chemical equation**: Symbolic representation of a chemical reaction
- **chemical change**: When atoms or particles break their bonds and form a new type of matter
- **reactants**: The starting material in a chemical reaction
- **products**: The substances formed during the chemical reaction

Review

1. What is a chemical equation? Identify the parts of a chemical equation.
2. Write a chemical equation for the chemical reaction in which calcium carbonate ($CaCO_3$) produces calcium oxide (CaO) and carbon dioxide (CO_2).
3. When is a chemical equation balanced?

Review the basic chemistry with these matching games at:

1. http://www.neok12.com/quiz/ATOM0001

- Chemical change
- Chemical equation
- reactants
- products

11.2 Chemical Reactions

Lesson Objectives

- Describe what happens in a chemical reaction, and identify types of chemical reactions.
- Explain the role of energy in chemical reactions, and define activation energy.
- State factors that affect the rate of chemical reactions.
- Explain the importance of enzymes in organisms, and describe how enzymes work.

Introduction

A chemical compound may be very different from the substances that combine to form it. For example, the element chlorine (Cl) is a poisonous gas, but when it combines with sodium (Na) to form sodium chloride (NaCl), it is no longer toxic. You may even eat it on your food. Sodium chloride is just table salt. What process changes a toxic chemical like chlorine into a much different substance like table salt?

What are Chemical Reactions?

A **chemical reaction** is a process that changes some chemical substances into other chemical substances. The substances that start a chemical reaction are called **reactants**. The substances that form as a result of a chemical reaction are called **products**. During the reaction, the reactants are used up to create the products. For example, when methane burns in oxygen, it releases carbon dioxide and water. In this reaction, the reactants are methane (CH_4) and oxygen (O_2), and the products are carbon dioxide (CO_2) and water (H_2O).

Chemical Equations

A chemical reaction can be represented by a chemical equation. Using the same example, the burning of methane gas can be represented by the equation:

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$.

The arrow in a chemical equation separates the reactants from the products and shows the direction in which the reaction occurs. If the reaction could also occur in the opposite direction, then two arrows, one pointing in each direction, would be used. On each side of the arrow, a mixture of chemicals is indicated by the chemical symbols joined by a plus sign (+). The numbers preceding some of the chemical symbols (such as $2\ O_2$) indicate how many molecules of the chemicals are involved in the reaction. (If there is no number in front of a chemical symbol, it means that just one molecule is involved.)

In a chemical reaction, the quantity of each element does not change. There is the same amount of each element at the end of the reaction as there was at the beginning. This is reflected in the chemical equation for the reaction. The equation should be balanced. In a balanced equation, the same number of atoms of a given element appear on each side of the arrow. For example, in the equation above, there are four hydrogen atoms on each side of the arrow.

Types of Chemical Reactions

In general, a chemical reaction involves the breaking and forming of chemical bonds. In the methane reaction above, bonds are broken in methane and oxygen, and bonds are formed in carbon dioxide and water. A reaction like this, in which a compound or element burns in oxygen, is called a **combustion reaction**. This is just one of many possible types of chemical reactions. Other types of chemical reactions include synthesis, decomposition, and substitution reactions.

- A **synthesis reaction** occurs when two or more chemical elements or compounds unite to form a more complex product. For example, nitrogen (N_2) and hydrogen (H_2) unite to form ammonia (NH_3):

$N_2 + 3\ H_2 \rightarrow 2\ NH_3$.

- A **decomposition reaction** occurs when a compound is broken down into smaller compounds or elements. For example, water (H_2O) breaks down into hydrogen (H_2) and oxygen (O_2):

$2\ H_2O \rightarrow 2\ H_2 + O_2$.

- A **substitution reaction** occurs when one element replaces another element in a compound. For example, sodium (Na^+) replaces hydrogen (H) in hydrochloric acid (HCl), producing sodium chloride (NaCl) and hydrogen gas (H_2):

$2\ Na^+ + 2\ HCl \rightarrow 2\ NaCl + H_2$.

Chemical Reactions and Energy

Some chemical reactions consume energy, whereas other chemical reactions release energy. Each of the energy changes that occur during a reaction are graphed in **Figure** 11.2. In the reaction on the left, energy is released. In the reaction on the right, energy is consumed.

Bill Nye discusses chemical reactions at http://www.youtube.com/watch?v=66kuhJkQCVM (2:05).

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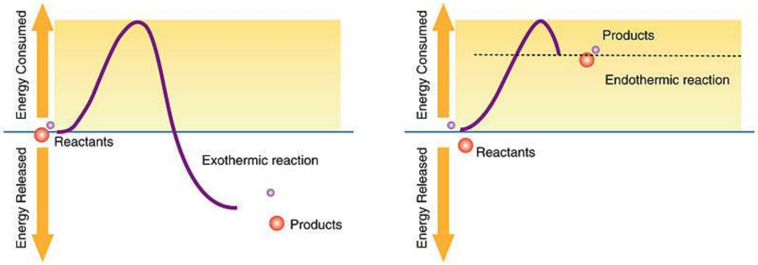


FIGURE 11.2

The reaction on the left releases energy. The reaction on the right consumes energy.

Exothermic Reactions

Chemical reactions that release energy are called **exothermic reactions**. An example is the combustion of methane described at the beginning of this lesson. In organisms, exothermic reactions are called **catabolic reactions**. Catabolic reactions break down molecules into smaller units. An example is the breakdown of glucose molecules for energy. Exothermic reactions can be represented by the general chemical equation:

Reactants → Products + Heat.

Endothermic Reactions

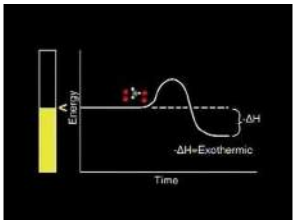
Chemical reactions that consume energy are called **endothermic reactions**. An example is the synthesis of ammonia, described above. In organisms, endothermic reactions are called **anabolic reactions**. Anabolic reactions construct molecules from smaller units. An example is the synthesis of proteins from amino acids. Endothermic reactions can be represented by the general chemical equation:

Reactants + Heat → Products.

Activation Energy

Regardless of whether reactions are exothermic or endothermic, they all need energy to get started. This energy is called **activation energy**. Activation energy is like the push you need to start moving down a slide. The push gives you enough energy to start moving. Once you start, you keep moving without being pushed again. The concept of activation energy is illustrated in **Figure** 11.3.

An overview of activation energy (**1b**) can be viewed at http://www.youtube.com/watch?v=VbIaK6PLrRM&feature=related (1:16).



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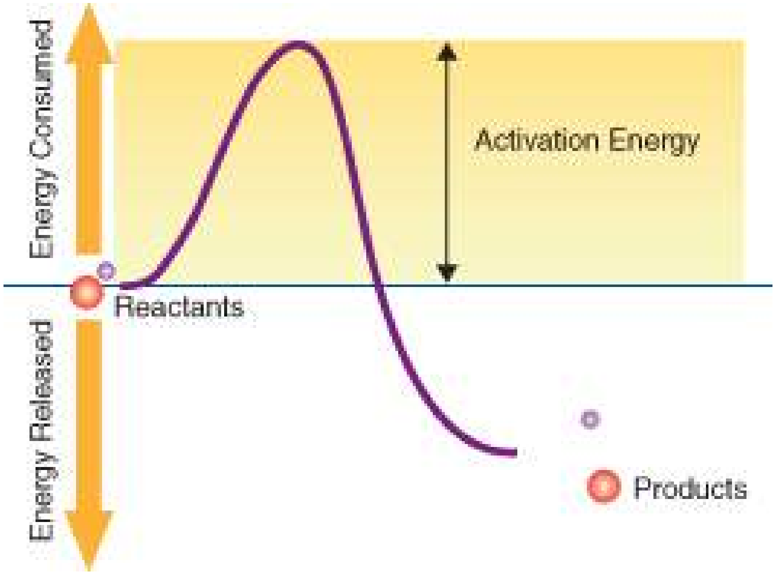

FIGURE 11.3

To start this reaction, a certain amount of energy is required, called the activation energy. How much activation energy is required depends on the nature of the reaction and the conditions under which the reaction takes place.

Why do reactions need energy to get started? In order for reactions to occur, three things must happen, and they all require energy:

- Reactant molecules must collide. To collide, they must move, so they need kinetic energy.

- Unless reactant molecules are positioned correctly, intermolecular forces may push them apart. To overcome these forces and move together requires more energy.
- If reactant molecules collide and move together, there must be enough energy left for them to react.

Rates of Chemical Reactions

The rates at which chemical reactions take place in organisms are very important. Chemical reactions in organisms are involved in processes ranging from the contraction of muscles to the digestion of food. For example, when you wave goodbye, it requires repeated contractions of muscles in your arm over a period of a couple of seconds. A huge number of reactions must take place in that time, so each reaction cannot take longer than a few milliseconds. If the reactions took much longer, you might not finish waving until sometime next year.

Factors that help reactant molecules collide and react speed up chemical reactions. These factors include the concentration of reactants and the temperature at which the reactions occur.

- Reactions are usually faster at higher concentrations of reactants. The more reactant molecules there are in a given space, the more likely they are to collide and react.
- Reactions are usually faster at higher temperatures. Reactant molecules at higher temperatures have more energy to move, collide, and react.

Enzymes and Biochemical Reactions

Most chemical reactions within organisms would be impossible under the conditions in cells. For example, the body temperature of most organisms is too low for reactions to occur quickly enough to carry out life processes. Reactants may also be present in such low concentrations that it is unlikely they will meet and collide. Therefore, the rate of most biochemical reactions must be increased by a catalyst. A catalyst is a chemical that speeds up chemical reactions. In organisms, catalysts are called **enzymes**.

Like other catalysts, enzymes are not reactants in the reactions they control. They help the reactants interact but are not used up in the reactions. Instead, they may be used over and over again. Unlike other catalysts, enzymes are usually highly specific for particular chemical reactions. They generally catalyze only one or a few types of reactions.

Enzymes are extremely efficient in speeding up reactions. They can catalyze up to several million reactions per second. As a result, the difference in rates of biochemical reactions with and without enzymes may be enormous. A typical biochemical reaction might take hours or even days to occur under normal cellular conditions without an enzyme but less than a second with the enzyme.

An overview of enzymes **(1b)** can be viewed at http://www.youtube.com/watch?v=E90D4BmaVJM&feature=related (9:43).

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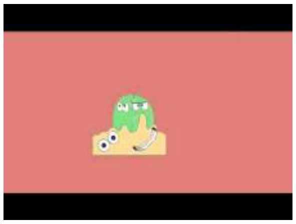
How Enzymes Work

How do enzymes speed up biochemical reactions so dramatically? Like all catalysts, enzymes work by lowering the activation energy of chemical reactions. This is illustrated in **Figure** 11.4. The biochemical reaction shown in the figure requires about three times as much activation energy without the enzyme as it does with the enzyme.

An animation of how enzymes work **(1b)** can be seen at http://www.youtube.com/watch?v=CZD5xsOKres&feature=related (2:02).



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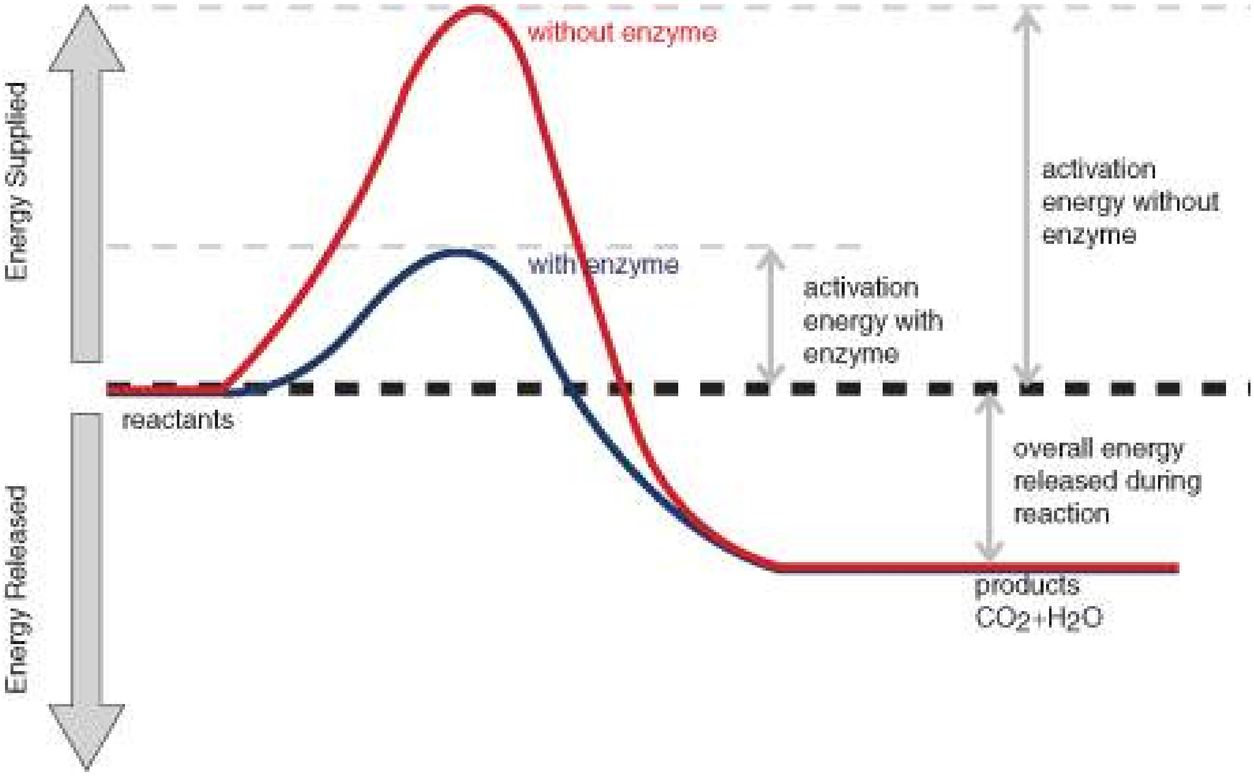

FIGURE 11.4

The reaction represented by this graph is a combustion reaction involving the reactants glucose ($C_6H_{12}O_6$) and oxygen (O_2). The products of the reaction are carbon dioxide (CO_2) and water (H_2O). Energy is also released during the reaction. The enzyme speeds up the reaction by lowering the activation energy needed for the reaction to start. Compare the activation energy with and without the enzyme.

Enzymes generally lower activation energy by reducing the energy needed for reactants to come together and react. For example:

- Enzymes bring reactants together so they don't have to expend energy moving about until they collide at random. Enzymes bind both reactant molecules (called substrate), tightly and specifically, at a site on the enzyme molecule called the active site (**Figure** 11.5).

- By binding reactants at the active site, enzymes also position reactants correctly, so they do not have to overcome intermolecular forces that would otherwise push them apart. This allows the molecules to interact with less energy.

- Enzymes may also allow reactions to occur by different pathways that have lower activation energy.

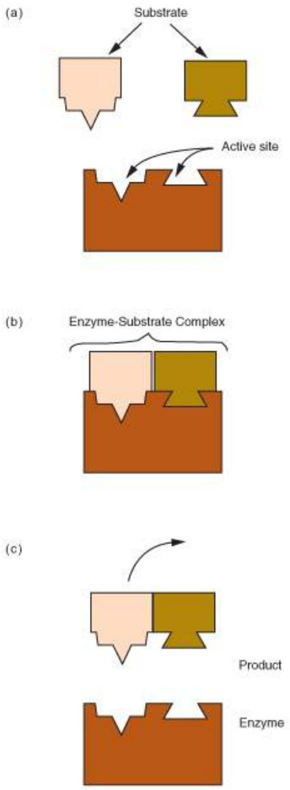

FIGURE 11.5

This enzyme molecule binds reactant molecules—called substrate—at its active site, forming an enzyme-substrate complex. This brings the reactants together and positions them correctly so the reaction can occur. After the reaction, the products are released from the enzyme's active site. This frees up the enzyme so it can catalyze additional reactions.

The activities of enzymes also depend on the temperature, ionic conditions, and the pH of the surroundings.

Some enzymes work best at acidic pHs, while others work best in neutral environments.

- Digestive enzymes secreted in the acidic environment (low pH) of the stomach help break down proteins into smaller molecules. The main digestive enzyme in the stomach is pepsin, which works best at a pH of about 1.5 (see the *Digestive and Excretory Systems* chapter). These enzymes would not work optimally at other pHs. Trypsin is another enzyme in the digestive system which break protein chains in the food into smaller parts. Trypsin works in the small intestine, which is not an acidic environment. Trypsin's optimum pH is about 8.

- Biochemical reactions are optimal at physiological temperatures. For example, most biochemical reactions work best at the normal body temperature of 98.6˚F. Many enzymes lose function at lower and higher temperatures. At higher temperatures, an enzyme's shape deteriorates and only when the temperature comes back to normal does the enzyme regain its shape and normal activity.

Importance of Enzymes

Enzymes are involved in most of the chemical reactions that take place in organisms. About 4,000 such reactions are known to be catalyzed by enzymes, but the number may be even higher.

In animals, an important function of enzymes is to help digest food. Digestive enzymes speed up reactions that break down large molecules of carbohydrates, proteins, and fats into smaller molecules the body can use (*See* Chapter: Digestive and Excretory Systems). Without digestive enzymes, animals would not be able to break down food molecules quickly enough to provide the energy and nutrients they need to survive.

Lesson Summary

- A chemical reaction is a process that changes some chemical substances into others. It involves breaking and forming chemical bonds. Types of chemical reactions include synthesis reactions and decomposition reactions.
- Some chemical reactions are exothermic, which means they release energy. Other chemical reactions are endothermic, which means they consume energy. All chemical reactions require activation energy, which is the energy needed to get a reaction started.
- Rates of chemical reactions depend on factors such as the concentration of reactants and the temperature at which reactions occur. Both factors affect the ability of reactant molecules to react.
- Enzymes are needed to speed up chemical reactions in organisms. They work by lowering the activation energy of reactions.

Review Questions

1. Identify the roles of reactants and products in a chemical reaction.
2. What is the general chemical equation for an endothermic reaction?
3. State two factors, other than enzymes, that speed up chemical reactions.
4. How do enzymes work to speed up chemical reactions?
5. What is wrong with the chemical equation below? How could you fix it? $CH_4 + O_2 \rightarrow CO_2 + 2\ H_2O$
6. What type of reaction is represented by the following chemical equation? Explain your answer. $2\ Na + 2\ HCL \rightarrow 2\ NaCl + H_2$
7. Why do all chemical reactions require activation energy?
8. Explain why organisms need enzymes to survive.

Further Reading / Supplemental Links

- Peter Atkins and Julio De Paula, *Physical Chemistry for the Life Sciences*. Oxford University Press, 2006.
- Rita Elkins, *Digestive Enzymes*. Woodland Publishing, 2007.
- James Keeler and Peter Wothers, *Why Chemical Reactions Happen*. Oxford University Press, 2003.
- George W. Roberts, *Chemical Reactions and Chemical Reactors*. Wiley, 2008.
- http://en.wikipedia.org

Summary Animations

- http://www.stolaf.edu/people/giannini/flashanimat

Vocabulary

activation energy	Energy needed for a chemical reaction to get started.
anabolic reaction	Endothermic reaction that occurs in organisms.
catabolic reaction	Exothermic reaction that occurs in organisms.
chemical reaction	Process that changes some chemical substances into other chemical substances.
combustion reaction	Type of chemical reaction in which a compound or element burns in oxygen.
decomposition reaction	Type of chemical reaction in which a compound is broken down into smaller compounds or elements.
endothermic reaction	Any chemical reaction that consumes energy.
enzyme	Chemical that speeds up chemical reactions in organisms.
exothermic reaction	Any chemical reaction that releases energy.
product	Substance that forms as a result of a chemical reaction.
reactant	Substance involved in a chemical reaction that is present at the beginning of the reaction.
substitution reaction	Type of chemical reaction in which one element replaces another element in a compound.
synthesis reaction	Type of chemical reaction in which elements or compounds unite to form a more complex product.

Points to Consider

Most chemical reactions in organisms take place in an environment that is mostly water.

- What do you know about water?
- Are you aware that water has unique properties?
- Do you know how water behaves differently from most other substances on Earth?
- Do you know why water is necessary for life?

11.3 References

1. CK-12 Foundation. . CC-BY-NC-SA 3.0
2. CK-12 Foundation. . CC-BY-SA
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CHAPTER 12 Chemical Equllibrium

Chapter Outline

Chm.3.1 Understand the factors affecting rate of reaction and chemical equilibrium.

Chm.3.1.1 Explain the factors that affect the rate of a reaction (temperature, concentration, particle size and presence of a catalyst). Chm.3.1.2 Explain the conditions of a system at equilibrium. Chm.3.1.3 Infer the shift in equilibrium when a stress is applied to a chemical system (Le Chatelier's Principle).

UNPACKED

Chm.3.1.1

- Understand qualitatively that reaction rate is proportional to number of effective collisions.
- Explain that nature of reactants can refer to their complexity and the number of bonds that must be broken and reformed in the course of reaction.
- Explain how temperature (kinetic energy), concentration, and/or pressure affects the number of collisions.
- Explain how increased surface area increases number of collisions.
- Explain how a catalyst lowers the activation energy, so that at a given temperature, more molecules will have energy equal to or greater than the activation energy.
  Chm.3.1.2

- Define chemical equilibrium for reversible reactions.
- Distinguish between equal rates and equal concentrations.
- Explain equilibrium expressions for a given reaction.
- Evaluate equilibrium constants as a measure of the extent that the reaction proceeds to completion.
- Determine the effects of stresses on systems at equilibrium. (Adding/ removing a reactant or product; adding/removing heat; increasing/decreasing pressure)
- Relate the shift that occurs in terms of the order/disorder of the system.

12.1 Chemical Equilibrium

- Define chemical equilibrium.
- List conditions for equilibrium.

Pull hard

A tug of war involves two teams at the ends of a rope. The goal is to pull the other team over a line in the middle. At first, there is a great deal of tension on the rope, but no apparent movement. A bystander might think that there is nothing happening. In reality, there is a great deal of tension on the rope as the two teams pull in opposite directions at the same time.

Chemical Equilibrium

Hydrogen and iodine gases react to form hydrogen iodide according to the following reaction:

$$\begin{aligned} & H_2(g)+I_2(g) \rightleftarrows 2HI(g) \\ \text{Forward reaction:}\quad & H_2(g)+I_2(g) \rightarrow 2HI(g) \\ \text{Reverse reaction:}\quad & 2HI(g) \rightarrow H_2(g)+I_2(g) \end{aligned}$$

Initially, only the forward reaction occurs because no HI is present. As soon as some HI has formed, it begins to decompose back into H_2 and I_2. Gradually, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually the rate of combination of H_2 and I_2 to produce HI becomes equal to the rate of decomposition of HI into H_2 and I_2. When the rates of the forward and reverse reactions have become equal to one another, the reaction has achieved a state of balance. **Chemical equilibrium** is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.

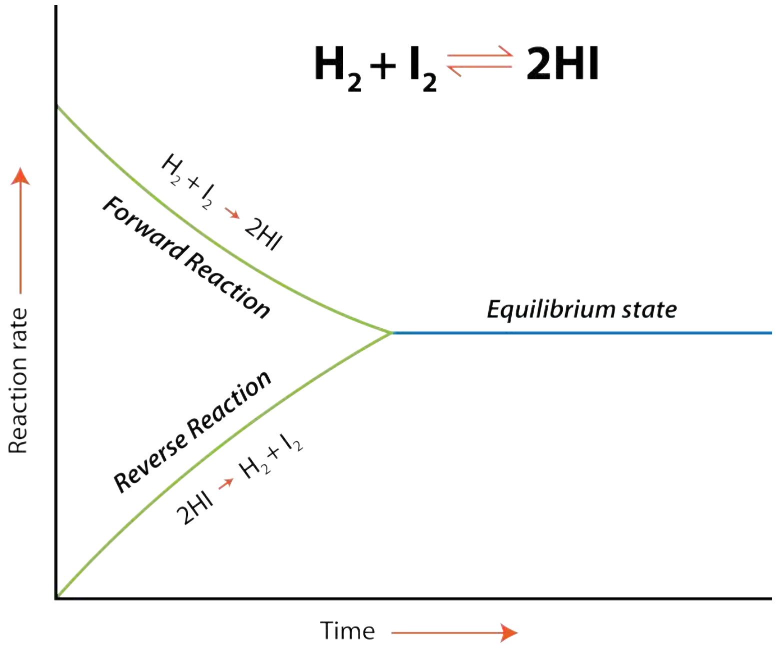


FIGURE 12.1

Equilibrium in reaction: $H_2(g) + I_2(g) \rightarrow 2HI(g)$.

Chemical equilibrium can be attained whether the reaction begins with all reactants and no products, all products and no reactants, or some of both. The **Figure** 12.2 shows changes in concentration of H_2, I_2, and HI for two different reactions. In the reaction depicted by the graph on the left (A), the reaction begins with only H_2 and I_2 present. There is no HI initially. As the reaction proceeds towards equilibrium, the concentrations of the H_2 and I_2 gradually decrease, while the concentration of the HI gradually increases. When the curve levels out and the concentrations all become constant, equilibrium has been reached. At equilibrium, concentrations of all substances are constant. In reaction B, the process begins with only HI and no H_2 or I_2. In this case, the concentration of HI gradually decreases while the concentrations of H_2 and I_2 gradually increase until equilibrium is again reached. Notice that in both cases, the relative position of equilibrium is the same, as shown by the relative concentrations of reactants and products. The concentration of HI at equilibrium is significantly higher than the concentrations of H_2 and I_2. This is true whether the reaction began with all reactants or all products. The position of equilibrium is a property of the particular reversible reaction and does not depend upon how equilibrium was achieved.

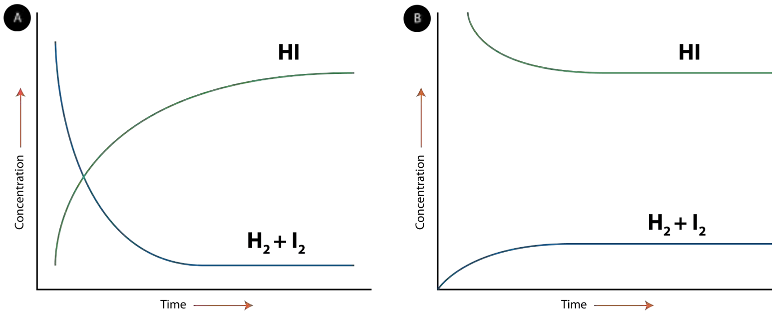


FIGURE 12.2

Equilibrium between reactants and products.

Conditions for Equilibrium and Types of Equilibrium

It may be tempting to think that once equilibrium has been reached, the reaction stops. Chemical equilibrium is a dynamic process. The forward and reverse reactions continue to occur even after equilibrium has been reached. However, because the rates of the reactions are the same, there is no change in the relative concentrations of reactants and products for a reaction that is at equilibrium. The conditions and properties of a system at equilibrium are summarized below.

1. The system must be closed, meaning no substances can enter or leave the system.
2. Equilibrium is a dynamic process. Even though we don't necessarily see the reactions, both forward and reverse are taking place.
3. The rates of the forward and reverse reactions must be equal.
4. The amount of reactants and products do not have to be equal. However, after equilibrium is attained, the amounts of reactants and products will be constant.

Summary

- The concept of chemical equilibrium is described.
- Conditions for chemical equilibrium are listed.

Practice

Read the material at the link below and answer the following questions:

http://www.chemguide.co.uk/physical/equilibria/introduction.html#top

1. What is a closed system?
2. What is a dynamic system?
3. What happens if you change the relative rates of the forward and back reactions?

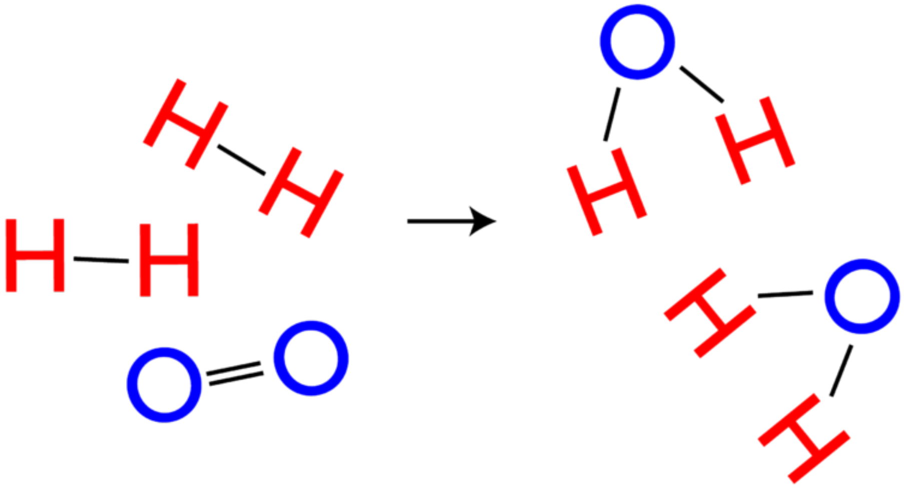
Review

1. What is chemical equilibrium?
2. In the reaction illustrated in the text, does the equilibrium concentration of HI equal the equilibrium concentrations of H_2 and I_2?
3. Does the position at equilibrium depend upon how the equilibrium was reached?

- **chemical equilibrium:** The state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.

12.2 Reversible Reactions

- Give examples of irreversible and reversible chemical reactions.
- Describe a reversible reaction that is in equilibrium.



This diagram represents a chemical reaction. In the reaction, two molecules of hydrogen gas (2 H-H) combine one molecule of oxygen gas (O=O) to produce two molecules of water (2 H-O-H). Look at the arrow in the diagram. What does it represent? It shows that the reaction normally occurs in that direction, that is, from the two gases to water. Does the opposite reaction ever occur? Can water change back to hydrogen and oxygen gases? In other words, is the reaction reversible? Read on to find out.

Irreversible Reactions

Some chemical reactions can occur in only one direction. These reactions are called irreversible reactions. The reactants can change to the products, but the products cannot change back to the reactants. These reactions are like making a cake. The ingredients of a cake—such as eggs and flour—are the reactants. They are mixed together and baked to form the cake, which is the product (see **Figure** 12.3). The cake can't be "unbaked" and "unmixed" to change it back to the raw eggs, flour, and other ingredients. So making a cake is irreversible.

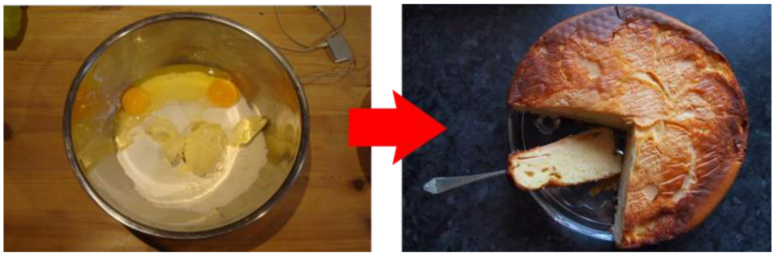


FIGURE 12.3

Combustion reactions are generally irreversible. Combustion occurs whenever a fuel burns. In this type of reaction, the fuel may combine with oxygen (in the air) and produces carbon dioxide and water vapor. The chemical equation for a combustion reaction is:

Fuel + $O_2 \rightarrow CO_2 + H_2O$

In a complete combustion reaction, fuel and oxygen are the reactants and the products are carbon dioxide and water. These two products cannot react to reform the fuel and oxygen, so the reaction is irreversible. The one-way arrow in the equation shows that the reaction can go in only one direction.

Reversible Reactions

Many chemical reactions can occur in both directions. These reactions are called reversible reactions. Not only can the reactants change to the products, but the products can change back to the reactants, at least under certain conditions. Consider again the reaction in which hydrogen gas and oxygen gas combine to produce water. This reaction is reversible if an electric current is applied to the water that is produced. You can see how this is done in the **Figure** 12.4. The current causes the water molecules to break down to individual hydrogen and oxygen atoms. Then these atoms recombine to form molecules of hydrogen gas and oxygen gas.

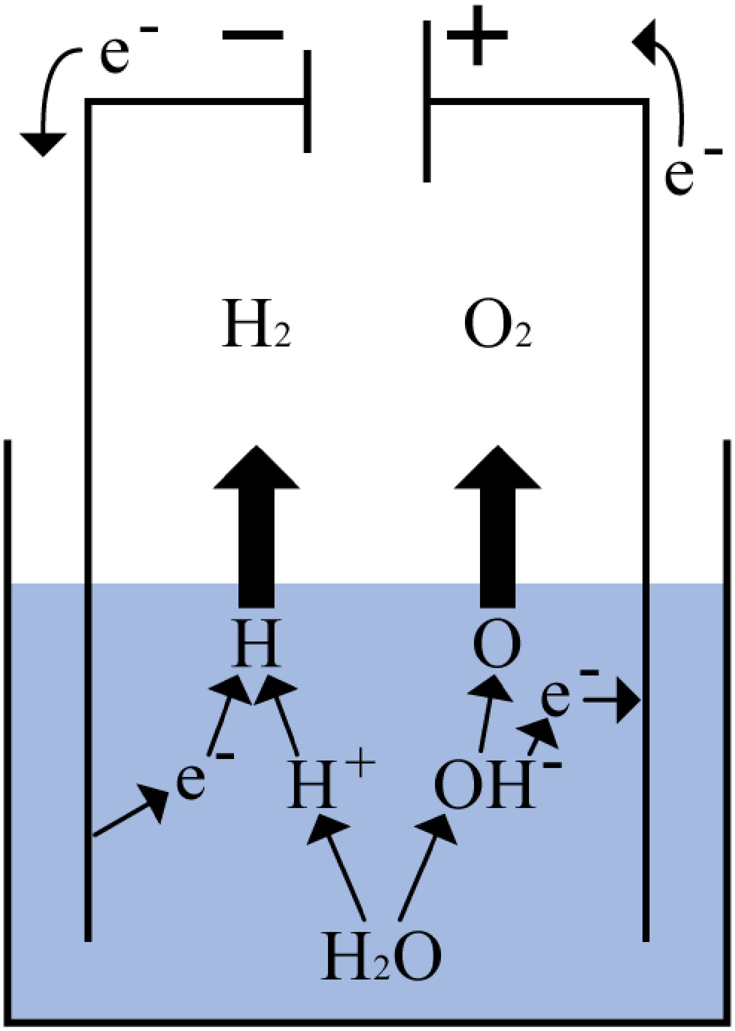


FIGURE 12.4

Q: In a chemical equation, a reversible reaction is represented with two arrows, one pointing in each direction. This shows that the reaction can go both ways. How would you represent the reversible reaction in which hydrogen and oxygen gases combine to produce water?

A: The chemical equation would be:

$H_2 + O_2 \rightleftharpoons H_2O$

Balancing Act

In a reversible reaction, both forward and reverse directions of the reaction generally occur at the same time. While reactants are reacting to produce products, products are reacting to produce reactants. Often, a point is reached at which forward and reverse directions of the reaction occur at the same rate. When this happens, there is no overall change in the amount of reactants and products, even though the reactions keep occurring in both directions. This point is called **equilibrium**. The term *equilibrium* means "state of balance," and it is used to refer to a state of balance between any opposing changes. You can see an animation of a chemical reaction reaching equilibrium at this URL: http://www.tutorvista.com/content/chemistry/chemistry-ii/chemical-equilibrium/chemical-equilibrium-animation.php .

Summary

- Irreversible chemical reactions can occur in only one direction. The reactants can change to the products, but the products cannot change back to the reactants.
- Reversible chemical reactions can occur in both directions. The reactants can change to the products, and the products can also change back to the reactants.
- Equilibrium occurs when forward and reverse directions of a reversible reaction occur at the same rate so there is no overall change in the amounts of reactants and products.

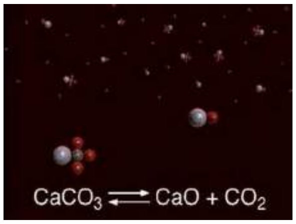
Vocabulary

- **equilibrium**: Balance between opposing changes, such as the forward and reverse directions of a chemical reaction.

Practice

Watch the video about chemical equilibrium at the following URL, and then answer the questions below.

http://www.youtube.com/watch?v=LMIbJ-B92Ho



MEDIA

Click image to the left for more content.

1. What is the reversible reaction presented in the video? What type of reaction is it?
2. When does the reverse reaction start occurring? What type of reaction is the reverse reaction?
3. What shifts the point of equilibrium in this chemical reaction?

Review

1. What is an irreversible chemical reaction? Give an example.

2. Describe a reversible chemical reaction.
3. When is a reversible chemical reaction in equilibrium?

12.3 Catalysts

- Define catalyst.
- Explain how catalysts influence the rate of a chemical reaction.

Where did the sun go?

Gasoline-powered vehicles emit a lot of harmful materials. Nitrogen oxides are formed when atmospheric nitrogen reacts with oxygen at the high temperatures found in a car engine. Carbon monoxide is a by-product of incomplete combustion of hydrocarbons. Evaporated and unused fuel releases volatile hydrocarbons into the atmosphere to help form smog. The presence of a catalytic converter in the exhaust system of the car causes these materials to react and be changed into less harmful products.

Catalysts

Sometimes a substance can be added to a chemical reaction and it will cause that reaction to undergo a dramatic increase in rate. Hydrogen peroxide is used as a disinfectant for scrapes and cuts and is found in many medicine cabinets as a 3% aqueous solution. Hydrogen peroxide naturally decomposes to produce water and oxygen gas, but the reaction is very slow. A bottle of hydrogen peroxide will last for several years before it needs to be replaced. However, the addition of just a small amount of manganese(IV) oxide to hydrogen peroxide will cause it to decompose completely in just a matter of minutes. A **catalyst** is a substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction. After the reaction occurs, a catalyst returns to its original state and so catalysts can be used over and over again. Because it is neither a reactant nor a product, a catalyst is shown in a chemical equation by being written above the yield arrow.

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

A catalyst works by changing the specific way in which the reaction occurs, called its mechanism. The important outcome from the use of a catalyst is that the overall activation energy of the reaction is lowered (see **Figure** 12.5). With a lower activation energy barrier, a greater percentage of reactant molecules are able to have effective collisions and the reaction rate increases.

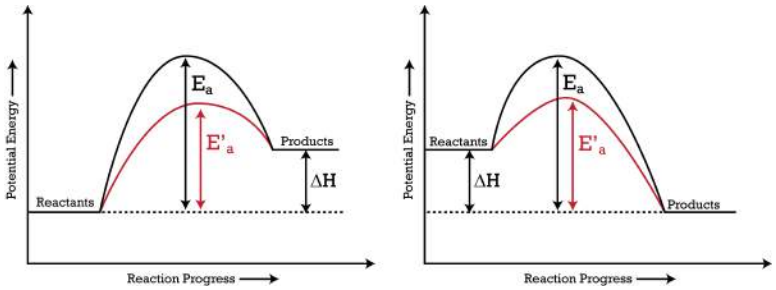


FIGURE 12.5

The addition of a catalyst to a reaction lowers the activation energy, increasing the rate of the reaction. The activation energy of the uncatalyzed reaction is shown by Ea, while the catalyzed reaction is shown by Ea'. The heat of reaction (ΔH) is unchanged by the presence of the catalyst.

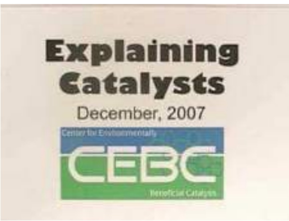
Catalysts are extremely important parts of many chemical reactions. Enzymes in your body act as nature's catalysts, allowing important biochemical reactions to occur at reasonable rates. Chemical companies constantly search for new and better catalysts to make reactions go faster and thus make the company more profitable.

Summary

- The function of a catalyst is described.

Practice

Watch the video at the link below and answer the following questions:



MEDIA

Click image to the left for more content.

http://www.youtube.com/watch?v=9KQdF1bnXHE

1. What dye is used on blue jeans?
2. What do enzymes do for the blue jeans?
3. List three issues that exist for uncatalyzed reactions.

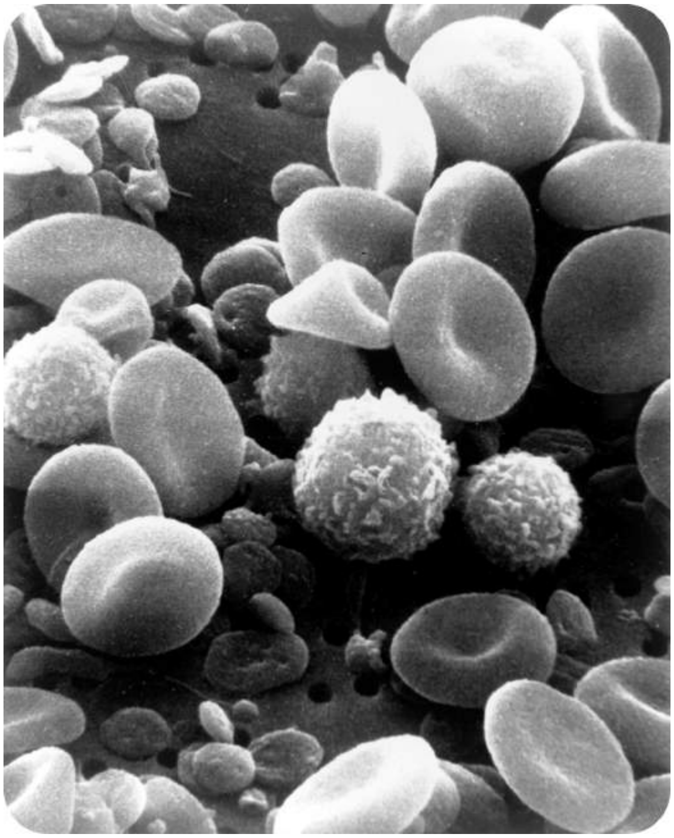
Review

1. What does a catalyst affect in a chemical reaction?
2. Is the ΔH for the process affected?
3. When writing a chemical equation, where do we indicate the catalyst?

- **catalyst:** A substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction.

12.4 Equilibrium Constant

- Define equilibrium constant.
- Write the equation for the general equilibrium constant.



Carbon monoxide poisoning

Red blood cells transport oxygen to the tissues so they can function. In the absence of oxygen, cells cannot carry out their biochemical responsibilities. Oxygen moves to the cells attached to hemoglobin, a protein found in the red cells. In cases of carbon monoxide poisoning, CO binds much more strongly to the hemoglobin, blocking oxygen attachment and lowering the amount of oxygen reaching the cells. Treatment involves the patient breathing pure oxygen to displace the carbon monoxide. The equilibrium reaction shown below illustrates the shift toward the right when excess oxygen is added to the system:

$$\text{Hb(CO)}_4(aq) + 4\text{O}_2(g) \leftrightharpoons \text{Hb(O}_2)_4(aq) + 4\text{CO}(g)$$

Equilibrium Constant

The general value of the equilibrium constant gives us information about whether the reactants or the products are favored at equilibrium. Since the product concentrations are in the numerator of the equilibrium expression, a $K_{eq} > 1$ means that the products are favored over the reactants. A $K_{eq} < 1$ means that the reactants are favored over the products.

Consider the hypothetical reversible reaction in which reactants A and B react to form products C and D. This equilibrium can be shown below, where the lower case letters represent the coefficients of each substance.

$$aA + bB \rightleftarrows cC + dD$$

As we have established, the rates of the forward and reverse reactions are the same at equilibrium, and so the concentrations of all of the substances are constant. Since that is the case, it stands to reason that a ratio of the concentrations for any given reaction at equilibrium maintains a constant value. The **equilibrium constant** (K_{eq}) is the ratio of the mathematical product of the concentrations of the products of a reaction to the mathematical product of the concentrations of the reactants of the reaction. Each concentration is raised to the power of its coefficient in the balanced chemical equation. For the general reaction above, the equilibrium constant expression is written as follows:

$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

The concentrations of each substance, indicated by the square brackets around the formula, are measured in molarity units (mol/L).

The value of the equilibrium constant for any reaction is only determined by experiment. As detailed in the above section, the position of equilibrium for a given reaction does not depend on the starting concentrations and so the value of the equilibrium constant is truly constant. It does, however, depend on the temperature of the reaction. This is because equilibrium is defined as a condition resulting from the rates of forward and reverse reactions being equal. If the temperature changes, the corresponding change in those reaction rates will alter the equilibrium constant. For any reaction in which a K_{eq} is given, the temperature should be specified.

Summary

- The equilibrium constant for a reversible reaction is described.

Practice

Read the material at the link below and answer the following questions:

1. What is a homogeneous equilibrium?
2. What is a heterogeneous equilibrium?
3. What goes in the numerator in an equilibrium expression?
4. Why is it important to write your equilibrium equation out before setting up the equilibrium constant?

Review

1. What does the equilibrium constant tell us?
2. What does it mean if the K_{eq} is >1?
3. What does it mean if the K_{eq} is <1?
4. Does the position of equilibrium depend on the starting concentrations?

- **equilibrium constant** (K_{eq})**:** The ratio of the mathematical product of the concentrations of the products of a reaction to the mathematical product of the concentrations of the reactants of the reaction. Each concentration is raised to the power of its coefficient in the balanced chemical equation.

12.5 References

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2. CK-12 Foundation - Christopher Auyeung. . CC-BY-NC-SA 3.0
3. CK-12 Foundation - Christopher Auyeung. . CC-BY-NC-SA 3.0
4. Ingredients: Gemma Bardsley; Cake: The Integer Club. Baking a cake is an irreversible reaction. CC BY 2.0
5. Zachary Wilson. Decomposition of water to form hydrogen and oxygen gas. CC BY-NC 3.0
6. Courtesy of Dr. Edwin P. Ewing, Jr., CDC. http://commons.wikimedia.org/wiki/File:SmogNY.jpg . Public Domain
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CHAPTER 13 Pressure, Temperature, Volume and Phase

Chapter Outline

Chm.2.1.1

- Explain physical equilibrium: liquid water-water vapor. Vapor pressure depends on temperature and concentration of particles in
  solution. (conceptual only –no calculations)
- Explain how the energy (kinetic and potential) of the particles of a substance changes when heated, cooled, or changing phase.
- Identify pressure as well as temperature as a determining factor for phase of matter.
- Contrast heat and temperature, including temperature as a measure of average kinetic energy, and appropriately use the units Joule,
  Celsius, and Kelvin

Chm.2.1.2

- Define and use the terms and/or symbols for: specific heat capacity, heat of fusion, heat of vaporization.
- Interpret the following:
  - heating and cooling curves (noting both significance of plateaus and the physical states of each segment
  - Phase diagrams for H2O and CO2,
- Complete calculations of: $q=mCp\Delta T$, $q = mHf$, and $q = mHv$ using heatling/cooling curve data.
- Explain phase change calculations in terms of heat absorbed or released (endothermic vs. exothermic processes).

Chm.2.1.3

- Draw phase diagrams of water and carbon dioxide (shows how sublimation occurs). Identify regions, phases and phase changes using a
  phase diagram.
- Use phase diagrams to determine information such as (1) phase at a given temperature and pressure, (2) boiling point or melting point at
  a given pressure, (3) triple point of a material.

Chm.2.1.4

- Recognize that, for a closed system, energy is neither lost nor gained only transferred between components of the system.
- Complete calculations of: $q=mCp\Delta T$, q = mHf , q = mHv, and q lost=(-q gain) in water, including phase changes, using laboratory data.

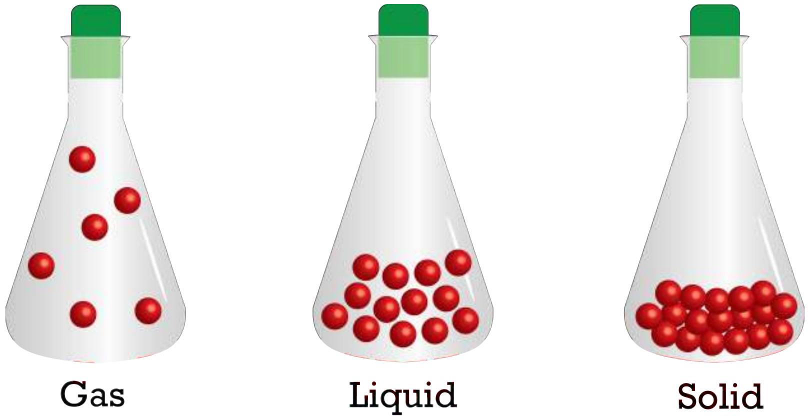
13.1 Change of State

- Define heat of fusion.
- Define heat of vaporization.
- Calculate heat transfers necessary for changes of state.

Before the internal combustion engine was invented, steam engines were the power source for ships, locomotives, tractors, lumber saws, and most industrial machines. Coal or wood was burned to boil water into steam, which ran the engine.

Change of State

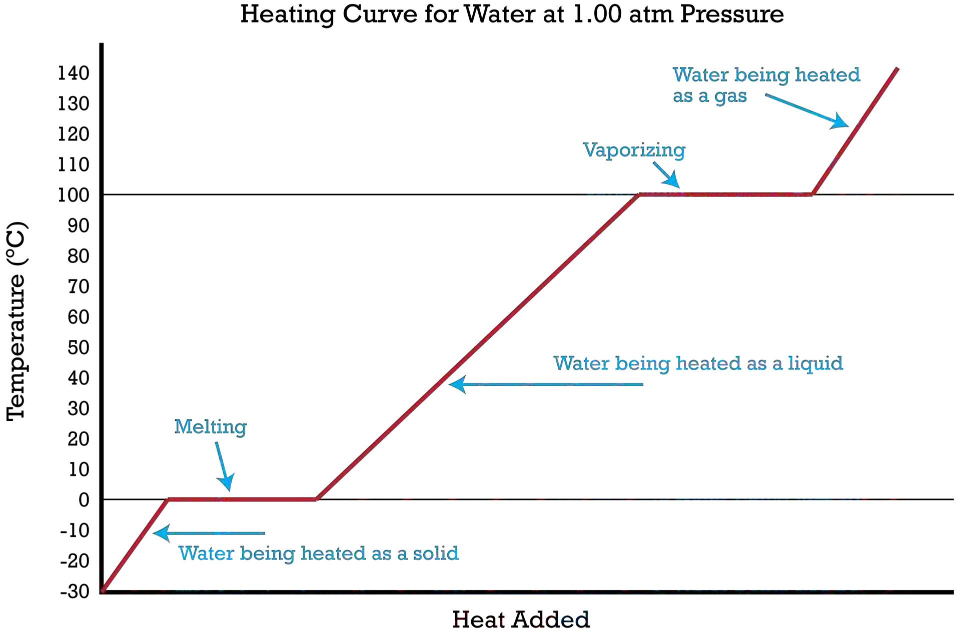
Most substances may exist in any of the three common states of matter. In the gaseous state, the molecular motion has completely overcome any attraction between the particles and the particles are totally separate from each other. There are large spaces between the particles and they move large distances between collisions. In the liquid state, the molecular motion and the molecular attractions are more balanced. While the particles stay more or less in contact with each other, they are still free to move and can slide past one another easily. In the solid state, the attractive forces dominate. The particles are pulled together into a tightly packed pattern which does not allow the particles to pass each other. The molecular motion in this form is essentially reduced to vibration in place. Increasing the temperature of a substance means increasing the molecular motion (kinetic energy) of the molecules in the substance. The phase in which a substance exists is the result of a competition between attractive forces and molecular motion.



For most substances, when the temperature of the solid is raised high enough, the substance changes to a liquid, and when the temperature of the liquid is raised high enough, the substance changes to a gas. We typically visualize a solid as tiny particles in constant motion held together by attractive forces. As we add heat to the solid, the motion, or the kinetic energy, of the particles increases. At some temperature, the motion of the particles becomes great enough to overcome the attractive forces. The thermal energy that was added to the solid up to this point was absorbed by the solid as kinetic energy, increasing the speed of the molecules. The lowest temperature at which the particles are able to exist in the liquid form is called the **melting point**.

In order for the molecules to actually separate from each other, more energy must be added. This energy, called **heat of fusion** or heat of melting, is absorbed by the particles as potential energy as the solid changes to a liquid. Recognize that, once the temperature of a solid has been raised to the melting point, it is still necessary for the solid to absorb additional thermal energy in the form of potential energy as the molecules separate.

The **boiling point** of a liquid is the temperature at which the particles have sufficient molecular motion to exist in the form of a gas. Once again, however, in order for the particles to separate to the gaseous form, they must absorb a sufficient amount of potential energy. The amount of potential energy necessary for a **phase change** to gaseous form is called the **heat of vaporization**. Consider the heating curve shown below.



The heating curve shown is for water but other substances have similarly shaped heating curves. Suppose you begin with solid water (ice) at -30°C and add heat at a constant rate. The heat you add in the beginning will be absorbed as kinetic energy and the temperature of the solid will increase. When you reach a temperature of 0°C (the melting point for water), the heat you add is no longer absorbed as kinetic energy. Instead, the added heat is absorbed as potential energy and the particles separate from each other. During the flat part of the curve labeled "melting", heat is being added constantly but the temperature does not increase. At the left edge of this flat line, the water is solid; by the time enough heat has been added to get to the right edge, the water is liquid, but maintains the same temperature. Once all the water is in the liquid form, the added heat will once again be absorbed as kinetic energy and the temperature will increase again. During the time labeled "water being heated as a liquid", all the added heat is absorbed as kinetic energy.

When a temperature of 100°C (the boiling point of water) is reached, the added heat is once again absorbed as potential energy and the molecules separate from liquid form into gaseous form. When all the substance has been converted into gas, the temperature will again begin to rise.

TABLE 13.1: Heats of Fusion and Vaporization of Some Common Substances

Substance	Heat of Fusion, H_f (J/kg)	Heat of Vaporization, H_v (J/kg)
Copper	2.05×10^5	5.07×10^6
Gold	6.30×10^4	1.64×10^6
Iron	2.66×10^5	6.29×10^6
Methanol	1.09×10^5	8.78×10^5
Water	3.34×10^5	2.26×10^6

When the temperature of a substance is changing, we can use the specific heat to determine the amount of heat that is being gained or lost. When a substance is changing phase, we can use the heat of fusion or heat of vaporization to determine the amount of heat being gained or lost. When a substance freezes from liquid to solid, the amount of heat given off is exactly the same as the amount of heat absorbed when the substance melts from solid to liquid. The equations for heat gained or lost are given here:

The heat gained or lost during a temperature change: $Q = mc\Delta t$.

The heat gained or lost during a phase change of solid to liquid: $Q = mH_f$.

The heat gained or lost during a phase change of liquid to gas: $Q = mH_v$.

Example Problem: 5000. Joules of heat is added to ice at 273 K. All the heat goes into changing solid ice into liquid water. How much ice is melted?

Solution: $m = \frac{Q}{H_f} = \frac{5000\ J}{3.34 \times 10^5\ J/kg} = 0.0150\ kg$

Example Problem: Beginning with 1.00 kg of ice at -20.0°C, heat is added until the substance becomes water vapor at 130.0°C. How much heat was added? The specific heat of ice is 2108 $J/kg°C$, the specific heat of liquid water is 4187 $J/kg°C$, and the specific heat of water vapor is 1996 $J/kg°C$.

Solution: 5 steps.

1. Calculate the heat required to raise the sample from -20.0°C to 0°C.
2. Calculate the heat required to melt the sample.
3. Calculate the heat required to raise the sample from 0°C to 100°C.
4. Calculate the heat required to vaporize the sample.
5. Calculate the heat required to raise the sample from 100°C to 130°C.

The solution is the sum of these steps.

1.$Q_{HS} = mc_{\text{ice}}\Delta t = (1.00\ kg)(2108\ J/kg.°C)(20.0°C) = 42160\ J$

2.$Q_{\text{Melt}} = mH_f = (1.00\ kg)(334000\ J/kg) = 334000\ J$

3.$Q_{HL} = mc_{\text{water}}\Delta t = (1.00\ kg)(4187\ J/kg.°C)(100.0°C) = 418700\ J$

4.$Q_{\text{Vap}} = mH_v = (1.00\ kg)(2260000\ J/kg) = 2260000\ J$

5.$Q_{\text{HV}} = mc_{\text{vapor}}\Delta t = (1.00\ kg)(1996\ J/kg.^\circ C)(30.0^\circ C) = 59880\ J$

Total Heat $= 3.11 \times 10^6\ J$

Summary

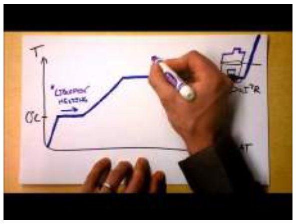
- Most substances may exist in any of the three common states of matter, solid, liquid, or gas.
- The phase in which a substance exists is the result of a competition between attractive forces and molecular motion.
- The potential energy absorbed by a solid as it changes to a liquid is called the heat of fusion or the heat of melting.
- The amount of potential energy necessary for a phase change to gaseous form is called the heat of vaporization.
- The heat gained or lost during a temperature change is given by, $Q = mc\Delta t$.
- The heat gained or lost during a phase change of solid to liquid is given by, $Q = mH_f$.
- The heat gained or lost during a phase change of liquid to gas is given by, $Q = mH_v$.

Practice

Questions

The following video explains heat of fusion and vaporization. Use this video to answer the questions that follow.

http://www.youtube.com/watch?v=6lAxBTLgYfU



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Click image to the left for more content.

1. For water, which takes more energy, melting or evaporating?
2. When are there two phases present at the same time in the pot?

Practice problems involving phase changes:

http://www.slideshare.net/angel4all1/specific-heat-and-phase-change-ditto

Review

Questions

1. A 200. g sample of water at 60.0°C is heated to water vapor at 140.0°C. How much heat was absorbed?
2. A 175 g lump of molten lead at its melting point (327°C) is placed into 55.0 g of water at 20.0°C. The specific heat of lead is 130 $J/kg \cdot ^\circ C$ and the H_f of lead is 20,400 J/kg.
   (a) When the lead has become solid but is still at the melting point, what is the temperature of the water?
   (b) When the lead and the water have reached equilibrium, what is the temperature of the mixture?

- **phase change:** When a substance changes from one state, or phase, of matter to another we say that it has undergone a change of state, or we say that it has undergone a change of phase.

- **heat of fusion:** The change in potential energy for the conversion of 1 mole or 1 gram of a solid to a liquid, at constant pressure and temperature and is usually denoted as ΔH_f.
- **heat of vaporization:** The change in potential energy for the conversion of 1 mole or 1 gram of a liquid to a gas, at constant pressure and temperature and is usually denoted as ΔH_v.

13.2 Heat of Vaporization

Lesson Objectives

- The student will be able to calculate energy changes during phase changes.
- The student will be able to explain the slopes of various parts of heating and cooling curves.

Introduction

In order to vaporize a liquid, heat must be added to raise the kinetic energy (temperature) to the phase change temperature and then more heat must be added to provide the necessary potential energy to separate the molecules in the liquid form to the gaseous form.

The Potential Energy Stored in Gases

The difference between the liquid phase and the gas phase of a substance is essentially the distance between the molecules. Since the molecules attract each other and they are separated by a greater distance in the gaseous phase than in the liquid phase, the molecules in gaseous phase possess more potential energy than in the liquid phase. When a substance changes from the liquid phase to the gaseous phase, work (the physics kind of work as in force times distance) must be done on the molecules to pull them away from each other. The work done separating the molecules is stored in the molecular structure as potential energy. If the molecules are allowed to go back together as in the liquid phase, the potential energy is released – exactly the same amount that was needed to separate the molecules. This potential energy stored in molecules in the gaseous phase is called the **heat of vaporization**. The heat of vaporization ($\triangle H_{\text{VAP}}$) for water is 540 calories/gram (2.26 kJ/g) at the normal boiling point. Because of the strength of the polar attractions holding water molecules together in the liquid form, water has a fairly high heat of vaporization. Ammonia, NH_3, and ethanol, C_2H_5OH, which are also polar molecules have heats of vaporization of 1.38 kJ/g and 0.84 kJ/g respectively.

Example

How much heat in kJ is necessary to vaporize 100. grams of ammonia at its boiling point?

Solution

Heat, $Q = \text{ (mass)}(\triangle H_{\text{VAP}}) = (100. \text{ g})(1.38 \text{ kJ/g}) = 138 \text{ kJ}$

The boiling point of ammonia is $-33°C$. It is very important to understand that the ammonia is at the boiling point before the heat of vaporization is added and after the heat of vaporization is added, the ammonia is STILL at the boiling point. All the energy involved in the heat of vaporization is absorbed by the substance as potential energy – none of it goes into kinetic energy and therefore, the temperature cannot change. A common question asked of students to determine if they understand this point is to ask which would produce a more severe burn, spilling boiling water at 100°C on your skin or being burned by gaseous water at 100°C? It may seem at first that since they are both at the same temperature, they would do the same damage, but in fact, the gaseous water would first deliver a tremendous amount of heat to your skin as the gas condensed to water (giving up the heat of vaporization) and after you had a burn from that, then you would have 100°C boiling water on the skin that was already burned.

Heating and Cooling Curves

The addition of heat before, during, and after a phase transition can be analyzed with the help of a heating curve. In the heating curve below, a sample of water at 20°C and 1.00 atm pressure had equal quantities of heat added to it per unit time.

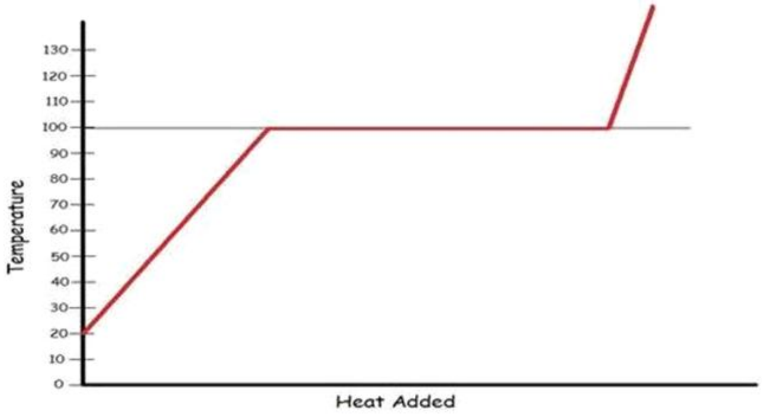


FIGURE 13.1

The heating curve for water at normal pressure.

Between the temperatures of 20°C and 100°C, all the heat added was absorbed as kinetic energy and therefore the temperature increased. Once the water reached the boiling point, even though the same amount of heat was added per unit time, the temperature did not increase. Thus, this energy added to the water DID NOT become kinetic energy. All the heat added to the sample during the time the slope of the line is zero went into potential energy. This energy represents the heat of vaporization and was used to do the work of separating the liquid molecules into the gaseous form (greater distance between molecules). During this flat line period, an observer would see that the water was changing into gas but all of it, both the part in liquid form and the part in gaseous form, would be at exactly 100°C. The temperature cannot increase until all the liquid has been converted to gas. At 1.00 atm pressure, it is impossible to get water hotter than 100°C. No matter how much heat you add to it, all that happens is that the water boils faster and converts to gas faster but its temperature will never exceed 100°C. It is sort of like the water decides where the heat goes, and the water decides that ALL the heat added will be used to change phase and none of it will be used to raise temperature. Once all the water is in the gaseous form, the heat once again becomes kinetic energy and the temperature of the gas rises. When the gas is cooled, it follows this same curve in reverse. As it is cooled, the same flat line appears while the heat of vaporization (heat of condensation now) is removed, and then the temperature may go down again.

Specific Heat

Thermodynamic data (melting point, boiling point, heat of melting, heat of vaporization) for almost all elements and thousands of compounds are determined by laboratory activity and placed in reference books. Much of this thermodynamic data is also available now on the internet. Another piece of useful thermodynamic data about substances is called **specific heat.** The specific heat for a substance is the amount of heat required to raise 1.00 gram of the substance by $1.00^\circ C$. The symbol C is used for specific heat and the value for liquid water is $4.18\ \text{J}/\text{g}^\circ C$.

Example

How much heat is required to raise the temperature of 25 grams of water from $15^\circ C$ to $55^\circ C$?

Solution

$$Q = (\text{mass})(C)(\triangle t) = (25\text{ g})(4.18\text{ J/g.}^\circ C)(40.^\circ C) = 4180\text{ J} = 4.18\text{ kJ}$$

Example

How much heat is required to raise 25 g of water from $25^\circ C$ to gaseous water at $100^\circ C$?

Solution

In this problem, you have to calculate the heat to raise the temperature from $25^\circ C$ to $100^\circ C$ and then calculate the heat of vaporization for the 25 g of water.

$$\begin{aligned}\text{Heat}_{\text{raising temp}} &= (\text{mass})(\text{specific heat})(\triangle t) = (25\text{ g})(4.18\text{ J/g.}^\circ C)(75^\circ C) = 7840\text{ J} = 7.48\text{ kJ}\\ \text{Heat}_{\text{vaporizing}} &= (\text{mass})(\text{heat of vaporization}) = (25\text{ g})(2.26\text{ kJ/g}) = 57\text{ kJ}\\ \text{Total heat} &= 7.48\text{ kJ} + 57\text{ kJ} = 64\text{ kJ}\end{aligned}$$

Lesson Summary

- The heat of vaporization of a liquid is the quantity of heat required to vaporize a unit mass of that liquid at constant temperature.
- The energy released when a gas condenses to a liquid is called the heat of condensation.
- The specific heat of a substance is the amount of heat required to raise the temperature of one gram of the substance by $1.0^\circ C$.

Review Questions

1. How much heat is required to vaporize 200. grams of water at $100.^\circ C$ and 1.00 atm pressure? $\triangle H_{\text{VAP}}$ for water is $2.25\ \text{kJ}/\text{g}$.
2. How much heat is required to raise 80.0 grams of water from $0^\circ C$ to $100.^\circ C$ with no phase change occurring? The specific heat of water is $4.18\ \text{J}/\text{g}.^\circ C$.

Further Reading / Supplemental Links

- Chemistry, Matter and Its Changes, Fourth Edition, Chapter 12: Intermolecular Attractions and the Properties of Liquids and Solids, James E. Brady and Fred Senese, John Wiley Sons, Inc. 2004. http://www.sparknotes.com/testprep/books/sat2/chemistry/chapter5section5.rhtml
- http://www.kentchemistry.com/links/Matter/HeatingCurve.htm

Vocabulary

heat of condensation
The quantity of heat released when a unit mass of a vapor condenses to liquid at constant temperature.

heat of vaporization
The quantity of heat required to vaporize a unit mass of liquid at constant temperature.

13.3 Heats of Fusion and Solidification

- Define heat of fusion.
- Define heat of solidification.
- Perform calculations of energy changes during the transition from solid to liquid or liquid to solid.

What makes the ice melt?

Suppose you hold an ice cube in your hand. It feels cold because heat energy leaves your hand and enters the ice cube. What happens to the ice cube? It melts. However, the temperature during a phase change remains constant. So the heat that is being lost by your hand does not raise the temperature of the ice above its melting temperature of 0°C. Rather, all the heat goes into the change of state. Energy is absorbed during the process of changing ice into water. The water that is produced also remains at 0°C until all of the ice is melted.

Heats of Fusion and Solidification

All solids absorb heat as they melt to become liquids. The gain of heat in this endothermic process goes into changing the state rather than changing the temperature. The **molar heat of fusion** (ΔH_{fus}) of a substance is the heat absorbed by one mole of that substance as it is converted from a solid to a liquid. Since the melting of any substance absorbs heat, it follows that the freezing of any substance releases heat. The **molar heat of solidification** $(\Delta H_{\text{solid}})$ of a substance is the heat released by one mole of that substance as it is converted from a liquid to a solid. Since fusion and solidification of a given substance are the exact opposite processes, the numerical value of the molar heat of fusion is the same as the numerical value of the molar heat of solidification, but opposite in sign. In other words, $\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$. The **Figure** 13.2 shows all of the possible changes of state along with the direction of heat flow during each process.

Every substance has a unique value for its molar heat of fusion, depending on the amount of energy required to disrupt the intermolecular forces present in the solid. When 1 mol of ice at 0°C is converted to 1 mol of liquid water

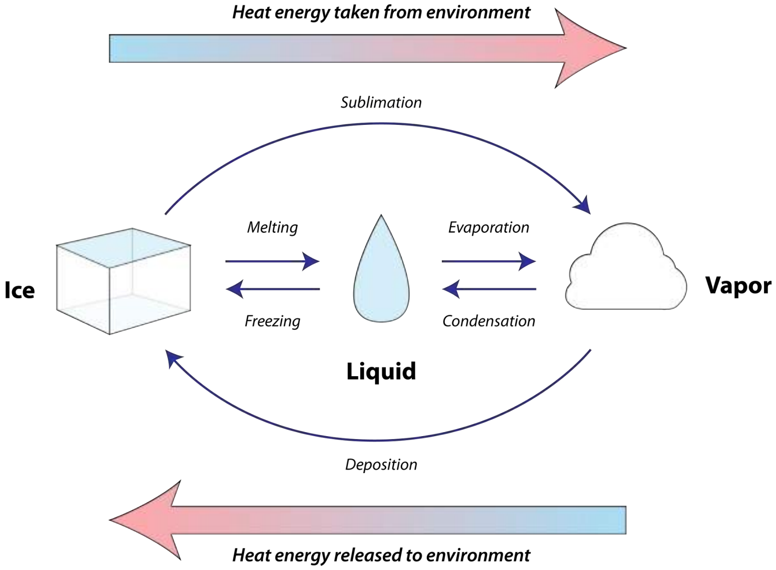


FIGURE 13.2

From left to right, heat is absorbed from the surroundings during melting, evaporation, and sublimation. Form right to left, heat is released to the surroundings during freezing, condensation, and deposition.

at 0°C, 6.01 kJ of heat are absorbed from the surroundings. When 1 mol of water at 0°C freezes to ice at 0°C, 6.01 kJ of heat are released into the surroundings.

$$\begin{aligned} \mathrm{H_2O}(s) &\rightarrow \mathrm{H_2O}(l) & \Delta H_{\text{fus}} &= 6.01 \text{ kJ/mol} \\ \mathrm{H_2O}(l) &\rightarrow \mathrm{H_2O}(s) & \Delta H_{\text{fus}} &= -6.01 \text{ kJ/mol} \end{aligned}$$

The molar heats of fusion and solidification of a given substance can be used to calculate the heat absorbed or released when various amounts are melted or frozen.

Enjoy the video about the heat of fusion at the link below:

MEDIA

Click image to the left for more content.

http://www.youtube.com/watch?v=jaaGqui9NVY

Sample Problem Heat of Fusion

Calculate the heat absorbed when 31.6 g of ice at 0°C is completely melted.

Step 1: List the known quantities and plan the problem.

Known

- mass = 31.6 g ice
- molar mass $H_2O(s)$ = 18.02 g/mol

- molar heat of fusion = 6.01 kJ/mol

Unknown

- $\Delta H = ?$ J

The mass of ice is first converted to moles. This is then multiplied by the conversion factor of $\left(\frac{6.01 \text{ kJ}}{1 \text{ mol}}\right)$ in order to find the kJ of heat absorbed.

Step 2: Solve.

$$31.6 \text{ g ice} \times \frac{1 \text{ mol ice}}{18.02 \text{ g ice}} \times \frac{6.01 \text{ kJ}}{1 \text{ mol ice}} = 10.5 \text{ kJ}$$

Step 3: Think about your result.

The given quantity is a bit less than 2 moles of ice, and so just less than 12 kJ of heat is absorbed by the melting process.

Summary

- Molar heats of fusion and solidification are defined.
- Calculations of heat changes during fusion and solidification are described.

Practice

Work problems 4-5 at the link below:

http://ths.sps.lane.edu/chemweb/unit4/problems/heatcalc/index.htm

Review

1. In the transition from liquid to solid, is energy absorbed or released?
2. In the transition from solid to liquid, is energy absorbed or released?
3. How much energy is released when one mole of water at 0°C changes from liquid to solid?

- **molar heat of fusion** (ΔH_{fus}) **:** The heat absorbed by one mole of that substance as it is converted from a solid to a liquid.
- **molar heat of solidification** $(\Delta H_{\text{solid}})$ **:** The heat released by one mole of that substance as it is converted from a liquid to a solid.

13.4 Vapor Pressure Curves

One of the first lessons in cooking is how to boil water. Yes, it sounds simple, but there are a couple of hints that speed things up. One hint is to put a lid on the pot. The picture above has water boiling uncovered with the steam escaping to the atmosphere. If the lid is on the pot, less water will be boiled off and the water will boil faster. The buildup of pressure inside the pot helps speed up the boiling process.

Vapor Pressure Curves

The boiling points of various liquids can be illustrated in a **vapor pressure curve** (**Figure** 13.3). A vapor pressure curve is a graph of vapor pressure as a function of temperature. To find the normal boiling point of a liquid, a horizontal line is drawn from the left at a pressure equal to standard pressure. At whatever temperature that line intersects the vapor pressure curve of a liquid is the boiling point of that liquid.

The boiling points of liquid also correlate to the strength of the intermolecular forces. Recall that diethyl ether has weak dispersion forces, which meant that the liquid has a high vapor pressure. The weak forces also mean that it does not require a large an input of energy to make diethyl ether boil and so it has a relatively low normal boiling point of 34.6°C. Water, with its much stronger hydrogen bonding, has a low vapor pressure and a higher normal boiling point of 100°C.

As stated earlier, boiling points are affected by external pressure. At higher altitudes, the atmospheric pressure is lower. With less pressure pushing down on the surface of the liquid, it boils at a lower temperature. This can also be seen from the vapor pressure curves. If one draws a horizontal line at a lower vapor pressure, it intersects each curve at a lower temperature. The boiling point of water is 100°C at sea level, where the atmospheric pressure is standard. In Denver, Colorado at 1600 m above sea level, the atmospheric pressure is about 640 mmHg and water boils at about 95°C. On the summit of Mt. Everest the atmospheric pressure is about 255 mmHg and water boils at only

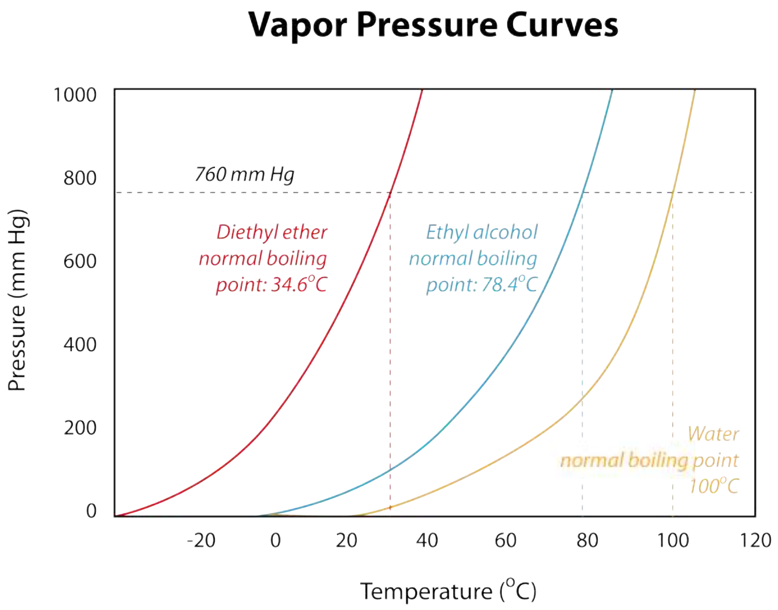


FIGURE 13.3
Vapor pressure curves.

70°C. On the other hand, water boils at greater than 100°C if the external pressure is higher than standard. Pressure cookers do not allow the vapor to escape and the vapor pressure increases. Since water now boils at a temperature above 100°C, the food cooks more quickly.

FIGURE 13.4
Pressure cooker.

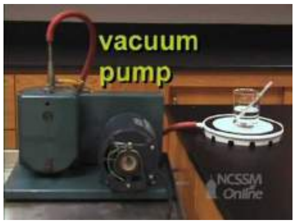
The effect of decreased air pressure can be demonstrated by placing a beaker of water in a vacuum chamber. At a low enough pressure, about 20 mmHg, water will boil at room temperature.

Summary

- A vapor pressure curve is a graph of vapor pressure as a function of temperature.
- Boiling points are affected by external pressure.

Practice

Use the video below to answer the following questions:



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Click image to the left for more content.

http://www.youtube.com/watch?v=Cshd5MVGpfk&feature=player_embedded .

1. What does a vacuum pump do?
2. What is the role of the O-ring?
3. What happens when the vacuum pump is turned on?
4. What happened when the vacuum pump was turned on?
5. What was the temperature of this boiling water?

Review

1. What does a vapor pressure curve show?
2. Why does diethyl ether have a low boiling point?
3. What intermolecular forces hold water molecules together?
4. Why does water boil at a lower temperature when at a high altitude?

- **vapor pressure curve:** A graph of vapor pressure as a function of temperature.

13.5 General Phase Diagram

- Define phase diagram.
- Define the triple point.
- Describe how to sue the phase diagram to indicate the state of a material and different temperatures and pressures.

Many rockets use a combination of kerosene and liquid oxygen for their fuel. Oxygen can be reduced to the liquid state either by cooling or by using high pressure. Since the oxygen is in a container essentially out in the open, maintaining a temperature of -183°C (the boiling point of oxygen) is not real practical. But high pressure can be used to force the oxygen into tanks and cause it to liquefy so it can then mix with the kerosene and provide a powerful ignition to move the rocket.

Phase Diagrams

The relationships among the solid, liquid, and vapor (gas) states of a substance can be shown as a function of temperature and pressure in single diagram. **A phase diagram** is graph showing the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas phases. Examine the general phase diagram shown in the **Figure** 13.5. In each of the three colored regions of the diagram, the substance is in a single state (or phase). The dark lines that act as the boundary between those regions represent the conditions under which the two phases are in equilibrium.

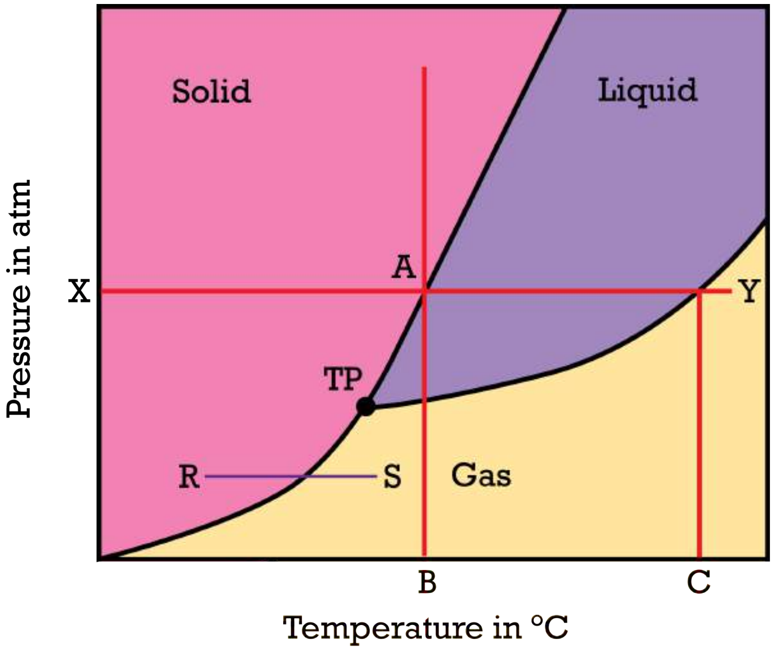

FIGURE 13.5

General phase diagram, which shows the state (phase) of a substance as a function of its temperature and pressure.

Find the X on the pressure axis and presume that the value of X is standard pressure of 1 atm. As one moves left to right across the red line, the temperature of the solid substance is being increased while the temperature remains constant. When point A is reached, the substance melts and the temperature B on the horizontal axis represents the normal melting point of the substance. Moving further to the right, the substance boils at point Y and so point C on the horizontal axis represents the normal boiling point of the substance. As the temperature increases at a constant pressure, the substance changes from solid to liquid to gas.

Start right above point B on the temperature axis and follow the red line vertically. At very low pressure, the particles of the substance are far apart from one another and the substance is in the gas state. As the pressure is increased, the particles of the substance are forced closer and closer together. Eventually the particles are pushed so close together that attractive forces cause the substance to condense into the liquid state. Continually increasing the pressure on the liquid will eventually cause the substance to solidify. For the majority of substances, the solid state is denser than the liquid state and so putting a liquid under great pressure will cause it to turn into a solid. The line segment $R-S$ represents the process of sublimation, where the substance changes directly from a solid to a gas. At a sufficiently low pressure, the liquid phase does not exist. The point labeled TP is called the **triple point**. The triple point is the one condition of temperature and pressure where the solid, liquid, and vapor states of a substance can all coexist at equilibrium.

Summary

- A phase diagram is a graph which the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas.
- The triple point is the one condition of temperature and pressure where the solid, liquid, and vapor states of a substance can all coexist at equilibrium.

Practice

Use the link below to answer the following questions:

http://chemwiki.ucdavis.edu/Physical_Chemistry/Physical_Properties_of_Matter/Phase_Transitions/Phase_Diagrams

1. What temperatures and pressures favor the formation of a solid?
2. What temperatures and pressures favor the formation of a gas?
3. What does the line *BC* represent?
4. What does the line *BD* represent?

Review

1. What is a phase diagram?
2. What is the triple point?
3. What does point *C* represent?

- **phase diagram:** A graph which the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas.
- **triple point:** The one condition of temperature and pressure where the solid, liquid, and vapor states of a substance can all coexist at equilibrium.

13.6 Phase Diagram for Water

- Describe the phase diagram for water.
- Define critical temperature.
- Define critical pressure.

How is it possible to create snowballs?

You need a special snow to make the best snowballs. This snow needs to be a little wet so the particles will stick together. Dry snow can be tightly pressed and will form snowballs because the higher pressure causes the snowflakes to melt somewhat. However, when you release the pressure, the snow goes back to a more solid form and the flakes no longer stick together.

Phase Diagram for Water

Water is a unique substance in many ways. One of these special properties is the fact that solid water (ice) is less dense than liquid water just above the freezing point. The phase diagram for water is shown below.

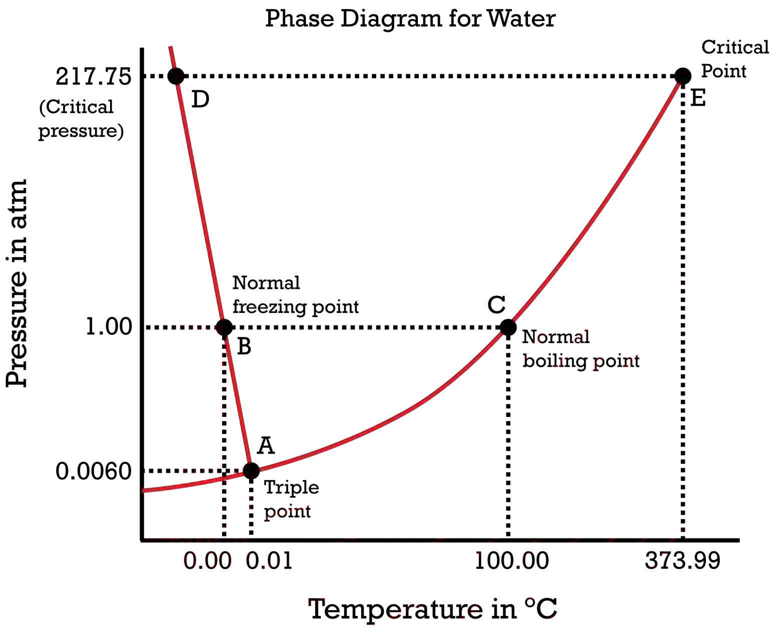

FIGURE 13.6

Phase diagram for water.

Notice one key difference between the general phase diagram and the phase diagram for water. In water's diagram, the slope of the line between the solid and liquid states is negative rather than positive. The reason is that water is an unusual substance in that its solid state is less dense than the liquid state. Ice floats in liquid water. Therefore, a pressure change has the opposite effect on those two phases. If ice is relatively near its melting point, it can be changed into liquid water by the application of pressure. The water molecules are actually closer together in the liquid phase than they are in the solid phase.

Refer again to water's phase diagram (**Figure** 13.6). Notice point E, labeled the **critical point**. What does that mean? At 373.99°C, particles of water in the gas phase are moving very, very rapidly. At any temperature higher than that, the gas phase cannot be made to liquefy, no matter how much pressure is applied to the gas. The **critical pressure** (P_c) is the pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid. For water, the critical pressure is very high, 217.75 atm. The critical point is the intersection point of the **critical temperature** and the critical pressure.

Summary

- Solid water is less dense than liquid water just above the freezing point.
- The critical temperature (T_c) of a substance is the highest temperature at which the substance can possibly exist as a liquid.
- The critical pressure (P_c) is the pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid.
- The critical point is the intersection point of the critical temperature and the critical pressure.

Practice

Read the section of the material at the link below that deals specifically with water:

http://www.chemguide.co.uk/physical/phaseeqia/phasediags.html

1. Why does the melting point of ice get lower at higher pressures?

2. Can you compress steam to form a liquid if the temperature is above 374°C?
3. What is the critical pressure of water?

Review

1. What happens to solid ice under high pressure near the freezing point?
2. What is the critical temperature?
3. What is the critical pressure?

- **critical point:** The intersection point of the critical temperature and the critical pressure.
- **critical pressure:** The pressure that must be applied to the gas at the critical temperature in order to turn it into a liquid.
- **critical temperature:** The highest temperature at which the substance can possibly exist as a liquid.

13.7 References

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CHAPTER 14

Gas Laws

Chapter Outline

- **Chm2.1.5**

- Identify characteristics of ideal gases.
- Apply general gas solubility characteristics.
- Apply the following formulas and concepts of kinetic molecular theory.

  1. 1 mole of any gas at STP=22.4 L
  2. Ideal gas equation (PV=nRT), Combined gas law (P1V1/T1 = P2V2/T2) and applications holding one variable constant: for PV=k,
     P1V1 = P2V2; for V/T=k, V1/T1= V2/T2; for P/T=k, P1/T1 = P2/T2. Note: Students should be able to derive and use these gas laws,
     but are not necessarily expected to memorize their names.
  3. Avogadro's law (n/V=k), n1/V1 = n1/V2
  4. Dalton's law (Pt=P1+P2+P3 ...)
  5. Vapor pressure of water as a function of temperature (conceptually).

14.1 Gas Laws

Lesson Objectives

- Describe various observed relationships between the pressure, volume, temperature, and amount of a gas, including Boyle's law, Charles's law, Gay-Lussac's law, and Avogadro's law. Be able to perform calculations using these relationships.
- Be able to derive a combined gas law, and utilize this expression to calculate volume, temperature, and pressure changes in gas systems.
- Describe the ideal gas equation PV=nRT and perform calculations using this relationship.

Lesson Vocabulary

- **Boyle's Law**: States that for a fixed amount of gas at a constant temperature, the volume occupied is inversely proportional to its pressure.
- **absolute zero**: The lower limit for possible temperatures, a value of 0 Kelvin or -273°C.
- **Charles's Law**: States that for a fixed amount of gas at a constant pressure, the volume is directly proportional to its absolute temperature.
- **Gay-Lussac's Law**: States that pressure is directly proportional to the absolute temperature.
- **Avogadro's Law**: Stats that if you compare two samples of an ideal gas at the same temperature, pressure, and volume, they contain the same number of molecules.
- **ideal gas law**: The equation of state that describes the relationship between temperature, pressure, volume, and amount of an ideal gas. Derived from a combination of Boyle's, Charles's, Gay-Lussac's and Avogadro's gas laws.

Check Your Understanding

- One of the earliest air pumps was made by Robert Boyle. **Figure** 14.1 is a diagram of his early pump.

What do you suppose the relationship is between pressure and volume as shown here?

Introduction

In the last lesson, we studied the properties and characteristics of gases. Many of these properties can be explained by the fact that gases are mostly empty space but contain large numbers of particles with certain masses and velocities. Now we are going to study how gases behave when exposed to changes in temperature, pressure, and volume. We will also see how the development of instruments that measure pressure and temperature led to new discoveries in our understanding of gases. The origins of the absolute temperature scale will be discussed, and we will examine the empirical relationships between volume, temperature, pressure, and amount that were discovered by early chemists.

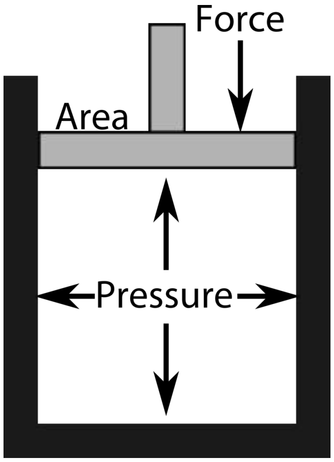


FIGURE 14.1
Air pump made by Robert Boyle.

FIGURE 14.2
Soap bubble in flight.

Boyle's Law: Pressure vs. Volume

With the development of the barometer in 1644 came a series of experiments and investigations aimed at understanding the nature of gases. Robert Boyle (1627-1691) was particularly interested in the relationships between the volume occupied by a sample of gas and the pressure it exerted on its container. He invented one of the earliest air pumps, which allowed him to create low-pressure environments. Boyle devised experiments to study the pressure-volume relationship. For example, he observed that reducing the volume of a balloon by half doubles its volume. In other words, volume is inversely proportional to pressure ($V \propto 1/P$).

For a fixed amount of gas at a constant temperature, the volume occupied is inversely proportional to its pressure. Stated mathematically, the product of the pressure and volume is equal to a constant:

$P \times V = k$

where k is a constant that depends on the temperature and amount of the gas in a given sample. This observation was published in 1662 and is now known as **Boyle's Law**.

This equation can be used to quantitatively predict how changes in pressure will affect the volume of a sample, and vice versa. Let's assume that a sample of gas undergoes a change to one of these variables while the temperature and amount are held constant. The initial pressure (P_1) and the initial volume (V_1) can be related to the final pressure (P_2) and the final volume (P_2) as follows:

$P_1 \times V_1 = k$

$P_2 \times V_2 = k$

Therefore,

$P_1 \times V_1 = P_2 \times V_2$

Example 14.3

A sample of oxygen gas has a volume of 150. mL when its pressure is 0.647 atm. What will the volume of the gas be at a pressure of 0.987 atm if the temperature remains constant?

Answer

$$\begin{aligned}
&P_1 \times V_1 = P_2 \times V_2 \\
&V_1 = 150.\text{ mL}, V_2 = ? \\
&P_1 = 0.647\text{ atm}, P_2 = 0.987\text{ atm} \\
&V_2 = \frac{P_1 \times V_1}{P_2} \\
&V_2 = \frac{(0.647\ \text{atm})(150.\ \text{mL})}{0.987\ \text{atm}} \\
&V_2 = 98.3\ \text{mL}
\end{aligned}$$

Temperature

The thermometer was developed a short time after the barometer made it possible to study pressure in a quantitative manner. Although crude thermometers had been available for quite some time, Daniel Fahrenheit constructed a more accurate thermometer in 1724 using mercury in glass. Because the height of the mercury in the thermometer was a relative measure, certain fixed points needed to be defined. Originally, Fahrenheit defined a 1:1:1 mixture of ice, salt, and ammonium chloride as 0°F, and the freezing/melting point of ice as 32°F. On this scale, the normal temperature of the human body is 96°F, and the boiling point of water is 212°C. The modern Fahrenheit scale is defined in terms of the normal freezing and boiling points of water.

Anders Celsius (1736) devised a slightly different scale, which also had two fixed values based on the freezing and boiling points of water. Originally, he suggested a value of 0°C for the boiling point and 100°C for the freezing point. Shortly after his death, the scale was reversed, resulting in the form that is used today. The Celsius scale is the standard way to measure temperature in most parts of the world. **Figure** 14.3 compares these two different scales.

Absolute Zero

The ideas of absolute zero and absolute temperature were conceived in 1848 by Lord Kelvin. Kelvin had observed that for every 1°C drop in temperature, a sample of gas contracted by a uniform amount, specifically, 1/273th of its volume at 0°C. If this data is extrapolated, or projected, to the point where the gas would reach a volume of zero, we reach a lower limit for possible temperatures. Kelvin interpreted this to mean that -273°C is the lowest achievable temperature. In contrast to other scales, where the temperature can be below zero, the value of zero Kelvin represented an **absolute zero**. **Figure** 14.4 Below shows Kelvin's extrapolation. The Kelvin scale is used for all gas law calculations.

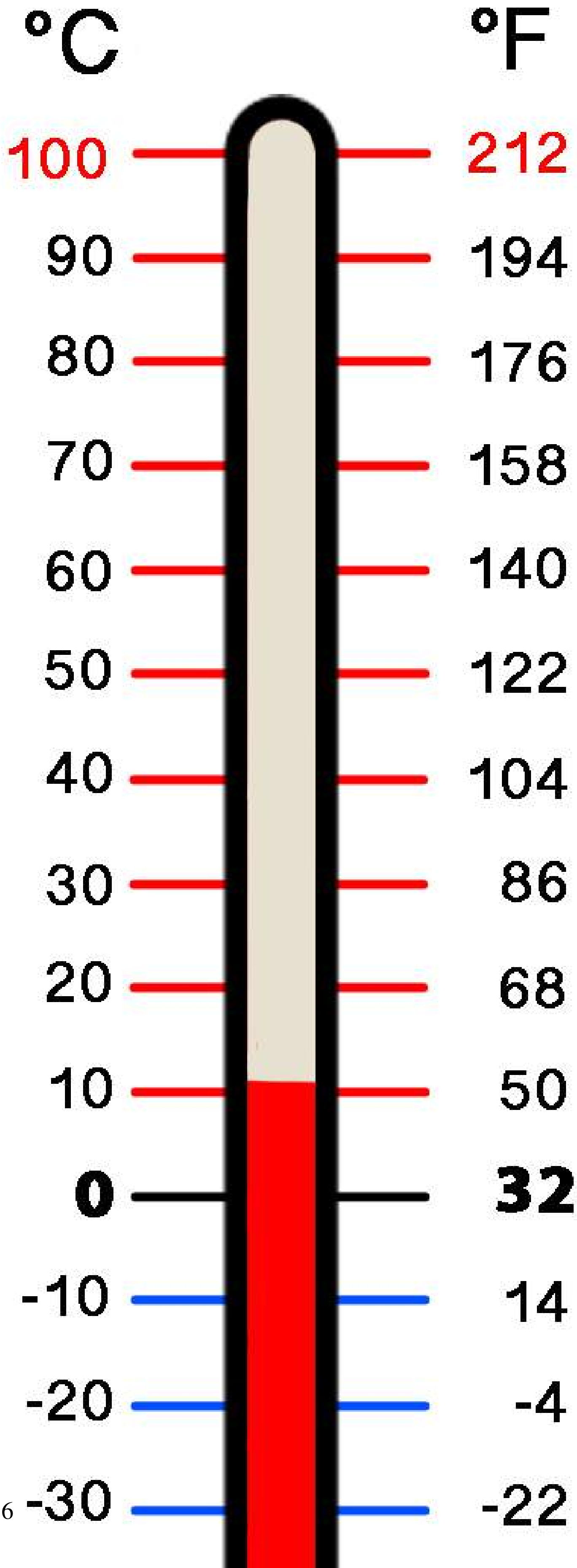

°C
°F
100
212
90
194
80
176
70
158
60
140
50
122
40
104
30
86
20
68
10
50
0
32
-10
14
-20
-4
-30
-22

FIGURE 14.3

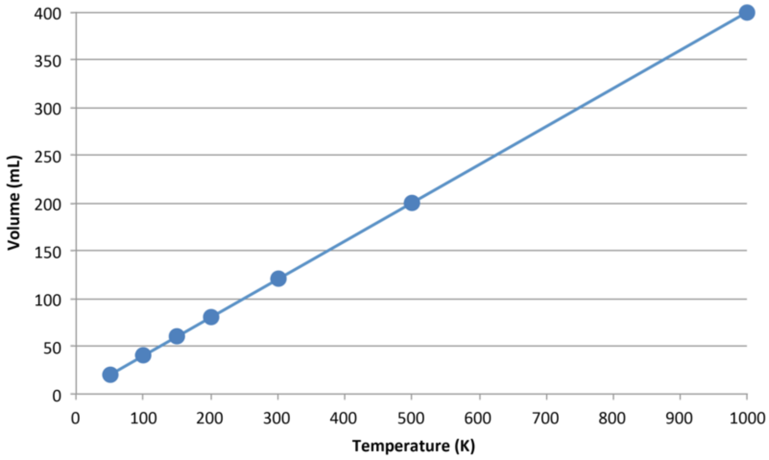


FIGURE 14.4

Temperature Relationships

Charles's Law

Various studies were done on the effects of temperature changes on both pressure and volume by two prominent French scientists, Jacques Charles (1746-1823) and Joseph Louis Gay-Lussac (1778-1850). In 1802, Charles and Gay-Lussac both proposed that for a fixed amount of gas at a constant pressure, the volume is directly proportional to its absolute temperature ($V \propto T$). **Figure** 14.5 illustrates the volume-temperature relationship.

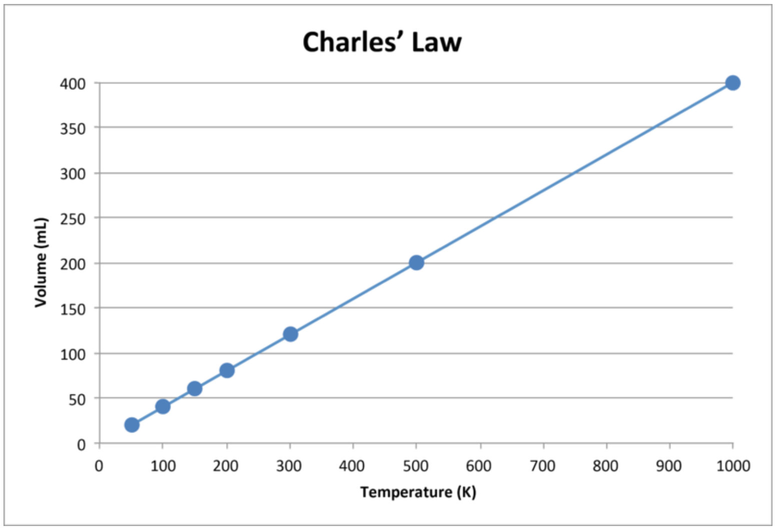


FIGURE 14.5

This relationship between volume and pressure is commonly known as **Charles's Law**. It can be expressed mathematically as follows:

$\frac{V}{T} = k$

where k is a constant that depends on the pressure and amount of the gas sample. As in the case of Boyle's Law, this can also be used to generate a way of comparing two different states for a given sample at constant pressure:

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Example 14.4

A sample of neon gas occupies a volume of 752 mL at 25°C. What volume will the gas occupy at 50°C if the pressure remains constant? Before you begin, think about whether or not the volume should go up or down.

Answer:

Don't forget that all temperatures must be converted to absolute temperature (Kelvin).

$$\begin{aligned}
T_1 &= 25^\circ\text{C} + 273 = 298\text{ K} \\
T_2 &= 50^\circ\text{C} + 273 = 323\text{ K} \\
V_1 &= 752\text{ mL} \\
V_2 &= ? \\
\frac{V_1}{T_1} &= \frac{V_2}{T_2} \\
V_2 &= \frac{V_1 T_2}{T_1} \\
V_2 &= \frac{(752\text{ mL})(323\text{ K})}{298\text{ K}} \\
V_2 &= 815\text{ mL}
\end{aligned}$$

Gay-Lussac's Law

Gay-Lussac's Law describes the relationship between pressure and absolute temperature:

$\frac{P}{T} = k$

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Gay-Lussac's law says that at a constant volume, the pressure of a gas is directly proportional to its absolute temperature. This relationship is also sometimes attributed to Guillaume Amontons (1663–1705). Although his work predated that of Gay-Lussac by over a century, the data on which he based his ideas were considerably less precise, due to the cruder instrumentation available at the time.

Example 14.5

The gas in an aerosol can exerts a pressure of 3.00 atm at 25°C. Directions on the can warn the user not to keep the can in a place where the temperature exceeds 52°C. What would the gas pressure in the can be at 52°C? Before you begin, do you think the pressure should go up, down or stay the same?

Answer:

Again, don't forget that all temperatures must be converted to absolute temperature (Kelvin).

$$
\begin{aligned}
T_1 &= 25^\circ\text{C} + 273 = 298\text{ K} \\
T_2 &= 52^\circ\text{C} + 273 = 325\text{ K} \\
P_1 &= 3.00\text{ atm} \\
P_2 &= ? \\
\frac{P_1}{T_1} &= \frac{P_2}{T_2} \\
P_2 &= \frac{P_1 T_2}{T_1} \\
P_2 &= \frac{(3.00\text{ atm})(325\text{ K})}{298\text{ K}} \\
P_2 &= 3.27\text{ atm}
\end{aligned}
$$

Note that the pressure increases, which would be expected if the temperature were to go up.

Combined Gas Law

We can combine the relationships described in Boyle's law, Charles's law, and Gay-Lussac's law to create a combined gas law that relates pressure, volume, and temperature.

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

We can use this expression to predict changes to these variables for a given chemical system.

Example 14.6

A helium-filled balloon has a volume of 50.0 L at 25°C and 1.08 atm. What volume will it have at 0.855 atm and 10.°°C?

$$
\begin{aligned}
P_1 &= 1.08\text{ atm}, P_2 = 0.855\text{ atm} \\
V_1 &= 50.0\text{ L}, V_2 = ? \\
T_1 &= 25^\circ\text{C} + 273 = 298\text{ K}, T_2 = 10^\circ\text{C} + 273 = 283\text{ K} \\
V_2 &= \frac{P_1 T_2 V_1}{P_2 T_1} \\
V_2 &= \frac{(1.08\text{ atm})(283\text{ K})(50.0\text{ L})}{(0.855\text{ atm})(298\text{ K})} = 60.0\text{ L}
\end{aligned}
$$

Avogadro's Law

In 1811, shortly after the work of Charles and Gay-Lussac, Amadeo Avogadro (1776-1856) was also studying the relationships between pressure, volume, and temperature. He postulated that if you compare two samples of an ideal gas at the same temperature, pressure, and volume, they contain the same number of molecules. This became known as **Avogadro's Law**. This means that at a given pressure and temperature, the volume of a sample is directly proportional to the number of molecules present.

$\frac{V}{n} = k$

Where:

V is the volume of the gas

n is the amount of gas particles (generally measured in moles)

k is a proportionality constant

Ideal Gas Law

If we add Avogadro's Law into the combined gas law, we get the following equation:

$\frac{PV}{Tn} = constant$

Unlike the constants for the other gas laws, this constant is given a special name. The ideal gas constant (R) has the same value for any sample of any ideal gas. A more common way to write the **ideal gas law** is the following:

$PV = nRT$

Recall that 1 mol of an ideal gas at 1 atm of pressure and 0°C (standard temperature and pressure) occupies a volume of 22.4 L. We can use this information to calculate the value of R:

$R = \frac{P \times V}{T \times n} = \frac{(1 \text{ atm})(22.4 \text{ L})}{(273 \text{ K})(1 \text{ mol})} = 0.0821 \text{ L} \cdot \text{atmK}^{-1} \cdot \text{mol}^{-1}$

Using more precise measurements, the ideal gas constant is found to have a value of 0.08206 L•atm•K^{-1}•mol^{-1}.

Example 14.7

Sulfur hexafluoride (SF_6) is a colorless, odorless, and very unreactive gas. Calculate the pressure (in atm) exerted by 1.82 moles of the gas at 69.5°C when confined to a rigid steel vessel with a volume of 5.43 L.

Answer:

$$\begin{aligned}
T_1 &= 69.5^\circ\text{C} + 273 = 342.5 \text{ K} \\
P &= \frac{nRT}{V} \\
P &= \frac{(1.82 \text{ mol})(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(342.5 \text{ K})}{5.43 \text{ L}} \\
P &= 9.42 \text{ atm}
\end{aligned}$$

Note that, in order for all of the units to cancel correctly, the pressure must be expressed in atmospheres, the volume must be in liters, the amount must be in moles, and the temperature must be in Kelvin. If different units were used, the ideal gas constant would have a different value.

Lesson Summary

- Pressure and volume are inversely proportional at a constant temperature.
- Volume and temperature are directly proportional at a constant pressure.
- Pressure and temperature are directly proportional at a constant volume.
- Pressure, volume, and temperature are related by the combined gas law.
- The volume occupied by a gas at a given temperature and pressure is proportional to the number of gas molecules present.
- For any ideal gas, PV=nRT, where R is a constant with a value of 0.08206 L•atm•K^{-1}•mol^{-1}.

Review Questions

1. A sample of a gas has a volume of 4.22 L at 27.8°C and 734 mmHg. What will its volume be at 24.6°C and 755 mmHg?
2. How many molecules of nitrogen gas are in a 5.20 L vessel at 136°C and 0.43 atm?
3. A 34.8 L tank of Ar gas is connected to an evacuated 87.5 L tank. If the final pressure is 640 mmHg, what must have been the original gas pressure in the 34.8 L tank?
4. A 54.0 L cylinder contains 798 g of chlorine gas at 29°C. What mass of chlorine must be released to reduce the pressure to 0.78 atm?
5. A sample of gas at 1 atm has a temperature of 15°C and a volume of 2.58 L. The temperature is then raised to 38°C at 1 atm.
    a. Does the volume of the gas increase or decrease?
    b. Calculate the new volume.
6. An automobile engine contains a vaporized sample of gasoline. The initial volume in the cylinder is 0.725 L. After compression, the volume is reduced to 0.075 L. The fuel mixture initially exerts a pressure of 1.00 atm. Calculate the pressure of the compressed fuel-air mixture (assuming temperature and the amount of gas remain constant).
7. If a gas is cooled from 333.0 K to 273.0 K at a constant volume, what would the final pressure be if the original pressure was 780.0 mm Hg?
8. A bag of potato chips is packaged at sea level (1.00 atm) and has a volume of 315 mL. If this bag of chips is transported to Denver (0.820 atm), what will the new volume of the bag be? Assume the temperature remains the same.
9. Calculate the volume (in L) occupied by 2.12 moles of nitric oxide (NO) at 6.54 atm and 76°C.

Further Reading/Supplemental Links

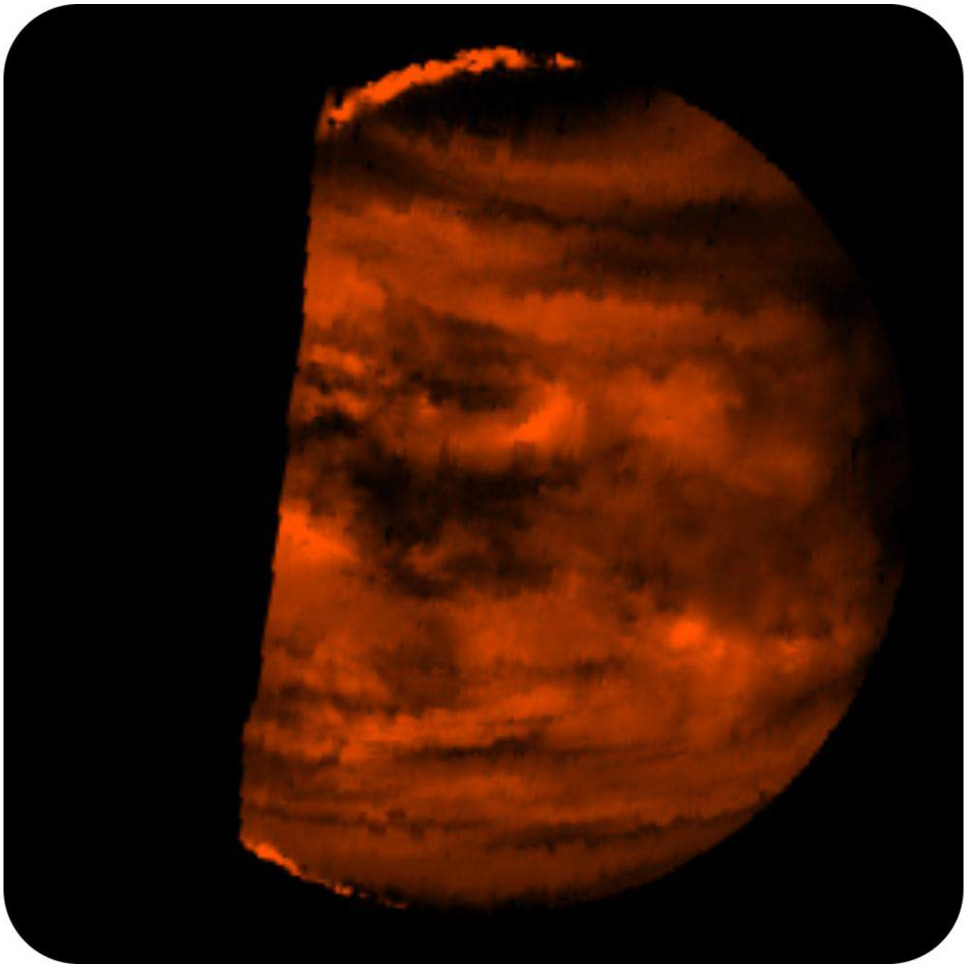
1. Charles's Law Practice Simulator: http://science.widener.edu/svb/tutorial/charleslaw.html
2. Nobel Prize 2001 for studies in absolute zero: http://www.colorado.edu/NewsServices/nobel/background.html
3. The gas laws: http://misterguch.brinkster.net/gaslaws.html

Points to Consider

- As we studied gases in this lesson, we assumed that all gases exhibit ideal behavior (collisions between particles are completely elastic, the particles do not attract or repel one another, and the volume of the gas particles is negligible compared to the volume they occupy). If some of these assumptions were not quite true, how would it affect the observed gas laws?
- What properties of gases are affected by the speeds of the individual particles?

14.2 Dalton's Law of Partial Pressures

- Define partial pressure.
- State Dalton's law of partial pressures.
- Use this law to calculate pressures of gas mixtures.



Is there oxygen available on Venus?

The atmosphere of Venus is markedly different from that of Earth. The gases in the Venusian atmosphere are 96.5% carbon dioxide and 3% nitrogen. The atmospheric pressure on Venus is roughly 92 times that of Earth, so the amount of nitrogen on Venus would contribute a pressure well over 2700 mm Hg. And there is no oxygen present, so we couldn't breathe there. Not that we would want to go to Venus –the surface temperature is usually over 460°C.

Dalton's Law of Partial Pressures

Gas pressure results from collisions between gas particles and the inside walls of their container. If more gas is added to a rigid container, the gas pressure increases. The identities of the two gases do not matter. John Dalton, the English chemist who proposed the atomic theory, also studied mixtures of gases. He found that each gas in a mixture exerts a pressure independently of every other gas in the mixture. For example, our atmosphere is composed of about 78% nitrogen and 21% oxygen, with smaller amounts of several other gases making up the rest. Since nitrogen makes up 78% of the gas particles in a given sample of air, it exerts 78% of the pressure. If the overall atmospheric pressure is 1.00 atm, then the pressure of just the nitrogen in the air is 0.78 atm. The pressure of the oxygen in the air is 0.21 atm.

The **partial pressure** of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture. The partial pressure of a gas is indicated by a P with a subscript that is the symbol or formula of that gas. The partial pressure of nitrogen is represented by P_{N_2}. **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases. Dalton's law can be expressed with the following equation:

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots$$

The figure below shows two gases that are in separate, equal-sized containers at the same temperature and pressure. Each exerts a different pressure, P_1 and P_2, reflective of the number of particles in the container. On the right, the two gases are combined into the same container, with no volume change. The total pressure of the gas mixture is equal to the sum of the individual pressures. If $P_1 = 300$ mmHg and $P_2 = 500$ mmHg, then $P_{\text{Total}} = 800$ mmHg.

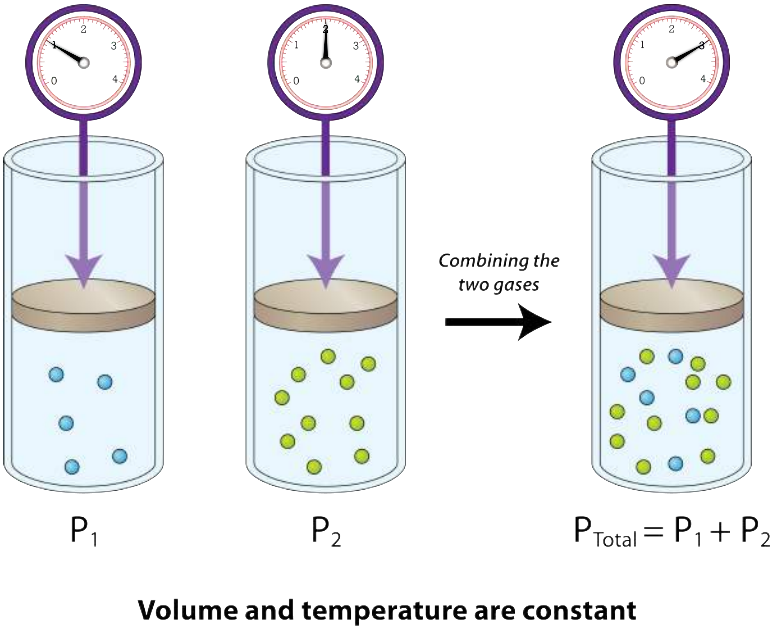


FIGURE 14.6
Dalton's law says that the pressure of a gas mixture is equal to the partial pressures of the combining gases.

Summary

- The total pressure in a system is equal to the sums of the partial pressures of the gases present.

Practice

Review the concepts at the link below and work the sample problems:

http://www.kentchemistry.com/links/GasLaws/dalton.htm

Review

1. What is the foundation for Dalton's law?
2. Argon makes up about 0.93% of our atmosphere. If the atmospheric pressure is 760 mm Hg, what is the pressure contributed by argon?
3. On a given day, the water vapor in the air is 2.5%. If the partial pressure of the vapor is 19.4 mm Hg, what is the atmospheric pressure?

- **Partial pressure:** The contribution that gas makes to the total pressure when the gas is part of a mixture.
- **Dalton's law of partial pressures:** The total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases.

14.3 Real and Ideal Gases

- Define a real gas.
- Describe differences between real gases and ideal gases.

Real and Ideal Gases

An ideal gas is one that follows the gas laws at all conditions of temperature and pressure. To do so, the gas would need to completely abide by the kinetic-molecular theory. The gas particles would need to occupy zero volume and they would need to exhibit no attractive forces what so ever toward each other. Since neither of those conditions can be true, there is no such thing as an ideal gas. A **real gas** is a gas that does not behave according to the assumptions of the kinetic-molecular theory. Fortunately, at the conditions of temperature and pressure that are normally encountered in a laboratory, real gases tend to behave very much like ideal gases.

Under what conditions then, do gases behave least ideally? When a gas is put under high pressure, its molecules are forced closer together as the empty space between the particles is diminished. A decrease in the empty space means that the assumption that the volume of the particles themselves is negligible is less valid. When a gas is cooled, the decrease in kinetic energy of the particles causes them to slow down. If the particles are moving at slower speeds, the attractive forces between them are more prominent. Another way to view it is that continued cooling the gas will eventually turn it into a liquid and a liquid is certainly not an ideal gas anymore (see liquid nitrogen in the **Figure** below). In summary, a real gas deviates most from an ideal gas at low temperatures and high pressures. Gases are most ideal at high temperature and low pressure.

FIGURE 14.7
Nitrogen gas that has been cooled to 77 K has turned to a liquid and must be stored in a vacuum insulated container to prevent it from rapidly vaporizing.

The **Figure** below shows a graph of $\frac{PV}{RT}$ plotted against pressure for 1 mol of a gas at three different temperatures - 200 K, 500 K, and 1000 K. An ideal gas would have a value of 1 for that ratio at all temperatures and pressures and the graph would simply be a horizontal line. As can be seen, deviations from an ideal gas occur. As the pressure begins to rise, the attractive forces cause the volume of the gas to be less than expected and the value of $\frac{PV}{RT}$ drops under 1. Continued pressure increase results in the volume of the particles to become significant and the value of $\frac{PV}{RT}$ rises to greater than 1. Notice, that the magnitude of the deviations from ideality is greatest for the gas at 200 K and least for the gas at 1000 K.

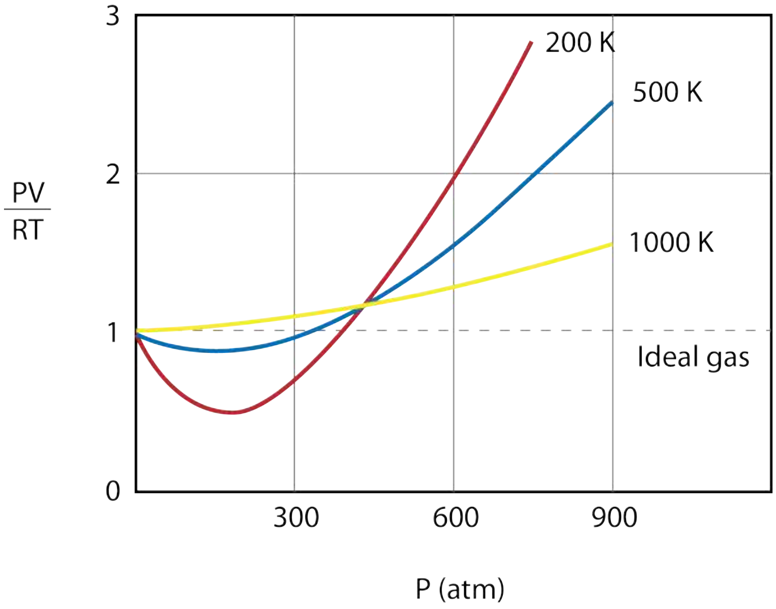


FIGURE 14.8

Real gases deviate from ideal gases at high pressures and at low temperatures.

The ideality of a gas also depends on the strength and type of intermolecular attractive forces that exist between the particles. Gases whose attractive forces are weak are more ideal than those with strong attractive forces. At the same temperature and pressure, neon is more ideal than water vapor because neon's atoms are only attracted by weak dispersion forces, while water vapor's molecules are attracted by relatively stronger hydrogen bonds. Helium is a more ideal gas than neon because its smaller number of electrons means that helium's dispersion forces are even weaker than those of neon.

Summary

- The properties of real gases and their deviations from ideality are described.

Practice

Use the link below to answer the following questions:

http://www.adichemistry.com/physical/gaseous/deviation/van-der-waals-equation.html

1. What is the compressibility factor for a perfect (ideal) gas?
2. What does it mean if $Z > 1$?
3. What does it mean if $Z < 1$?

Review

1. What becomes more significant as the pressure increases?
2. Do the attractive forces between gas particles become more prominent at higher or lower temperatures?
3. Would HCl gas be more or less ideal than helium?

- **real gas:** A gas that does not behave according to the assumptions of the kinetic-molecular theory.

14.4 Avogadro's Hypothesis and Molar Volume

- State Avogadro's hypothesis.
- Define standard temperature and pressure.
- Define molar volume.

How do scuba divers know if they will run out of gas?

Knowing how much gas is available for a dive is crucial to the survival of the diver. The tank on the diver's back is equipped with gauges to tell how much gas is present and what the pressure is. A basic knowledge of gas behavior allows the diver to assess how long to stay under water without developing problems.

Avogadro's Hypothesis and Molar Volume

Volume is a third way to measure the amount of matter, after item count and mass. With liquids and solids, volume varies greatly depending on the density of the substance. This is because solid and liquid particles are packed close together with very little space in between the particles. However, gases are largely composed of empty space between the actual gas particles.

In 1811, Amedeo Avogadro explained that the volumes of all gases can be easily determined. **Avogadro's hypothesis** states that equal volumes of all gases at the same temperature and pressure contain equal numbers of particles. Since the total volume that a gas occupies is made up primarily of the empty space between the particles, the actual size of the particles themselves is nearly negligible. A given volume of a gas with small light particles such as hydrogen (H_2) contains the same number of particles as the same volume of a heavy gas with large particles such as sulfur hexafluoride, SF_6.

Gases are compressible, meaning that when put under high pressure, the particles are forced closer to one another. This decreases the amount of empty space and reduces the volume of the gas. Gas volume is also affected

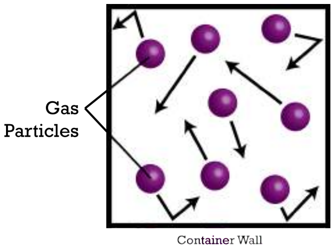


FIGURE 14.9

Gas particles are very small compared to the large amounts of empty space between them.

by temperature. When a gas is heated, its molecules move faster and the gas expands. Because of the variation in gas volume due to pressure and temperature changes, the comparison of gas volumes must be done at one standard temperature and pressure. **Standard temperature and pressure (STP)** is defined as 0°C (273.15 K) and 1 atm pressure. The **molar volume** of a gas is the volume of one mole of a gas at STP. At STP, one mole (6.02×10^{23} representative particles) of any gas occupies a volume of 22.4 L (**Figure** 14.10).

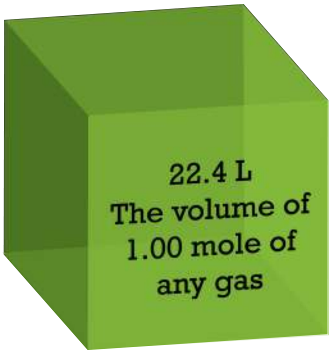


FIGURE 14.10

A mole of any gas occupies 22.4 L at standard temperature and pressure (0 °C and 1 atm).

The **Figure** 14.11 illustrates how molar volume can be seen when comparing different gases. Samples of helium (He), nitrogen (N_2), and methane (CH_4) are at STP. Each contains 1 mole or 6.02×10^{23} particles. However, the mass of each gas is different and corresponds to the molar mass of that gas: 4.00 g/mol for He, 28.0 g/mol for N_2, and 16.0 g/mol for CH_4.

Summary

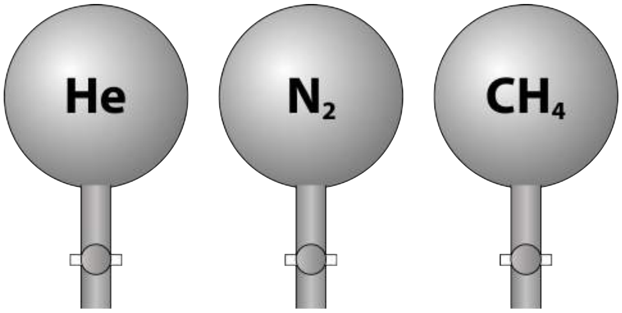
- Equal volumes of gases at the same conditions contain the same number of particles.
- Standard temperature and pressure are defined.

Practice

Use the link below to answer the following questions:

http://chemed.chem.purdue.edu/demos/main_pages/4.6.html

1. What was the volume of each gas that was weighed?
2. What did the experiment find?
3. What was the relationship between gas weight and molecular weight?



Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

FIGURE 14.11

Avogadro's hypothesis states that equal volumes of any gas at the same temperature and pressure contain the same number of particles. At standard temperature and pressure, 1 mole of any gas occupies 22.4 L.

Review

1. What do we know about the space actually taken up by a gas?
2. Why do we need to do all our comparisons at the same temperature and pressure?
3. How can we use this information?

- **Avogadro's hypothesis:** Equal volumes of all gases at the same temperature and pressure contain equal numbers of particles.
- **molar volume:** The volume of one mole of a gas at STP.
- **standard temperature and pressure (STP):** 0°C (273.15 K) and 1 atm pressure.

14.5 References

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CHAPTER 15 Water and Solutions

Chapter Outline

Chm.3.2.3

- Compute concentration (molarity) of solutions in moles per liter.
- Calculate molarity given mass of solute and volume of solution.
- Calculate mass of solute needed to create a solution of a given molarity and volume.
- Solve dilution problems: M1V1 = M2V2.

Chm.3.2.4

- Identify types of solutions (solid, liquid, gaseous, aqueous).
- Define solutions as homogeneous mixtures in a single phase.
- Distinguish between electrolytic and nonelectrolytic solutions.
- Summarize colligative properties (vapor pressure reduction, boiling point elevation, freezing point depression, and osmotic pressure).
  Chm.3.2.5

- Use graph of solubility vs. temperature to identify a substance based on solubility at a particular temperature.
- Use graph to relate the degree of saturation of solutions to temperature.
  Chm.3.2.6

- Develop a conceptual model for the solution process with a cause and effect relationship involving forces of attraction between solute and solvent particles. A material is insoluble due to a lack of attraction between particles.
- Describe the energetics of the solution process as it occurs and the overall process as exothermic or endothermic.
- Explain solubility in terms of the nature of solute-solvent attraction, temperature and pressure (for gases).

15.1 Physical Properties of Water

- Define surface tension.
- Define vapor pressure.
- Explain the phsyical properties of water in terms of hydrogen bonding.

What is this pan used for?

Water loss to the atmosphere is a significant problem in many parts of the world. When water supplies are low, anything that can be done to decrease water loss is important for farmers. An evaporation pan (seen above) can be used to measure how fast water evaporates in a given location. This information can be used as part of projects to develop ways to cut down on evaporation and increase the amount of usable water in a region.

Properties of Water

Compared to other molecular compounds of relatively low molar mass, ice melts at a very high temperature. A great deal of energy is required to break apart the hydrogen-bonded network of ice and return it to the liquid state. Likewise, the boiling point of water is very high. Most molecular compounds of similar molar mass are gases at room temperature.

Surface Tension

Water has a high **surface tension** (attraction between molecules at the surface of a liquid) because of its hydrogen bonding. Liquids that cannot hydrogen bond do not exhibit nearly as much surface tension. Surface tension can be seen by the curved meniscus that forms when water is in a thin column such as a graduated cylinder or a buret.

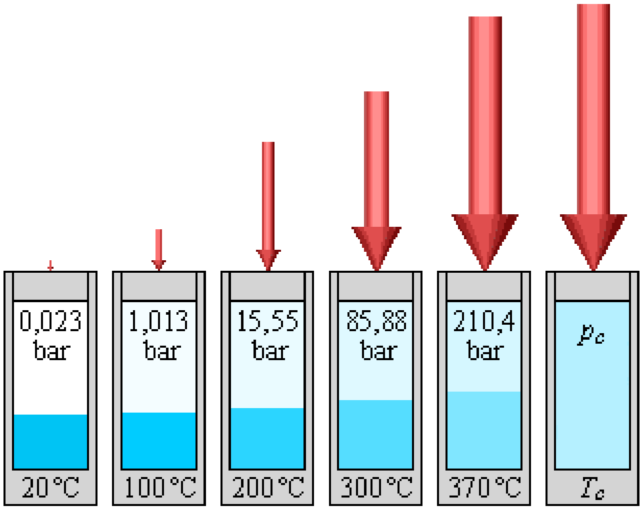
FIGURE 15.1

The meniscus of water in a graduated cylinder forms because of water's hydrogen bonding.

Vapor Pressure

The **vapor pressure** of a liquid is the pressure of the vapor produced by evaporation of a liquid or solid above the liquid or solid in a closed container. The hydrogen bonding between liquid water molecules explains why water has an unusually low vapor pressure. Relatively few molecules of water are capable of escaping the surface of the liquid and enter the vapor phase. Evaporation is slow and thus the vapor exerts a low pressure in a closed container. Low vapor pressure is an important physical property of water, since lakes, oceans, and other large bodies of water would all tend to evaporate much more quickly otherwise.

Vapor pressure is influenced by temperature. As the temperature increases, more molecules are released from the surface of the liquid. This increases movement above the liquid surface, increasing the pressure in the vapor stage. The **Figure ??** illustrates the effect of temperature on vapor pressure.



Summary

- Water has high surface tension because of extensive hydrogen bonding.
- The vapor pressure of water is low due to hydrogen bonding.
- Vapor pressure increases as temperature increases.

Practice

Use the link below to answer the following questions:

http://science.howstuffworks.com/environmental/earth/geophysics/h2o7.htm

1. Why does water bead up on waxy surfaces?
2. Explain capillary action.
3. Define cohesion.
4. Define adhesion.

Review

1. What is surface tension?
2. What is vapor pressure?
3. How does temperature affect vapor pressure?

- **surface tension:** The attraction between molecules at the surface of a liquid.
- **vapor pressure:** The pressure of the vapor produced by evaporation of a liquid or solid above the liquid or solid in a closed container.

15.2 Properties of Solutions

- Describe how solutes affect the properties of solvents in solutions.
- Give examples of freezing point depression and boiling point elevation.

Why hasn't the ocean water in this photo turned to ice? The water in the glacier on shore is frozen solid, but the water in the ocean is still in a liquid state.

Q: What is it about ocean water that keeps it from freezing when the temperature falls below the freezing point of pure water?

A: Ocean water is salty.

How Solutes Affect Solvents

Salt water in the ocean is a solution. In a solution, one substance, called the solute, dissolves in another substance, called the solvent. In ocean water, salt is the solute and water is the solvent. When a solute dissolves in a solvent, it changes the physical properties of the solvent. In particular, the solute generally lowers the freezing point of the solvent, which is called freezing point depression, and raises the boiling point of the solvent, which is called boiling point elevation. For example, adding either salt to water lowers the freezing point and raises the boiling point of the water. To learn the why these effects occur, watch the excellent video at this URL: http://www.youtube.com/watch?v=z9LxdqYntlU .

Freezing Point Depression

Pure water freezes at 0 °C, but the salt water in the ocean freezes at -2.2 °C because of freezing point depression. We take advantage of the freezing point depression of salt in water by putting salt on ice to melt it. That's why the truck in the **Figure** 15.2 is spreading salt on an icy road.

FIGURE 15.2

Did you ever see anyone add a fluid to their car radiator? The fluid might be antifreeze, like in the **Figure** 15.3. Antifreeze lowers the temperature of the water in the car radiator so it won't freeze, even when the temperature falls far below 0 °C. For example, a 50 percent antifreeze solution won't freeze unless the temperature goes below -37 °C.

FIGURE 15.3

Boiling Point Elevation

Antifreeze could also be called "antiboil" because it also raises the boiling point of the water in a car radiator. Hot weather combined with a hot engine can easily raise the temperature of the water in the radiator above 100 °C, which is the boiling point of pure water. If the water boils, it could cause the engine to overheat and become seriously damaged. However, if antifreeze has been added to the water, the boiling point is much higher. For example a 50 percent antifreeze solution has a boiling point of 129 °C. Unless the water gets hotter than this, it won't boil and ruin the engine.

Summary

- When a solute dissolves in a solvent, it changes the physical properties of the solvent.
- A solute generally lowers the freezing point of a solvent, which is called freezing point depression. For example, spreading salt on an icy road melts the ice.
- A solute generally raises the boiling point of a solvent, which is called boiling point elevation. For example, adding antifreeze to the water in a car radiator prevents the water from boiling.

Practice

Do the animated experiment at the following URL. First, select water as the solvent and sodium chloride as the solute. Then, determine the boiling and freezing points of a solution containing different masses of the solute, while holding the mass of solvent constant at 200 grams. Test at least five different masses of the solute, and record your results in a data table. Finally, write a brief summary of what your data reveal about boiling point elevation and freezing point depression of saltwater solutions. http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/propOfSoln/colligative.html

Review

1. What is freezing point depression?
2. Give an example of boiling point elevation.
3. Assume you are going to boil water to cook spaghetti. If you add salt to the water, how will this affect the temperature at which the water boils? How might it affect the time it takes the spaghetti to cook?

15.3 Saturated and Unsaturated Solutions

- Define saturated solution.
- Define unsaturated solution.
- Define solution equilibrium.

How do you make sure a compound is pure?

When compounds are synthesized, they often have contaminating materials mixed in with them. The process of recrystallization can be used to remove these impurities. The crystals are dissolved in a hot solvent, forming a solution. When the solvent is cooled the compound is no longer as soluble and will precipitate out of solution, leaving other materials still dissolved.

Saturated and Unsaturated Solutions

Table salt (NaCl) readily dissolves in water. Suppose that you have a beaker of water to which you add some salt, stirring until it dissolves. So you add more and that dissolves. You keep adding more and more salt, eventually reaching a point that no more of the salt will dissolve no matter how long or how vigorously you stir it. Why? On the molecular level, we know that action of the water causes the individual ions to break apart from the salt crystal and enter the solution, where they remain hydrated by water molecules. What also happens is that some of the dissolved ions collide back again with the crystal and remain there. **Recrystallization** is the process of dissolved solute returning to the solid state. At some point the rate at which the solid salt is dissolving becomes equal to the rate at which the dissolved solute is recrystallizing. When that point is reached, the total amount of dissolved salt remains unchanged. **Solution equilibrium** is the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate. The solution equilibrium for the dissolving of sodium chloride can be represented by one of two equations.

$$\mathrm{NaCl}(s) \rightleftarrows \mathrm{NaCl}(aq)$$

While this shows the change of state back and forth between solid and aqueous solution, the preferred equation also shows the dissociation that occurs as an ionic solid dissolves.

$$\text{NaCl}(s) \rightleftarrows \text{Na}^+(aq) + \text{Cl}^-(aq)$$

When the solution equilibrium point is reached and no more solute will dissolve, the solution is said to be saturated. A **saturated solution** is a solution that contains the maximum amount of solute that is capable of being dissolved. At 20°C, the maximum amount of NaCl that will dissolve in 100. g of water is 36.0 g. If any more NaCl is added past that point, it will not dissolve because the solution is saturated. What if more water is added to the solution instead? Now more NaCl would be capable of dissolving in the additional solvent. An **unsaturated solution** is a solution that contains less than the maximum amount of solute that is capable of being dissolved. **Figure** 15.4 illustrates the above process and shows the distinction between unsaturated and saturated.

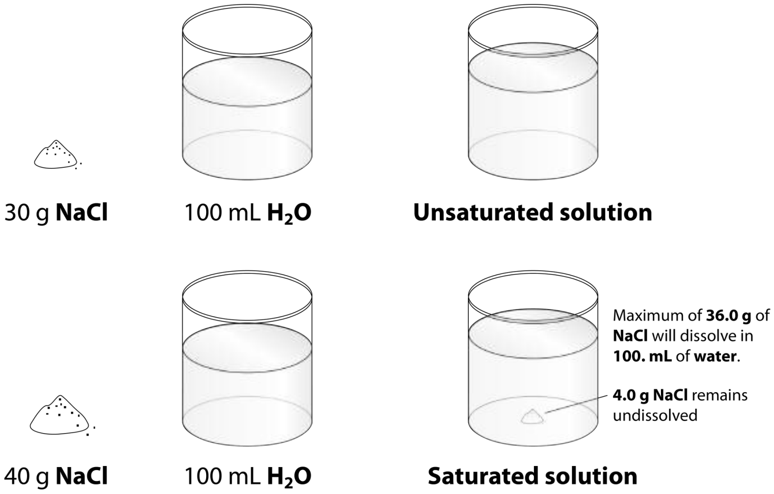


FIGURE 15.4

When 30.0 g of NaCl is added to 100 ml of water, it all dissolves, forming an unsaturated solution. When 40.0 g is added, 36.0 g dissolves and 4.0 g remains undissolved, forming a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium but which has extra undissolved solute at the bottom of the container must be saturated.

Summary

- Saturated and unsaturated solutions are defined.
- Solution equilibrium exists when the rate of dissolving equals the rate of recrystallization.

Practice

Watch the video at the link below and answer the following questions:

MEDIA

Click image to the left for more content.

http://www.youtube.com/watch?v=gawS3sBHMQw

1. What is the initial solution used?
2. What is the heat source for evaporation?
3. Why does the salt precipitate out of solution?

Review

1. Why is the preferred equation for solution equilibrium of NaCl an equilibrium between solid NaCl and the ions.
2. If I add water to a saturated sucrose solution, what will happen?
3. If I heat a solution and remove water, I see crystals at the bottom of the container. What happened?

- **recrystallization:** The process of dissolved solute returning to the solid state.
- **saturated solution:** A solution that contains the maximum amount of solute that is capable of being dissolved.
- **solution equilibrium:** The physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate.
- **unsaturated solution:** A solution that contains less than the maximum amount of solute that is capable of being dissolved.

15.4 How Temperature Influences Solubility

- Describe the influence of temperature on the solubility of solids in water.
- Describe the influence of temperature on the solubility of gases in water.

What happens to the fish in the water next to a nuclear power plant?

Nuclear power plants require large amounts of water to generate steam for the turbines and to cool the equipment. They will usually be situated near bodies of water to use that water as a coolant, returning the warmer water back to the lake or river. This increases the overall temperature of the water, which lowers the quantity of dissolved oxygen, affecting the survival of fish and other organisms.

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per 100 g of solvent. The solubility of sodium chloride in water is 36.0 g per 100 g water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. We will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (see **Figure** 15.5).

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep and so an increase in temperature dramatically increases the solubility of KNO_3.

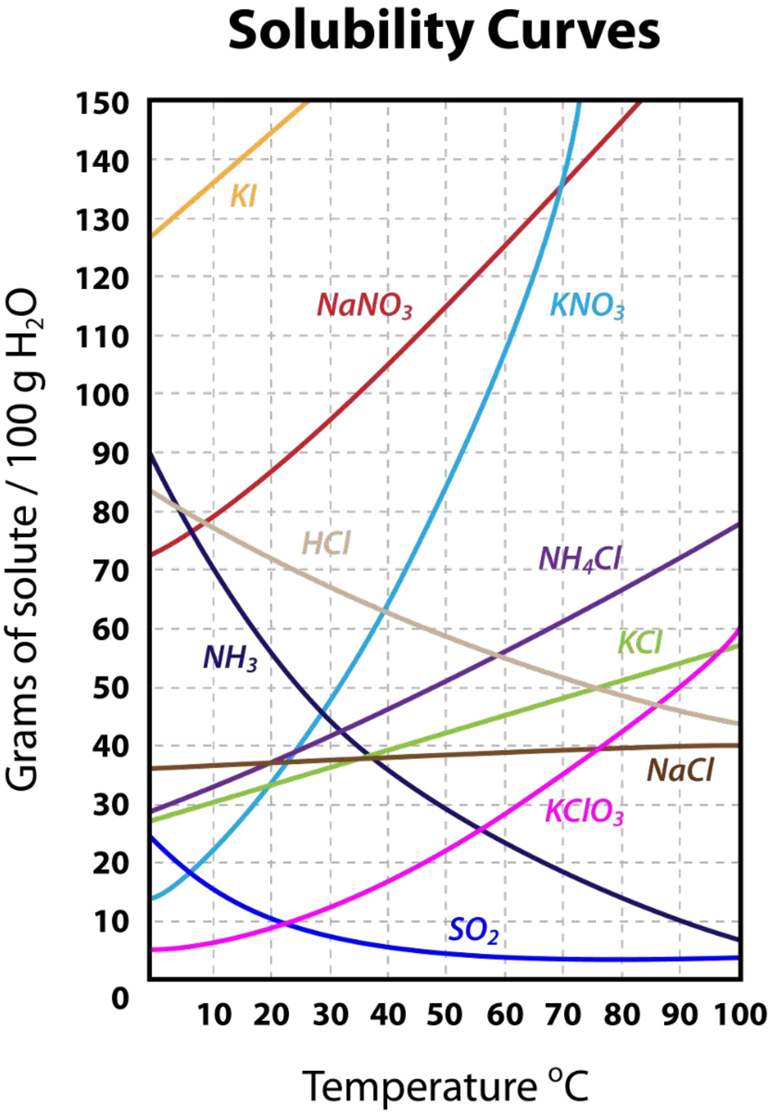


FIGURE 15.5

Solubility curves for several compounds.

Several substances –HCl, NH_3, and SO_2 –have solubility that decreases as temperature decreases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increases. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that 80 g of KNO_3 is added to 100 g of water at 30°C. According to the solubility curve, approximately 48 g of KNO_3 will dissolve at 30°C. This means that the solution will be saturated since 48 g is less than 80 g. We can also determine that there will be 80 - 48 = 32 g of undissolved KNO_3 remaining at the bottom of the container. Now suppose that this saturated solution is heated to 60°C. According to the curve, the solubility of KNO_3 at 60°C is about 107 g. Now the solution is unsaturated since it contains only the original 80 g of dissolved solute. Now suppose the solution is cooled all the way down to 0°C. The solubility at 0°C is about 14 g, meaning that 80 - 14 = 66 g of the KNO_3 will recrystallize.

Summary

- The solubility of a solid in water increases with an increase in temperature.

- Gas solubility decreases as the temperature increases.

Practice

Read the material at the link below and answer the questions. Try not to look at the answers until you do your own work:

http://chemwiki.ucdavis.edu/Physical_Chemistry/Physical_Properties_of_Matter/Solutions/SOLUBILITY/Temperature_Effects_On_The_Solubility_Of_Gases

Review

1. Why does the solubility of a gas decrease as the temperature increases?
2. Is the solubility of NaCl affected by solvent?
3. What is the solubility of KNO_3 at 50°C?

- **solubility:** The amount of a substance that is required to form a saturated solution in a given amount of solvent at a specified temperature.
- **solubility curve:** A graph of the solubility of substances as a function of temperature.

15.5 Molarity

- Define molarity.
- Perform calculations involving molarity.

How many molecules can be found in a reaction?

Chemists deal with amounts of molecules every day. Our reactions are described as so many molecules of compound A reacting with so many molecules of compound B to form so many molecules of compound C. When we determine how much reagent to use, we need to know the number of molecules in a given volume of the reagent. Percent solutions only tell us the number of grams, not molecules. A 100 mL solution of 2% NaCl will have a very different number of molecules than a 2% solution of CsCl. So we need another way to talk about numbers of molecules.

Molarity

Chemists primarily need the concentration of solutions to be expressed in a way that accounts for the number of particles that react according to a particular chemical equation. Since percentage measurements are based on either mass or volume, they are generally not useful for chemical reactions. A concentration unit based on moles is preferable. The **molarity (M)** of a solution is the number of moles of solute dissolved in one liter of solution. To calculate the molarity of a solution, you divide the moles of solute by the volume of the solution expressed in liters.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}}$$

Note that the volume is in liters of solution and not liters of solvent. When a molarity is reported, the unit is the symbol M and is read as “molar”. For example a solution labeled as 1.5 M NH_3 is read as “1.5 molar ammonia solution”.

Sample Problem: Calculating Molarity

A solution is prepared by dissolving 42.23 g of NH_4Cl into enough water to make 500.0 mL of solution. Calculate its molarity.

Step 1: List the known quantities and plan the problem.

Known	Unknown
mass $= 42.23$ g NH_4Cl	molarity $=$? M
molar mass $NH_4Cl = 53.50$ g/mol	
volume solution $= 500.0$ mL $= 0.5000$ L	

The mass of the ammonium chloride is first converted to moles. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters.

Step 2: Solve.

$$42.23 \text{ g } NH_4Cl \times \frac{1 \text{ mol } NH_4Cl}{53.50 \text{ g } NH_4Cl} = 0.7893 \text{ mol } NH_4Cl$$

$$\frac{0.7893 \text{ mol } NH_4Cl}{0.5000 \text{ L}} = 1.579 \text{ M}$$

Step 3: Think about your result.

The molarity is 1.579 M, meaning that a liter of the solution would contain 1.579 mol NH_4Cl. Four significant figures are appropriate.

In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams. See sample problem 16.3.

Sample Problem:

A chemist needs to prepare 3.00 L of a 0.250 M solution of potassium permanganate ($KMnO_4$). What mass of $KMnO_4$ does she need to make the solution?

Step 1: List the known quantities and plan the problem.

Known

- molarity = 0.250 M
- volume = 3.00 L
- molar mass $KMnO_4$ = 158.04 g/mol

Unknown

- mass $KMnO_4$ = ? g

Moles of solute is calculated by multiplying molarity by liters. Then, moles is converted to grams.

Step 2: Solve.

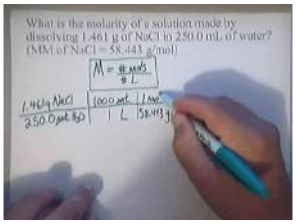
$$\text{mol KMnO}_4 = 0.250 \text{ M KMnO}_4 \times 3.00 \text{ L} = 0.750 \text{ mol KMnO}_4$$

$$0.750 \text{ mol KMnO}_4 \times \frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4} = 119 \text{ g KMnO}_4$$

Step 3: Think about your result.

When 119 g of potassium permanganate is dissolved into water to make 3.00 L of solution, the molarity is 0.250 M.

Watch a video of molarity calculations:



MEDIA

Click image to the left for more content.

http://www.youtube.com/watch?v=8oTqwBAvbnY

Summary

- Calculations using the concept of molarity are described.

Practice

Read the material and work the problems at the site below:

http://www.occc.edu/kmbailey/Chem1115Tutorials/Molarity.htm

Review

1. What does M stand for?
2. What does molarity tell us that percent solution information does not tell us?
3. What do we need to know about a molecule in order to carry out molarity calculations?

- **molarity (M):** The number of moles of solute dissolved in one liter of solution.

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CHAPTER 16

Acids and Bases

Chapter Outline

Chm.3.2.1

- Distinguish between acids and bases based on formula and chemical properties.
- Differentiate between concentration (molarity) and strength (degree of dissociation). No calculation involved.
- Use pH scale to identify acids and bases.
- Interpret pH scale in terms of the exponential nature of pH values in terms of concentrations.
- Relate the color of indicator to pH using pH ranges provided in a table.
- Compute pH, pOH, [H+], and [OH-].

Chm.3.2.2

Distinguish properties of acids and bases related to taste, touch, reaction with metals, electrical conductivity, and identification with indicators such as litmus paper and phenolphthalein.

Chm.3.2.3

- Perform 1:1 titration calculations: $M_1V_1 = M_2V_2$
- Determine the concentration of an acid or base using titration. Interpret titration curve for strong acid/strong base.

16.1 Acids and Bases

Lesson Objectives

- Define acids and bases in terms of the ions that are produced when each type of compound is dissolved in water.
- Be able to name acids when given their chemical formulas.
- Write the correct formula for an acid when given its name.
- Name and write formulas for bases.

Lesson Vocabulary

- acid
- base
- binary acid
- oxoacid

Check Your Understanding

Recalling Prior Knowledge

- What are cations and anions, and how are they formed?
- How do ionic compounds and molecular compounds compare structurally?
- How do you make sure that the formula for an ionic compound is electrically neutral?

Acids and bases are two very important classes of chemicals that you probably use every day. Many foods and beverages contain acids, such as citrus fruits and juices, soda, and vinegar (**Figure** 16.1). Bases are very prevalent in cleaners such as ammonia, and most drain cleaners use a strong base as their active ingredient. Acid-base chemistry will be covered in great detail in a later chapter. In this lesson, we will focus on the nomenclature used for these two types of compounds.

Acids

An acid can be defined in several ways. For the purposes of this chapter, the most straightforward description is that an **acid** *is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions* (H^+) *when dissolved in water*.

This is a different type of compound than the others we have seen in this chapter. Acids are molecular, which means that in their pure state, they exist as individual molecules and do not adopt the extended three-dimensional structures

FIGURE 16.1

(A) Vinegar comes in a variety of types, but they all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

exhibited by ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. For example, the following chemical equation describes what happens when the acid HCl is dissolved in water:

$$\mathrm{HCl} \rightarrow \mathrm{H}^+ + \mathrm{Cl}^-$$

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (everything except the H) constitutes the anion that is formed after the acid dissolves. Organic acids are also an important class of compounds, but the rules for naming them are different, and they will primarily be discussed in a later chapter. A **binary acid** *is an acid that consists of hydrogen and one other element*. The most common binary acids contain a halogen. An **oxoacid** *is an acid that consists of hydrogen, oxygen, and a third element*. The third element is usually a nonmetal.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. Recall from earlier in the chapter that anions can either be monatomic or polyatomic. The names of all monatomic ions end in *–ide*. The majority of polyatomic ions end in either *–ate* or *–ite*, though there are a few exceptions, such as the cyanide ion (CN^-). It is the suffix of the anion that determines how the acid is named, as described in the rules and the table below (**Table** 16.1).

The three different suffixes that are possible for the anions lead to the three rules below.

1. When the anion ends in *–ide*, the acid name begins with the prefix *hydro-*. The root of the anion name goes in the blank (e.g., *chlor-* for chloride), followed by the suffix *–ic acid*. HCl is named hydrochloric acid, because Cl^- is a chloride ion. HCN is hydrocyanic acid because CN^- is a cyanide ion.
2. When the anion ends in *–ate*, the name of the acid is the root of the anion followed by the suffix *–ic acid*. There is no prefix. H_2SO_4 is sulfuric acid (not sulfic) because ${SO_4}^{2-}$ is the sulfate ion.
3. When the anion ends in *–ite*, the name of the acid is the root of the anion followed by the suffix *–ous acid*. Again, there is no prefix. HNO_2 is nitrous acid because ${NO_2}^-$ is the nitrite ion.

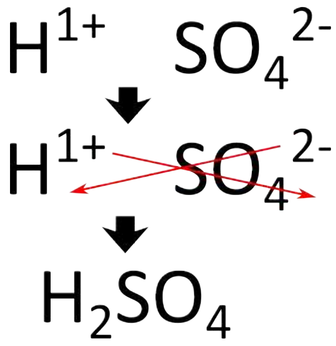
Note that the root for a sulfur-containing oxoacid is *sulfur-* instead of just *sulf-*. The same is true for a phosphorus-containing oxoacid, which uses *phosphor-* as its root instead of simply *phosph-*.

TABLE 16.1: Naming System for Acids

Anion Suffix	Example	Name of acid	Example
-ide	chloride (Cl^-)	hydro_____ic acid	hydrochloric acid (HCl)
-ate	sulfate ($SO_4{}^{2-}$)	_____ic acid	sulfuric acid H_2SO_4)
-ite	nitrite ($NO_2{}^-$)	_____ous acid	nitrous acid (HNO_2)

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H+ ions. Since H^+ ions carry a single positive charge, the number of H^+ ions in the formula is equal to the magnitude of the negative charge on the anion. Two examples from the table above (**Table** 16.1) illustrate this point. The chloride ion carries a 1− charge, so only one H is needed in the formula of the acid (HCl). The sulfate ion carries a 2− charge, so two hydrogen atoms are needed in the formula of the acid (H_2SO_4). Another way to think about writing the correct formula is to utilize the crisscross method, as shown below for sulfuric acid.



Bases

The simplest way to define a **base** *is an ionic compound that produces hydroxide ions when dissolved in water*. One of the most commonly used bases is sodium hydroxide (**Figure** 16.2).

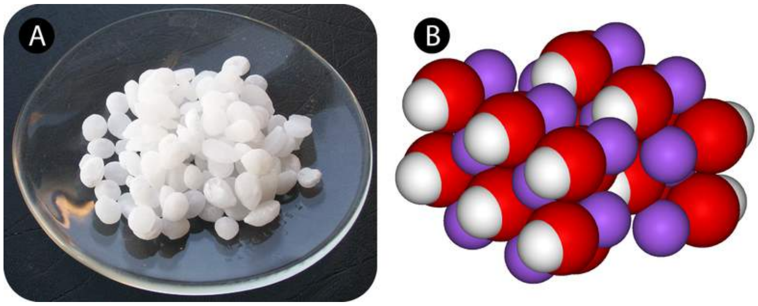


FIGURE 16.2

(A) Sodium hydroxide, a base, is a solid that is typically produced as small white pellets. (B) The structure of sodium hydroxide is an extended three-dimensional network. The purple spheres are the sodium ions (Na^+). The red and white spheres are oxygen and hydrogen atoms, respectively, which are bonded together to form hydroxide ions (OH^-).

Names and Formulas of Bases

Inorganic bases are named in the same way as other ionic compounds. Since they all contain the OH^- anion, the names of these bases end in hydroxide. Some examples of names and formulas for bases are shown below (**Figure** 16.2).

TABLE 16.2: Examples of Bases

Formula	Name
NaOH	sodium hydroxide
$Ca(OH)_2$	calcium hydroxide
NH_4OH	ammonium hydroxide

Notice that because bases are ionic compounds, the number of hydroxide ions in the formula does not affect the name. The compound must be neutral, so the charges of the ions are balanced just as for other ionic compounds. The sodium ion (Na^+) requires one OH^- ion to balance the charge, so the formula of sodium hydroxide is NaOH. The calcium ion (Ca^{2+}) requires two OH^- ions to balance the charge, so the formula of calcium hydroxide is $Ca(OH)_2$. Remember that the hydroxide ion is a polyatomic ion, so it must be put in parentheses when the formula contains more than one.

Lesson Summary

- Acids are molecular compounds that dissolve in water to produce hydrogen ions and an anion. Bases are ionic compounds consisting of hydroxide ions and a cation.
- The naming rules for acids are based on the suffix of the anion. Formulas for acids are written by balancing out the charge of the anion with the appropriate number of hydrogen ions.
- Naming and formula writing for bases follows the same guidelines as for other ionic compounds.

Lesson Review Questions

Reviewing Concepts

1. What ion must be produced when an acid dissolves in water?
2. What ion must be present for a compound to be considered a base?
3. Identify each of the following compounds as either an acid, a base, or neither.
    a. RbOH
    b. HBr
    c. MgO
    d. CH_4
    e. H_3PO_3
    f. $Sr(OH)_2$

Problems

4. Name the following acids.
    a. HF

   b. $HClO_2$
   c. H_2CrO_4
   d. H_2SO_3
   e. H_3PO_4
   f. $HClO_4$
   g. H_2S
   h. HNO_3

5. Write correct formulas for the following acids.

   a. carbonic acid
   b. hydroiodic acid
   c. chloric acid
   d. phosphorous acid
   e. oxalic acid
   f. hypochlorous acid
   g. hydrobromic acid
   h. permanganic acid

6. Write names or formulas for these compounds.

   a. LiOH
   b. $Mg(OH)_2$
   c. $Fe(OH)_3$
   d. nickel(II) hydroxide
   e. aluminum hydroxide
   f. silver hydroxide

7. Identify each compound below as either a (1) ionic compound, (2) molecular compound, (3) acid, or (4) base. Then, name it properly.

   a. HCN
   b. KNO_3
   c. N_2O_4
   d. $Sr(OH)_2$
   e. $HBrO_3$
   f. SO_3
   g. HCH_3COO
   h. CsOH
   i. $Pb_3(PO_4)_2$
   j. AsH_3

Further Reading / Supplemental Links

- Kristi Lew, *Acids and Bases* (*Essential Chemistry*). Chelsea House, 2008.
- Chris Oxlade, *Acids and Bases*. Heinemann-Raintree, 2007.
- An Introduction to Chemistry (http://preparatorychemistry.com/Bishop_acid_nomenclature_help.htm

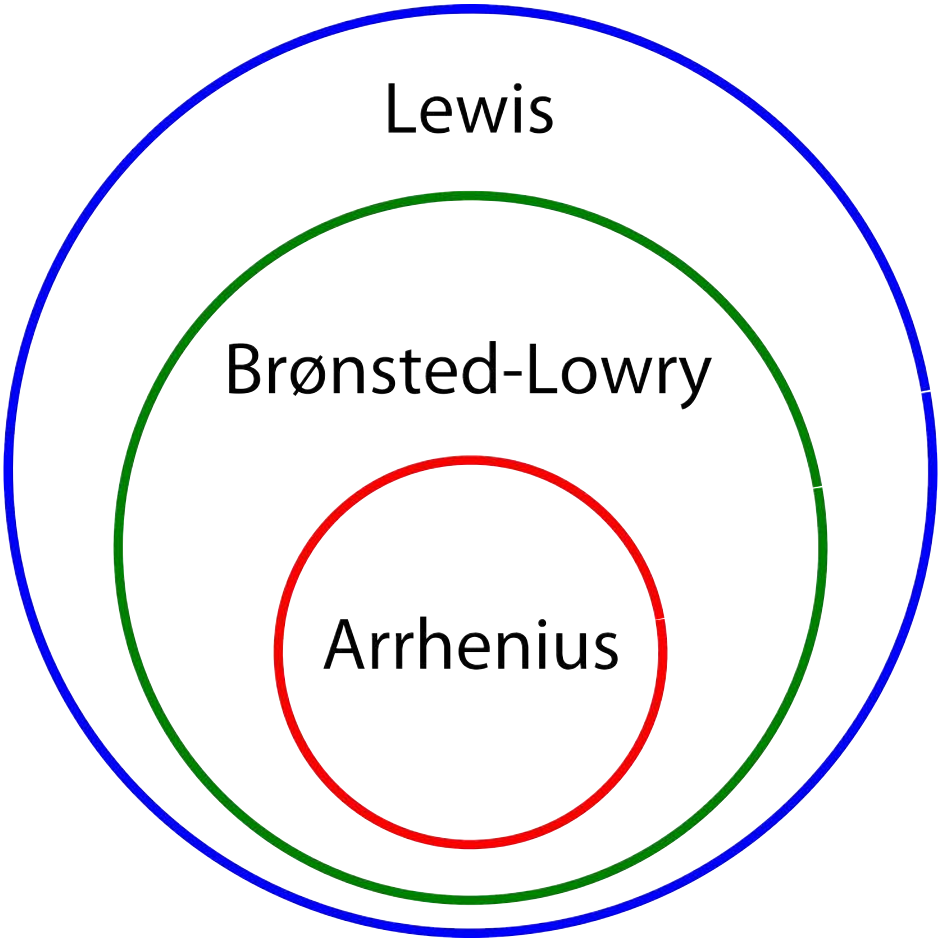
Points to Consider

The extended three-dimensional structure of ionic compounds is commonly called a crystal lattice. The interactions of the ions within a crystal lattice lead directly to many important physical and chemical properties of ionic compounds.

- How is a crystal lattice formed?
- Why is a crystal lattice the most stable structural form for an ionic compound?
- What physical properties do ionic compounds have in common as a result of their structure?

16.2 Lewis Acids and Bases

- Define Lewis acids and bases.
- Give examples of Lewis acids and bases.



The big picture gets bigger

Ideas in science never stay static. One discovery builds on another. Our concepts of acids and bases have grown from the fundamental ideas of Arrhenius to Brønsted-Lowry to Lewis. Each step adds to our understanding of the world around us and makes the "big picture" even bigger.

Lewis Acids and Bases

Gilbert Lewis (1875-1946) proposed a third theory of acids and bases that is even more general than either the Arrhenius or Brønsted-Lowry theories. A **Lewis acid** is a substance that accepts a pair of electrons to form a covalent bond. A **Lewis base** is a substance that donates a pair of electrons to form a covalent bond. So, a Lewis acid-base reaction is represented by the transfer of a pair of electrons from a base to an acid. A hydrogen ion, which lacks any electrons, accepts a pair of electrons. It is an acid under both the Brønsted-Lowry and Lewis definitions. Ammonia consists of a nitrogen atom as the central atom with a lone pair of electrons. The reaction between ammonia and the hydrogen ion can be depicted as shown in the **Figure** 16.3.

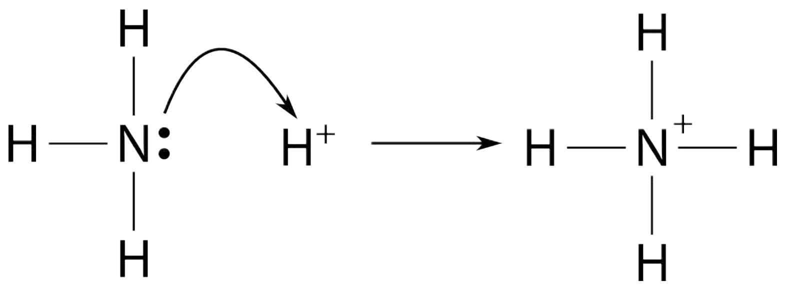


FIGURE 16.3
Reaction between ammonia and proton.

The lone pair on the nitrogen atom is transferred to the hydrogen ion, making the NH_3 a Lewis base while the H^+ is a Lewis acid.

Some reactions that do not qualify as acid-base reactions under the other definitions do so under only the Lewis definition. An example is the reaction of ammonia with boron trifluoride.

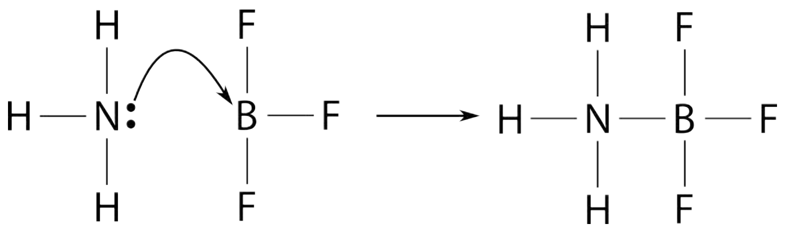


FIGURE 16.4
Ammonia and boron trifluoride.

Boron trifluoride is the Lewis acid, while ammonia is again the Lewis base. As there is no hydrogen ion involved in this reaction, it qualifies as an acid-base reaction only under the Lewis definition. The **Table** 16.3 summarizes the three acid-base theories.

TABLE 16.3: **Acid-Base Definitions**

Type	**Acid**	**Base**
Arrhenius	H^+ ions in solution	OH^- ions in solution
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

Summary

- Lewis acids and bases are defined.
- Examples of Lewis acids and bases are given.

Practice

Do the practice problems associated with Section 3 Lewis Theory at the link below:

http://staarhelper.com/Mr_B_Science/daily_lessons/cacid.htm

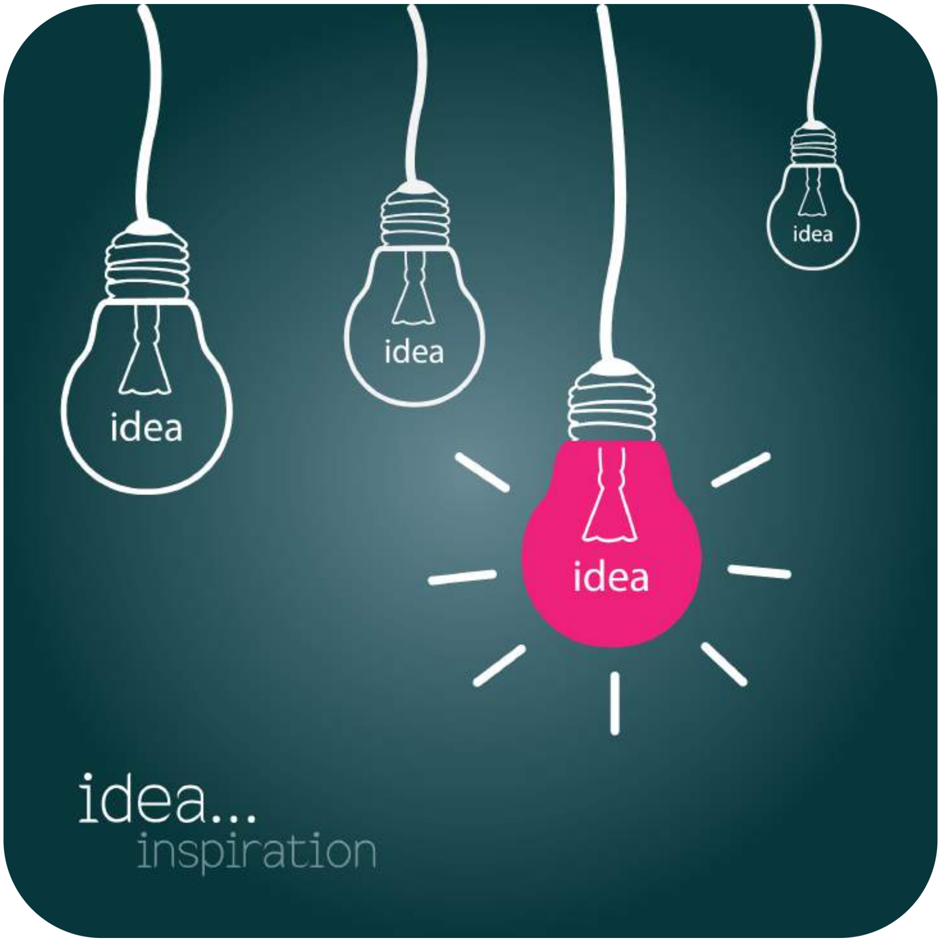
Review

1. What is the difference between a Lewis acid or base and a Brønsted Lowry acid or base?
2. What is required to happen in Lewis acid-base reactions?
3. Do all Lewis acid-base reactions involve protons?

- **Lewis acid:** A substance that accepts a pair of electrons to form a covalent bond.

16.3 Brønsted-Lowry Acids and Bases

- Define Brønsted-Lowry acid.
- Define Brønsted-Lowry base.
- Give examples of Brønsted-Lowry acids and bases.



A new theory

The Arrhenius concept of acids and bases was a significant contribution to our understanding of acids and bases. It replaced and expanded the original idea of Lavoisier that all acids contained oxygen. However, the Arrhenius theory had its shortcomings also. It did not take into account the role of the solvent. In addition, this concept did not deal with acid-base behavior in solvents such as benzene where there could be no ionization. So, a new theory needed to be formed, which built on the findings of Arrhenius but also went beyond them.

Brønsted-Lowry Acids and Bases

The Arrhenius definition of acids and bases is somewhat limited. There are some compounds whose properties suggest that they are either acidic or basic, but which do not qualify according to the Arrhenius definition. An example is ammonia (NH_3). Its aqueous solution turns litmus blue, it reacts with acids, and displays all the other properties of a base. However, it does not contain the hydroxide ion. In 1923, a broader definition of acids and bases was independently proposed by Danish chemist Johannes Brønsted (1879-1947) and English chemist Thomas Lowry (1874-1936). A **Brønsted-Lowry acid** is a molecule or ion that donates a hydrogen ion in a reaction. A **Brønsted-Lowry base** is a molecule or ion that accepts a hydrogen ion in a reaction. A hydrogen ion is commonly referred to as a proton, and so acids and bases are proton donors and proton acceptors respectively according to the Brønsted-Lowry definition. All substances that are categorized as acids and bases under the Arrhenius definition are also defined as such under the Brønsted-Lowry definition. The new definition, however, includes some substances that are left out according to the Arrhenius definition.

What kind of molecule would qualify as a Brønsted-Lowry base? These molecules need to be able to accept a hydrogen ion (or proton). Two possibilities come to mind: an anion that can form a neural compound with a proton, or a molecule in which one or more atoms has lone-pair electrons. The most obvious anion is the Arrhenius base OH^-. This ion can form a water molecule with a proton by accepting the proton. The acetate anion CH_3COO^- is another anion that can combine with a proton to form acetic acid. Lone-pair electron groups would include the nitrogen atom (see **Figure** 16.5). The two electrons at the top of the nitrogen atom are not connected in any type of bond, but they do interact readily with a bare proton.

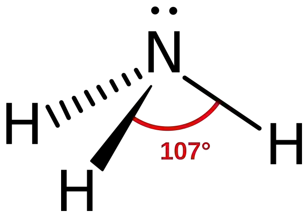


FIGURE 16.5

Lone pair electrons on nitrogen.

Oxygen is another atom with lone pair electrons that can function as Brønsted-Lowry bases.

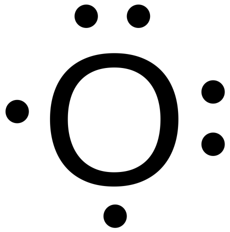


FIGURE 16.6

Electrons in an oxygen atom.

The two single electrons (to the left and bottom of the atom) can form single covalent bonds with other atoms while the two pairs of double electrons (top and right) are available to interact with a hydrogen ion.

Summary

- Brønsted-Lowry acids and bases are defined.

Practice

Read the material at the link below and answer the following questions:

http://www.chemteam.info/AcidBase/Bronsted-Lowry-AcidBase.html

1. Why is water considered a base in the reaction between water and HCl?
2. Why is H_3O^+ considered an acid?
3. Why is the chloride anion considered a base?

Review

1. What is a Brønsted-Lowry acid?
2. What is a Brønsted-Lowry base?
3. How does ammonia function as a Brønsted-Lowry base?

- **Brønsted-Lowry acid:** A molecule or ion that donates a hydrogen ion in a reaction.
- **Brønsted-Lowry base:** A molecule or ion that accepts a hydrogen ion in a reaction.

16.4 Acids and Bases Review

Lesson Objectives

- Define and give examples for the following terms: acid, base, binary acid, oxoacid.
- Determine the name of an inorganic acid or base when given the formula.
- Determine the formula of an inorganic acid or base when given its name.

Lesson Vocabulary

- **acid**: Any compound that produces hydrogen ions (H+) when dissolved in water.
- **binary acid**: Acids in which one or more acidic hydrogen atoms are bound directly to a single atom.
- **oxoacid**: A strong acid produced by combining oxoanions with one or more hydrogen ions.
- **base**: A compound that produces the hydroxide (OH-) ion when dissolved in water.

Check Your Understanding

1. Name the following compounds:
    a. $NaNO_3$
    b. BF_3
    c. $FeSO_3$
2. Are the following compounds molecular or ionic?
    a. H_2O
    b. CH_4
    c. $BaSO_4$

Introduction

In this chapter, we first looked at the naming conventions for ionic compounds, which exist as extended networks of cations and anions. For most of the compounds that we considered, the cation was a monatomic metal (e.g., Na^+, Mg^{2+}, Fe^{3+}), and the anion was a monatomic nonmetal (e.g., Cl^-, O^{2-}, N^{3-}) or a polyatomic ion, which often contains multiple oxygen atoms (e.g., ${SO_4}^{2-}$). We then looked at molecular compounds, in which atoms are held together into individual molecules by covalent bonds. Now we are going to consider acids and bases, which share characteristics with both ionic and molecular compounds.

Acids

There are multiple ways to define what an acid is, but for the purposes of this book, we will define an **acid** as any compound that produces hydrogen ions (H^+) when dissolved in water. Based on this definition, all acids contain at least one hydrogen atom, but not all hydrogen-containing compounds are acids. When isolated as a pure material, most acids exist as molecular substances. However, when dissolved in water, one or more of the hydrogen atoms acts as an H^+ ion that transfers to water, leaving behind the remainder of the molecule as an anion. This reaction with water can be represented by the following generic equation, where HA represents an acid:

$HA_{(aq)} + H_2O_{(l)} \rightarrow A^-_{(aq)} + H_3O^+_{(aq)}$

As you can see, the acid reacts with a molecule of water to produce a hydronium ion (H_3O^+) and the A^- anion. (Note: The H^+ ion is sometimes referred to as a proton. This makes sense when you consider that the most common form of the hydrogen atom consists of one proton and one electron. When the single electron is removed to make a cation, only a proton is left behind. As a result, the above reaction is sometimes referred to as a proton transfer.) A specific example of this process can be seen in the following animation:

http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/animations/HCl(aq).html

In this animation, hydrochloric acid (HCl) reacts with water to produce the hydronium ion and the chloride ion. Even though HCl exists as a molecular gas in the absence of water, it produces ions when water is present. Acids have some unique properties and reactivity patterns that we will discuss in future chapters. For now, we will focus on the ways in which they are named.

Binary Acids

Binary acids are acids in which one or more acidic hydrogen atoms are bound directly to a single atom. As a result, the anion left behind when a binary acid is dissolved in water is a monatomic anion. Examples include hydrogen chloride ($HCl_{(g)}$ and hydrogen sulfide ($H_2S_{(g)}$). Both of these substances are molecular gases in their pure form, but change their properties and their names when dissolved in water. To name a binary acid, start with the name of the anion left behind after the acidic hydrogens have been removed. Then, add the prefix *hydro-* and replace the suffix *-ide* with *-ic acid*. For example, HCl produces Cl^- ions when dissolved in water, so it would therefore be named hydrochloric acid. Some other common binary acids are shown in the following **Table** 16.4:

TABLE 16.4: **Common Binary Acids**

Formula	Name	Anion	Name
$HF_{(aq)}$	hydrofluoric acid	F^-	fluoride
$HCl_{(aq)}$	hydrochloric acid	Cl^-	chloride
$HBr_{(aq)}$	hydrobromic acid	Br^-	bromide
$HI_{(aq)}$	hydroiodic acid	I^-	iodide
$H_2S_{(aq)}$	hydrosulfuric acid	S^{2-}	sulfide

Most of the binary acids listed here are monoprotic, because they have only one acidic hydrogen. Hydrosulfuric acid, on the other hand is diprotic. Its hydrogen ions are transferred to two water molecules in two subsequent reactions.

Oxoacids

So far we have looked at acids that leave behind monatomic anions. However, many strong acids leave behind polyatomic anions as well. In particular, many of the oxoanions we looked at earlier can combine with one or more hydrogen ions (enough to make a neutral molecule) to produce strong acids called **oxoacids**. A common example of an oxoacid is nitric acid (HNO_3), which can be thought of as a hydrogen ion (H^+) combining with a nitrate ion ($NO_3{}^-$). If we simply named this as an ionic compound, we could name the compound hydrogen nitrate. However, because nitric acid exists as individual molecules and not an ionic structure, we use a different convention. To name an oxoacid, look at the anion that will be left behind when all acidic hydrogens have been removed. If it ends in *-ate*, replace that suffix with *-ic acid*. If it ends in *-ite*, replace that suffix with *-ous acid*. The following **Table** 16.5 lists some common oxoacids and their formulas:

TABLE 16.5: Common Oxoacids

Formula	Name	Anion	Name
$HClO_4$	*per*chloric acid	${ClO_4}^-$	*per*chlor*ate*
$HClO_3$	chlor*ic* acid	${ClO_3}^-$	chlor*ate*
$HClO_2$	chlor*ous* acid	${ClO_2}^-$	chlor*ite*
$HClO$	*hypo*chlor*ous* acid	ClO^-	*hypo*chlor*ite*
HNO_3	nitr*ic* acid	${NO_3}^-$	nitr*ate*
HNO_2	nitr*ous* acid	${NO_2}^-$	nitr*ite*
H_2SO_4	sulfur*ic* acid	${SO_4}^{2-}$	sulfate
H_2SO_3	sulfur*ous* acid	${SO_3}^{2-}$	sulf*ite*
H_2CO_3	carbon*ic* acid	${CO_3}^{2-}$	carbon*ate*
H_3PO_4	phosphor*ic* acid	${PO_4}^{3-}$	phosph*ate*

Example 7.3

Name the following compounds:

1. HIO_3
2. $NaBrO_2$
3. $Ca_3(PO_4)_2$
4. H_3PO_3

Answer:

1. iodic acid
2. sodium bromite
3. calcium phosphate
4. phosphorous acid

Bases

Bases can also be defined in multiple ways, but for now, we will define a **base** as a compound that produces the hydroxide (OH^-) ion when dissolved in water. Most of the common strong bases that you will need to deal with are simply ionic compounds in which a metal cation is combined with the hydroxide anion. These bases are named in the same way as any other ionic compound. For example, NaOH would be named sodium hydroxide, and $Ca(OH)_2$ is calcium hydroxide.

Some common bases are listed in the **Table** 16.6:

TABLE 16.6: Examples of Bases

Formula	Name
$NaOH$	sodium hydroxide
$Ca(OH)_2$	calcium hydroxide
NH_4OH	ammonium hydroxide

Lesson Summary

- Acids are molecular compounds that dissolve in water to produce hydronium ions and an anion.

$HA_{(aq)} + H_2O_{(l)} \rightarrow A^-_{(aq)} + H_3O^+_{(aq)}$

- The naming rules for acids are based on the suffix of the anion. Formulas for acids are written by balancing out the charge of the anion with the appropriate number of hydrogen ions.
- Bases are ionic compounds consisting of hydroxide ions and a cation. Naming and formula writing for bases follows the same guidelines as for other ionic compounds.

Review Questions

1. How do acids behave in water?
2. What defines a binary acid?
3. What defines an oxoacid?
4. How do bases behave in water?

5. Complete the following **Table** 16.7:

TABLE 16.7: Review Question 1

#	Acid Name	Formula of Acid	Name of Anion
1.	hydrobromic acid	HBr	bromide
2.	carbonic acid	H_2CO_3	carbonate
3.		HCl	
4.			sulfite
5.	chlorous acid		
6.	nitric acid		
7.			sulfide
8.		HNO_2	
9.	chromic acid		
10.			phosphate

6. Name the following acids:
   (a) HF
   (b) HI
   (c) H_2S
   (d) H_3PO_4
   (e) H_2SO_4
7. Write the formulas for the following acids:
   (a) sulfurous acid
   (b) hydrosulfuric acid
   (c) nitric acid
   (d) carbonic acid
   (e) chloric acid

Further Reading / Supplemental Links

- Chemical Nomenclature. (2012), from http://www.chemteam.info/Nomenclature/Nomenclature.html
- Video on acid-base nomenclature: http://www.youtube.com/watch?v=CVgi74kswPA

Points to Consider

- Vinegar is an acid that can be produced from the aerobic fermentation of wine. In fact, vinegar is most likely the oldest known acid. It is commonly used as a food additive (to give things an acidic or sour taste) and as a mild cleaning agent.

16.5 Introduction to the pH Concept

- Explain why some acids and bases are stronger than others.
- Define pH.
- Show how to use the pH scale.
- Explain why pH is important to living things.

This scientist is collecting and testing samples of water from the river. One of the properties of the water she is testing is acidity. She wants to know how acidic the water is because water that is too acidic can harm the health of water organisms.

Strength of Acids and Bases

Acids are ionic compounds that produce positively charged hydrogen ions (H^+) when dissolved in water. Acids taste sour and react with metals. Bases are ionic compounds that produce negatively charged hydroxide ions (OH^-) when dissolved in water. Bases taste bitter and do not react with metals. Examples of acids are vinegar and battery acid. The acid in vinegar is weak enough to safely eat on a salad. The acid in a car battery is strong enough to eat through skin. Examples of bases include those in antacid tablets and drain cleaner. Bases in antacid tablets are weak enough to take for an upset stomach. Bases in drain cleaner are strong enough to cause serious burns.

Q: What do you think causes these differences in the strength of acids and bases?

A: The strength of an acid or a base depends on how much of it breaks down into ions when it dissolves in water.

Concentration of Ions

The strength of an acid depends on how many hydrogen ions it produces when it dissolves in water. A stronger acid produces more hydrogen ions than a weaker acid. For example, sulfuric acid (H_2SO_4), which is found in car batteries, is a strong acid because nearly all of it breaks down into ions when it dissolves in water. On the other hand, acetic acid (CH_3CO_2H), which is the acid in vinegar, is a weak acid because less than 1 percent of it breaks down into ions in water.

The strength of a base depends on how many hydroxide ions it produces when it dissolves in water. A stronger base produces more hydroxide ions than a weaker base. For example, sodium hydroxide (NaOH), a base in drain cleaner, is a strong base because all of it breaks down into ions when it dissolves in water. Calcium carbonate ($CaCO_3$), a base in antacids, is a weak base because only a small percentage of it breaks down into ions in water.

The pH Scale

The strength of acids and bases is measured on a scale called the pH scale, which is shown in the **Figure** 16.7. By definition, **pH** represents the **acidity**, or hydrogen ion (H^+) concentration, of a solution. Pure water, which is neutral, has a pH of 7. With a higher the concentration of hydrogen ions, a solution is more acidic and has a lower pH. Acids have a pH less than 7, and the strongest acids have a pH close to zero. Bases have a pH greater than 7, and the strongest bases have a pH close to 14. It's important to realize that the pH scale is based on powers of ten. For example, a solution with a pH of 8 is 10 times more basic than a solution with a pH of 7, and a solution with a pH of 9 is 100 times more basic than a solution with a pH of 7.

Q: How much more acidic is a solution with a pH of 4 than a solution with a pH of 7?

A: A solution with a pH of 4 is 1000 ($10 \times 10 \times 10$, or 10^3) times more acidic than a solution with a pH of 7.

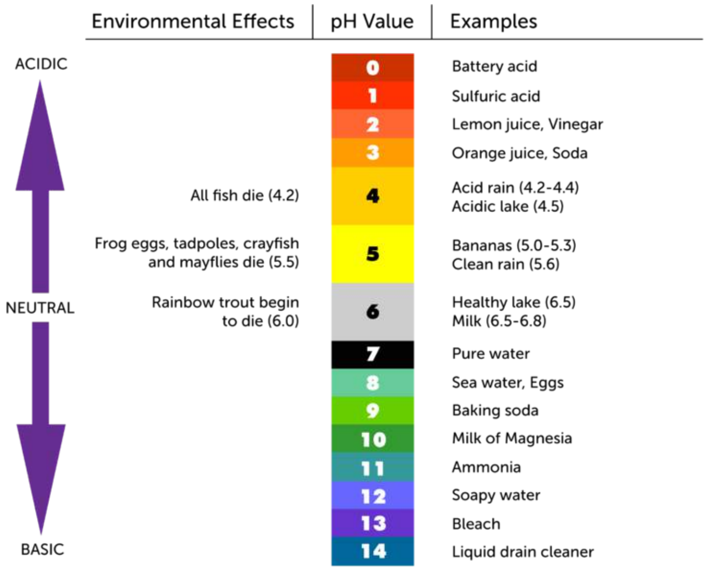


FIGURE 16.7

Q: Which solution on the pH scale in the **Figure** 16.7 is the weakest acid? Which solution is the strongest base?

A: The weakest acid on the scale is milk, which has a pH value between 6.5 and 6.8. The strongest base on the scale is liquid drain cleaner, which has a pH of 14.

For a deeper understanding of pH and the pH scale, watch the video at this URL: http://www.youtube.com/watch?v=M8tTELZD5Ek .

For an interactive exploration of pH, including an interactive pH scale, go to this URL: http://www.miamisci.org/ph/index.html .

Why pH Matters

Acidity is an important factor for living things. For example, many plants grow best in soil that has a pH between 6 and 7. Fish may also need a pH between 6 and 7. Certain air pollutants form acids when dissolved in water droplets in the air. This results in acid fog and acid rain, which may have a pH of 4 or even lower. The pH chart in the **Figure** 16.7 and the **Figure** 16.8 reveal some of the adverse effects of acid fog and rain. Acid rain not only kills trees. It also lowers the pH of surface waters such as ponds and lakes. As a result, the water may become too acidic for fish and other water organisms to survive.

FIGURE 16.8

Acid fog and acid rain killed the trees in this forest.

Even normal (clean) rain is somewhat acidic. That's because carbon dioxide (CO_2) in the air dissolves in raindrops, producing a weak acid called carbonic acid (H_2CO_3), which has a pH of about 5.5. When rainwater soaks into the ground, it can slowly dissolve rocks, particularly those containing calcium carbonate. This is how water forms underground caves.

Q: How do you think acid rain might affect buildings and statues made of stone?

A: Acid rain dissolves and damages stone buildings and statues. The **Figure** 16.9 shows a statue that has been damaged by acid rain.

Summary

- The strength of an acid or base is called acidity. It depends on how much of the substance breaks down into ions when it dissolves in water.
- Acidity is measured by pH, which is the concentration of hydrogen ions in a solution.
- Acidity is an important factor for living things because most can survive only within a relatively narrow range of acidity.

Vocabulary

- **acidity**: Concentration of hydrogen ions in a solution.
- **pH**: Measure of the acidity, or hydrogen ion (H^+) concentration, of a substance.

FIGURE 16.9

Practice

Try arranging substances in order by their pH values at the following URL.

http://www.quia.com/rd/1975.html?AP_rand=1450177352

Review

1. What determines how acidic or basic a solution is?
2. What is pH? What is the pH of a neutral substance?
3. How much more acidic is soapy water than pure water? (*Hint*: See the pH chart in the **Figure** 16.7.)
4. Why is the pH of the environment important for living things?

16.6 The pH Concept

Lesson Objectives

The student will:

- calculate $[\mathrm{H}^+]$ and $[\mathrm{OH}^-]$ for a solution of acid or base.
- define autoionization.
- state the $[\mathrm{H}^+]$, $[\mathrm{OH}^-]$, and K_w values for the autoionization of water.
- define pH and describe the pH scale.
- write the formulas for pH and pOH and express their values in a neutral solution at 25°C.
- explain the relationships among pH, pOH, and K_w.
- calculate $[\mathrm{H}^+]$, $[\mathrm{OH}^-]$, pH, and pOH given the value of any one of the other values.
- explain the relationship between the acidity or basicity of a solution and the hydronium ion concentration, $[\mathrm{H_3O}^+]$, and the hydroxide ion concentration, $[\mathrm{OH}^-]$, of the solution.
- predict whether an aqueous solution is acidic, basic, or neutral from $[\mathrm{H_3O}^+]$, $[\mathrm{OH}^-]$, or the pH.

Vocabulary

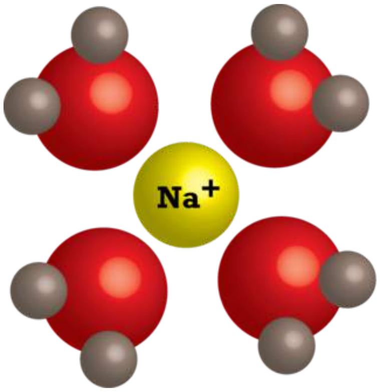
- autoionization
- hydronium ion
- ion product constant for water (K_w)
- pH
- pOH

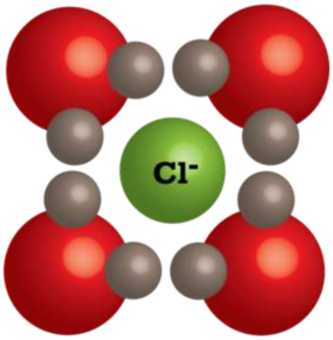
Introduction

We have learned many properties of water, such as pure water does not conduct electricity. The reason pure water does not conduct electricity is because the concentration of ions present when water ionizes is small. In this lesson, we will look a little closer at this property of water and how it relates to acids and bases.

The Hydronium Ion

Recall that ions in solution are hydrated. That is, water molecules are loosely bound to the ions by the attraction between the charge on the ion and the oppositely charged end of the polar water molecules, as illustrated in the figure below. When we write the formula for these ions in solution, we do not show the attached water molecules. It is simply recognized by chemists that ions in solution are always hydrated.





As with any other ion, a hydrogen ion dissolved in water will be closely associated with one or more water molecules. This fact is sometimes indicated explicitly by writing the **hydronium ion**, H_3O^+, in place of the hydrogen ion, H^+. Many chemists still use $H^+_{(aq)}$ to represent this situation, but it is understood that this is just an abbreviation for what is really occurring in solution. You are likely to come across both, and it is important for you to understand that they are actually describing the same entity. When using the hydronium ion in a chemical equation, you may need to add a molecule of water to the other side so that the equation will be balanced. This is illustrated in the equations below. Note that you are not really adding anything to the reaction. The (aq) symbol indicates that the various reaction components are dissolved in water, so writing one of these water molecules out explicitly in the equation does not change the reaction conditions.

$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ (not showing hydronium)
$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$ (showing hydronium)

Relationship Between [H

Even totally pure water will contain a small amount of H^+ and OH^-. This is because water undergoes a process known as autoionization. **Autoionization** occurs when the same reactant acts as both the acid and the base. Look at the reaction below.

$$H_2O_{(aq)} + H_2O_{(aq)} \rightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

The ionization of water is frequently written as:

$$H_2O_{(l)} \rightarrow H^+ + OH^-.$$

The equilibrium constant expression for this dissociation would be $K_w = [H^+][OH^-]$. From experimentation, chemists have determined that in pure water, $[H^+] = 1 \times 10^{-7}$ mol/L and $[OH^-] = 1 \times 10^{-7}$ mol/L.

Because this is a particularly important equilibrium, the equilibrium constant is given a subscript to differentiate it from other reactions. K_w, also known as the **ion product constant for water**, always refers to the autoionization of water. We can the calculate K_w because we know the value of $[H^+]$ and $[OH^-]$ for pure water at 25°C.

$$K_w = [H^+][OH^-]$$
$$K_w = (1 \times 10^{-7})(1 \times 10^{-7})$$
$$K_w = 1 \times 10^{-14}$$

A further definition of acids and bases can now be made:

When $[H_3O^+] = [OH^-]$ (as in pure water), the solution is *neutral*.
When $[H_3O^+] > [OH^-]$, the solution is an *acid*.
When $[H_3O^+] < [OH^-]$, the solution is a *base*.

Stated another way, an acid has a $[H_3O^+]$ that is greater than 1×10^{-7} and a $[OH^-]$ that is less than 1×10^{-7}. A base has a $[OH^-]$ that is greater than 1×10^{-7} and a $[H_3O^+]$ that is less than 1×10^{-7}.

The equilibrium between H^+, OH^-, and H_2O will exist in all water solutions, regardless of anything else that may be present in the solution. Some substances that are placed in water may become involved with either the hydrogen or hydroxide ions and alter the equilibrium state. However, as long as the temperature is kept constant at 25°C, the equilibrium will shift to maintain the equilibrium constant, K_w, at exactly 1×10^{-14}.

For example, a sample of pure water at 25°C has $[H^+]$ equal to 1×10^{-7} M and $[OH^-] = 1 \times 10^{-7}$ M. The K_w for this solution, of course, will be 1×10^{-14}. Suppose some HCl gas is added to this solution so that the H^+ concentration increases. This is a stress to the equilibrium system. Since the concentration of a product is increased, the reverse reaction rate will increase and the equilibrium will shift toward the reactants. The concentrations of both ions will be reduced until equilibrium is re-established. If the final $[H^+] = 1 \times 10^{-4}$ M, we can calculate the $[OH^-]$ because we know that the product of $[H^+]$ and $[OH^-]$ at equilibrium is always 1×10^{-14}.

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$
$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{1 \times 10^{-4}} = 1 \times 10^{-10}\ \text{M}$$

Suppose, on the other hand, something is added to the solution that reduces the hydrogen ion concentration. As soon as the hydrogen ion concentration begins to decrease, the reverse rate decreases and the forward rate will shift the equilibrium toward the products. The concentrations of both ions will be increased until equilibrium is re-established. If the final hydrogen ion concentration is 1×10^{-12} M, we can calculate the final hydroxide ion concentration.

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$
$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{1 \times 10^{-12}} = 1 \times 10^{-2}\ \text{M}$$

Using the K_w expression and our knowledge of the K_w value, as long as we know either the $[H^+]$ or the $[OH^-]$ in a water solution, we can always calculate the value for the other one.

Example:

What would be the $[H^+]$ for a grapefruit found to have a $[OH^-]$ of 1.26×10^{-11} mol/L? Is the solution acidic, basic, or neutral?

Solution:

$$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$$
$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]} = \frac{1 \times 10^{-14}}{1.26 \times 10^{-11}} = 7.94 \times 10^{-4}\ \text{M}$$

Since the $[H^+]$ in this solution is greater than 1×10^{-7} M, the solution is acidic.

pH and pOH

There are a few very concentrated acid and base solutions used in industrial chemistry and laboratory situations. For the most part, however, acid and base solutions that occur in nature, used in cleaning, and used in biochemistry applications are relatively dilute. Most of the acids and bases dealt with in laboratory situations have hydrogen ion concentrations between 1.0 M and 1.0×10^{-14} M. Expressing hydrogen ion concentrations in exponential numbers can become tedious, so a Danish chemist named Søren Sørensen developed a shorter method for expressing acid strength or hydrogen ion concentration with a non-exponential number. This value is referred to as **pH** and is defined by the following equation:

$$\text{pH} = -\log[\text{H}^+],$$

where $\text{p} = -\log$ and H refers to the hydrogen ion concentration. The p from pH comes from the German word *potenz*, meaning power or the exponent of. Rearranging this equation to solve for $[\text{H}^+]$, we get $[\text{H}^+] = 10^{-\text{pH}}$. If the hydrogen ion concentration is between 1.0 M and 1.0×10^{-14} M, the value of the pH will be between 0 and 14.

Example:

Calculate the pH of a solution where $[\text{H}^+] = 0.01\ \text{mol/L}$.

Solution:

$$\begin{aligned}
\text{pH} &= -\log(0.01)\\
\text{pH} &= -\log(1 \times 10^{-2})\\
\text{pH} &= 2
\end{aligned}$$

Example:

Calculate the $[\text{H}^+]$ if the pH is 4.

Solution:

$$\begin{aligned}
[\text{H}^+] &= 10^{-\text{pH}}\\
[\text{H}^+] &= 10^{-4}\\
[\text{H}^+] &= 1 \times 10^{-4}\ \text{mol/L}
\end{aligned}$$

Example:

Calculate the pH of saliva, where $[\text{H}^+] = 1.58 \times 10^{-6}\ \text{mol/L}$.

Solution:

$$\begin{aligned}
\text{pH} &= -\log[\text{H}^+] = -\log(1.58 \times 10^{-6})\\
\text{pH} &= 5.8
\end{aligned}$$

Example:

Fill in the rest of **Table** 16.8.

TABLE 16.8: **Hydrogen ion concentration and corresponding pH**

$[\text{H}^+]$ in mol/L	$-\log[\text{H}^+]$	pH
0.1	1.00	1.00
0.2	0.70	0.70
1×10^{-5}	?	?
?	?	6.00
0.065	?	?
?	?	9.00

Solution:

The completed table is shown below (**Table** 16.9).

TABLE 16.9: Hydrogen ion concentration and corresponding pH

$[H^+]$ in mol/L	$-\log[H^+]$	pH
0.1	1.00	1.00
0.2	0.70	0.70
1.00×10^{-5}	5	5
1.00×10^{-6}	6.00	6.00
0.065	1.19	1.19
1.00×10^{-9}	9.00	9.00

An acid with pH = 1, then, is stronger than an acid with pH = 2 by a factor of 10. Simply put, lower pH values correspond to higher H^+ concentrations and more acidic solutions, while higher pH values correspond to higher OH^- concentrations and more basic solutions. This is illustrated in the figure below. It should be pointed out that there are acids and bases that fall outside the pH range depicted. However, we will confine ourselves for now to those falling within the 0-14 range, which covers $[H^+]$ values from 1.0 M all the way down to 1×10^{-14} M.

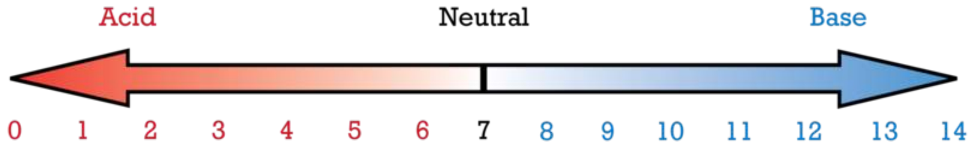
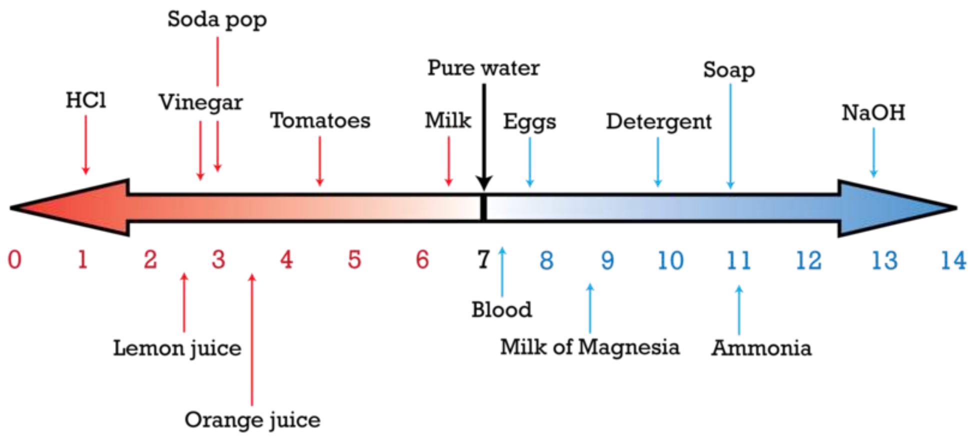


TABLE 16.10: pH versus Acidity

pH level	Solution
pH < 7	Acid
pH = 7	Neutral
pH > 7	Basic

Have you ever cut an onion and had your eyes water up? This is because of a compound with the formula C_3H_6OS that is found in onions. When you cut the onion, a variety of reactions occur that release a gas. This gas can diffuse into the air and eventfully mix with the water found in your eyes to produce a dilute solution of sulfuric acid. This is what irritates your eyes and causes them to water. There are many common examples of acids and bases in our everyday lives. Look at the pH scale below to see how these common examples relate in terms of their pH.



Even though both acidic and basic solutions can be expressed by pH, an equivalent set of expressions exists for the concentration of the hydroxide ion in water. This value, referred to as **pOH**, is defined as:

$$\text{pOH} = -\log[\text{OH}^-]$$

If the pOH is greater than 7, the solution is acidic. If the pOH is equal to 7, the solution is neutral. If the pOH is less than 7, the solution is basic.

If we take the negative log of the complete K_w expression, we obtain:

$$\begin{aligned}
&K_w = [\text{H}^+][\text{OH}^-]\\
&-\log K_w = (-\log[\text{H}^+]) + (-\log[\text{OH}^-])\\
&-\log(1 \times 10^{-14}) = (-\log[\text{H}^+]) + (-\log[\text{OH}^-])\\
&14 = \text{pH} + \text{pOH}
\end{aligned}$$

Therefore, the sum of the pH and the pOH is always equal to 14 (at 25°C). Remember that the pH scale is written with values from 0 to 14 because many useful acid and base solutions fall within this range. Now let's go through a few examples to see how this calculation works for problem-solving in solutions with an added acid or base.

Example:

What is the $[\text{H}^+]$ for a solution of NH_3 whose $[\text{OH}^-] = 8.23 \times 10^{-6}$ mol/L?

Solution:

$$\begin{aligned}
&[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}\\
&[\text{H}_3\text{O}^+] = \frac{1.00\times10^{-14}}{[\text{OH}^-]} = \frac{1.00\times10^{-14}}{8.23\times10^{-6}} = 1.26 \times 10^{-9}\ \text{M}
\end{aligned}$$

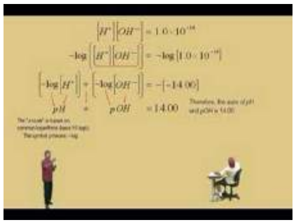
Example:

Black coffee has a $[\text{H}_3\text{O}^+] = 1.26 \times 10^{-5}$ mol/L. What is the pOH?

Solution:

$$\begin{aligned}
&\text{pH} = -\log[\text{H}^+] = -\log 1.26 \times 10^{-5} = 4.90\\
&\text{pH} + \text{pOH} = 14\\
&\text{pOH} = 14 - \text{pH} = 14 - 4.90 = 9.10
\end{aligned}$$

For a classroom demonstration of pH calculations **(5d, 5f; 1e IE Stand.)**, see http://www.youtube.com/watch?v=lca_puB1R8k (9:45).



MEDIA

Click image to the left for more content.

Lesson Summary

- Autoionization is the process where the same molecule acts as both an acid and a base.

- Water ionizes to a very slight degree according to the equation $H_2O_{(l)} \leftrightharpoons [H^+] + [OH^-]$.
- In pure water at 25°C, $[H^+] = [OH^-] = 1.00 \times 10^{-7}$ M.
- The equilibrium constant for the dissociation of water, K_w, is equal to 1.00×10^{-14} at 25°C.
- $\text{pH} = -\log[H^+]$
- $\text{pOH} = -\log[OH^-]$
- $\text{p}K_w = -\log K_w$
- $\text{pH} + \text{pOH} = \text{p}K_w = 14.0$

Further Reading / Supplemental Links

The websites below have more information about pH.

- http://www.johnkyrk.com/pH.html
- http://purchon.com/chemistry/ph.htm

Review Questions

1. What is the $[H^+]$ ion concentration in a solution of 0.350 mol/L H_2SO_4?
   a. 0.175 mol/L
   b. 0.350 mol/L
   c. 0.700 mol/L
   d. 1.42×10^{-14} mol/L
2. A solution has a pH of 6.54. What is the concentration of hydronium ions in the solution?
   a. 2.88×10^{-7} mol/L
   b. 3.46×10^{-8} mol/L
   c. 6.54 mol/L
   d. 7.46 mol/L
3. A solution has a pH of 3.34. What is the concentration of hydroxide ions in the solution?
   a. 4.57×10^{-4} mol/L
   b. 2.19×10^{-11} mol/L
   c. 3.34 mol/L
   d. 10.66 mol/L
4. A solution contains 4.33×10^{-8} M hydroxide ions. What is the pH of the solution?
   a. 4.33
   b. 6.64
   c. 7.36
   d. 9.67
5. Fill in **Table** 16.11 and rank the solutions in terms of increasing acidity.

TABLE 16.11: **Table for Problem 5**

Solutions	$[H^+]$ (mol/L)	$-\log\ [H^+]$	pH
A	0.25	0.60	0.60

TABLE 16.11: (continued)

Solutions	$[H^+]$ (mol/L)	$-\log\,[H^+]$	pH
B	?	2.90	?
C	1.25×10^{-8}	?	?
D	0.45×10^{-3}	?	?
E	?	1.26	?

6. It has long been advocated that red wine is good for the heart. Wine is considered to be an acidic solution. Determine the concentration of hydronium ions in wine with pH 3.81.
7. What does the value of K_w tell you about the autoionization of water?
8. If the pH of an unknown solution is 4.25, what is the pOH?
    a. $10^{-4.25}$
    b. $10^{-9.75}$
    c. 9.75
    d. $14.0 - 10^{-9.75}$
9. A solution contains a hydronium ion concentration of 3.36×10^{-4} mol/L. What is the pH of the solution?
    a. 3.36
    b. 3.47
    c. 10.53
    d. none of the above
10. A solution contains a hydroxide ion concentration of 6.43×10^{-9} mol/L. What is the pH of the solution?
    a. 5.80
    b. 6.48
    c. 7.52
    d. 8.19
11. An unknown solution was found in the lab. The pH of the solution was tested and found to be 3.98. What is the concentration of hydroxide ions in this solution?
    a. 3.98 mol/L
    b. 0.67 mol/L
    c. 1.05×10^{-4} mol/L
    d. 9.55×10^{-11} mol/L

16.7 Titration

Lesson Objectives

The student will:

- explain what an acid/base indicator is.
- explain how acid-base indicators work.
- explain the difference between natural and synthetic indicators.
- explain how indicators are used in the lab.
- explain what a titration is.
- describe how titrations can be used to determine the concentration of an acid or a base in solution.
- explain the difference between the equivalence point and the end point.
- define a standard solution in terms of acid-base titrations.
- calculate the concentration of an acid or base solution using a standard solution.
- calculate the concentration of unknown acid or base when given the concentration of the other and the volume needed to reach the equivalence point in a titration.

Vocabulary

- endpoint
- equivalence point
- natural indicator
- standard solution
- synthetic indicator
- titrant
- titration
- titration curve

Introduction

The typical laboratory procedure for determining the concentration of acid and/or base in a solution is to complete a titration. There are three main types of titration experiments. As we go through this lesson, we will take apply the knowledge we have obtained about acids and bases, chemical reactions, and molarity calculations to the concept of titrations.

Indicators

Recall from the chapter on "Acids-Bases" that an indicator is a substance that changes color at a specific pH and is used to indicate the pH of the solution. One example of an indicator is litmus paper. Litmus paper is paper that has been dipped in a substance that will undergo a color change when it is exposed to either an acid or a base. If red litmus paper turns blue, the solution is basic (pH >7), and if blue litmus turns red the solution is acidic (pH <7).

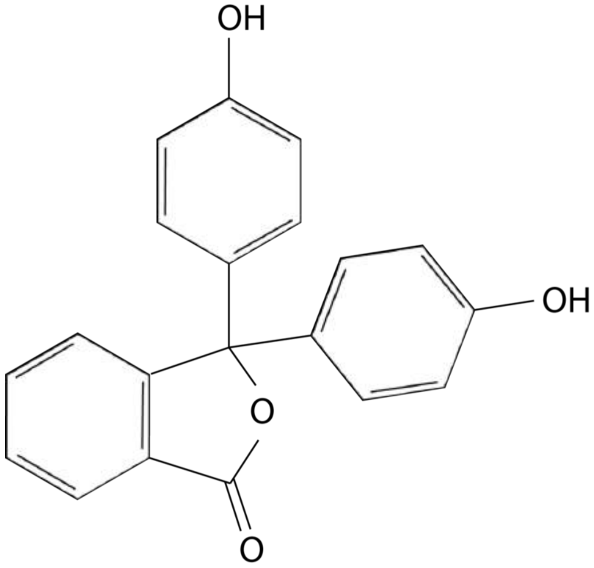
A **natural indicator** is an indicator that is a naturally occurring substance. For example, the juice from red cabbage can be used to prepare an indicator paper. It contains the chemical anthrocyanin, which is the active ingredient in the indicator. Red beets, blueberries, and cranberries are other great examples of naturally occurring indicators. These are all due to the same anthocyanin molecule found in the red cabbage.

Some flowers are also natural indicators. Hydrangea is a common garden plant with flowers that come in many colors, depending on the pH of the soil. A hydrangea plant with blue flowers indicates that the soil is acidic, while creamy white flowers mean the soil is neutral and pink flowers mean the soil is basic.

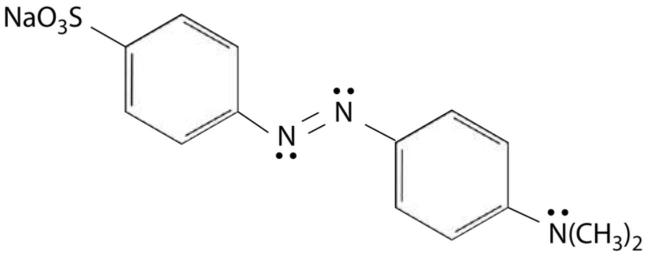
FIGURE 16.10

A hydrangea plant with blue flowers. What does the flower color indicate about the pH of the soil?

Synthetic indicators are compounds created in a chemistry lab rather than compounds found in nature. Both naturally occurring indicators and synthetic indicators are weak organic acids or bases. For example, a common synthetic indicator used in most chemistry laboratories is phenolphthalein. The chemical structure of phenolphthalein is shown in the figure below.

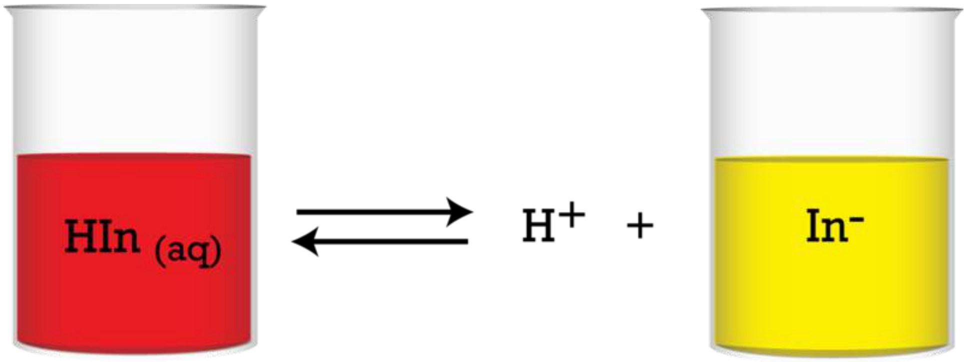


This indicator changes color at a pH of 8.2. Below 8.2 it is colorless, and above 8.2 it is bright pink. There are many common synthetic indicators that are useful in the chemistry laboratory. When dealing with a more acidic range, chemistry students may use methyl orange. The structure for methyl orange is shown below.

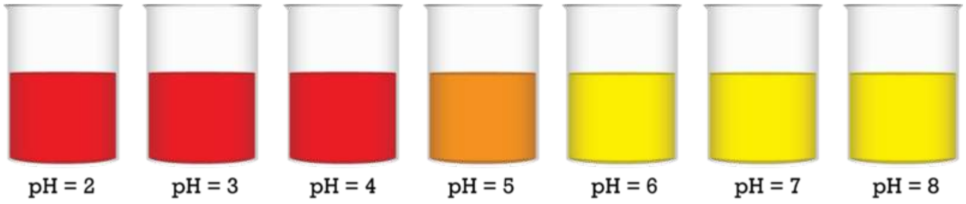


Methyl orange changes color from pH 3.2 to 4.4. Below 3.2, the color of the indicator is red. Above 4.4, the color of the indicator is yellow. In between 3.2 and 4.4, there are various shades of orange, hence the name.

There are two requirements for a substance to function as an acid-base indicator: 1) the substance must have an equilibrium affected by hydrogen ion concentration, and 2) the two forms of the compound on opposite sides of the equilibrium must have different colors. Most indicators function in the same general manner and can be presented by a generic indicator equation. In the equation below, we represent in the indicator ion with a hydrogen ion attached as HIn, and we represent the indicator ion without the hydrogen attached as In^-.



Since the indicator itself is a weak acid, the equilibrium between the protonated form and the anionic form is controlled by the hydrogen ion concentration. For the example above, the protonated form is colored red and the anionic form is colored yellow. If we add hydrogen ion to the solution, the equilibrium will be driven toward the reactants and the solution will turn red. If we add base to the solution (reduce hydrogen ion concentration), the equilibrium will shift toward the products and the solution will turn yellow. It is important to note that if this indicator changes color at pH = 5, then at all pH values less than 5, the solution will be red and at all pH values greater than 5, the solution will be yellow. Therefore, putting this indicator into a solution and having the solution turn yellow does NOT tell you the pH of the solution, it only tells you that the pH is greater than 5. At pH values less than 5, the great majority of the indicator molecules are in the red form and the solution will be red. At pH values greater than 5, the great majority of the indicator particles will be in the yellow form and the solution will be yellow. The equilibrium between these indicator particles is such that the particles will be 50% red form and 50% yellow form at exactly pH = 5. Therefore, at pH = 5, the actual color of the solution will be a 50-50 mixture of red and yellow particles and the solution will be orange, as demonstrated in the figure below.



Many indicators are available to help determine the pH of solutions. A list of the most common indicators is found in **Table** 16.12, along with their respective color change pH values and corresponding color changes.

TABLE 16.12: **Colors and pH Ranges for Common Indicators**

Indicator	pH Range	Color Change
Methyl Violet	0.0 – 1.6	Yellow - Blue
Thymol Blue	1.2 – 2.8	Red - Yellow
Orange IV	1.3 – 3.0	Red - Yellow
Methyl Orange	3.2 – 4.4	Red - Orange
Bromophenol Blue	3.0 – 4.7	Orange/Yellow - Violet
Congo Red	3.0 – 5.0	Blue - Red
Bromocresol Green	3.8 – 5.4	Yellow - Blue
Methyl Red	4.8 – 6.0	Red - Yellow
Litmus	5.0 – 8.0	Red - Blue
Chlorophenol Red	4.8 – 6.2	Yellow - Red
Bromothymol Blue	6.0 – 7.6	Yellow - Blue
Phenol Red	6.4 – 8.2	Yellow - Red/Violet
Thymol Blue	8.0 – 9.6	Yellow - Blue
Phenolphthalein	8.2 – 10.0	Colorless - Pink
Alizarin Yellow R	10.1 – 12.0	Yellow - Red
Methyl Blue	10.6 – 13.4	Blue - Pale Violet
Indigo Carmine	11.4 – 13.0	Blue - Yellow

There are many more indicators than are shown in **Table** 16.12, but these are ones that you may find in common chemistry classroom laboratories. One example of an indicator not found in the table is known as the universal indicator. The universal indicator is a solution that has a different color for each pH from 0 – 14. Universal indicator is produced by creatively mixing many of the individual indicators together so that a different color is achieved for each different pH. It is used for many types of experiments to determine if solutions are acids or bases and where on the pH scale the substance belongs. The chart below indicates the colors of universal indicator for different pH values.

Example:

If the pH of the solution is 4.8, what would be the color of the solution if the following indicators were added?

1. Universal indicator
2. Bromocresol Green
3. Phenol red

Solution:

1. Universal indicator = Orange to orange-yellow
2. Bromocresol Green = green (midway pH = 4.6)
3. Phenol red = yellow

Example:

A solution found in the laboratory was tested with a number of indicators. These were the results:

- Phenolphthalein was colorless
- Bromocresol green was blue
- Methyl red was yellow
- Phenol red was yellow

What was the pH of the solution?

Solution:

- Phenolphthalein was colorless, $\text{pH} < 8.0$
- Bromocresol green was blue, $\text{pH} > 5.4$
- Methyl red was yellow, $\text{pH} > 6.0$
- Phenol red was yellow, $\text{pH} < 6.4$

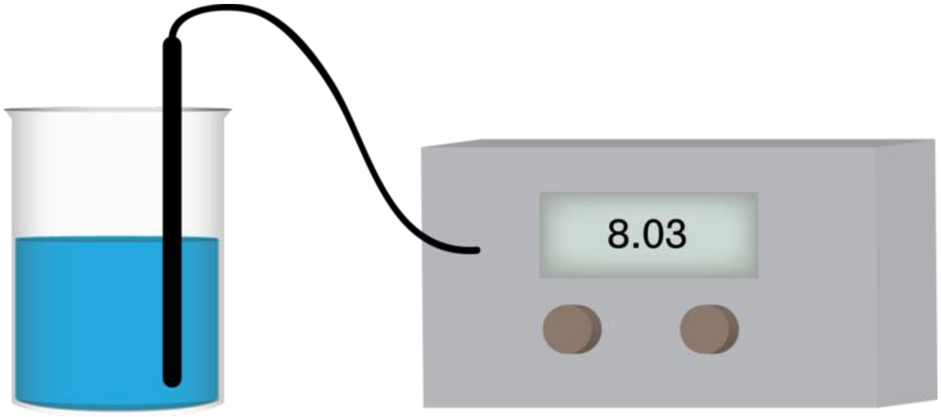
Therefore, the pH of the solution must be between 6.0 and 6.4.

The Titration Process

One of the properties of acids and bases is that they neutralize each other. In the laboratory setting, an experimental procedure where an acid is neutralized by a base (or vice versa) is known as titration. **Titration** is the addition of a known concentration of base (or acid), also called the **titrant**, to a solution of acid (or base) of unknown concentration. Since both volumes of the acid and base are known, the concentration of the unknown solution is then mathematically determined.

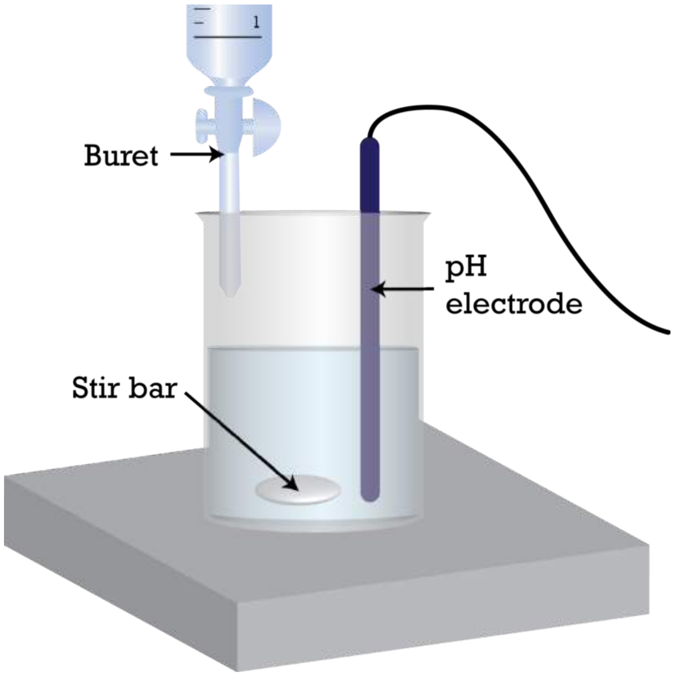
When doing a titration, you need to have a few pieces of equipment. A burette like the one shown below is used to accurately dispense the volume of the solution of known concentration. An Erlenmeyer flask is used to hold a known volume of the solution whose concentration is unknown. A few drops of the indicator are added to the flask before you begin the titration. The **endpoint** is the point where the indicator changes color, which tells us that the acid is neutralized by the base. The **equivalence point** is the point where the number of moles of acid exactly equals the number of moles of base.

Some laboratories have pH meters that measures this point more accurately than the indicator. The diagram below shows a simplified version of a pH meter with the probe from the meter immersed in a mildly alkaline solution (pH = 8.03). The two knobs on the meter are used to calibrate the instrument.



An example of a typical electronic pH meter with the attached probes is shown below. The main purpose of a pH meter in this experiment is to measure the changes in pH as the titration goes from start to finish.

A typical titration setup is shown below. The burette is upright and ready to drip the solution into the flask holding the solution of unknown concentration and the few drops of indicator. When the indicator changes color, the number of moles of acid equals the number of moles of base and the acid (or base) has been neutralized.



There are three types of titrations that are normally performed in the laboratory in order to determine the unknown concentration of the acid or base. These three types are:

1. Strong acid vs. Strong base
2. Strong acid vs. Weak base
3. Weak acid vs. Strong base

In these titrations, a pH meter may be used to measure the changes in the pH as the titration goes to completion. If so, a titration curve can be constructed. A **titration curve** is a graph of the pH versus the volume of titrant added. Let's take a look at how each of these types of titrations differs in terms of their pH curves and their pH at the equivalence point.

(1) Strong Acid vs. Strong Base

For a strong acid vs. a strong base titration, let's assume the strong base is the titrant. Therefore, the Erlenmeyer flask contains the strong acid and a few drops of your indicator. The initial pH of the solution in the flask will likely be low since the solution is a strong acid. As the base is added, the acid is slowly neutralized. At first the change in pH is minimal. This is due to the fact that the flask has a much greater number of H_3O^+ ions than OH^- ions available from the added titrant.

As more and more base is added, more OH^- ions are added and thus more H_3O^+ ions get neutralized. Let's stop here and look at the reaction. The equation below shows the total ionic equation of a reaction between a strong acid and a strong base:

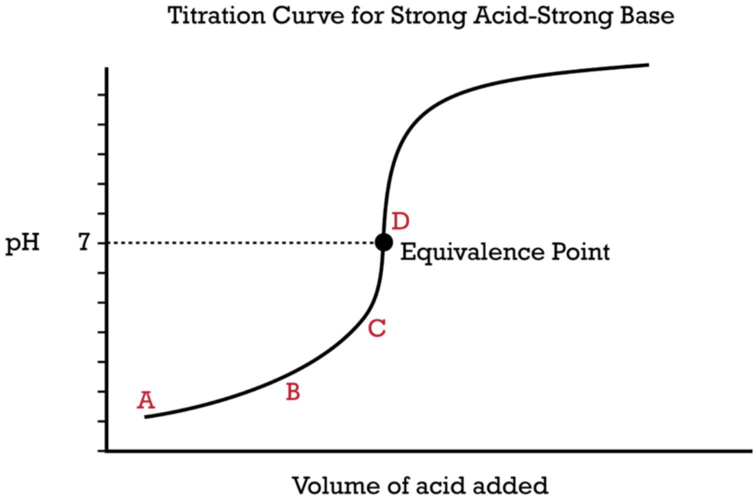
$$H^+_{(aq)} + Cl^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

The next equation shows the net ionic equation for the reaction between the strong acid and the strong base:

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

As we add more OH^- ions, more H_3O^+ (or H^+) ions are being neutralized. Since these two ions react to form water, a neutral solution will eventually be formed. For a strong acid and a strong base, this means the pH = 7.0 at the point of neutralization. If we continue to add the titrant (containing OH^- ions) after all of the H_3O^+ ions have been neutralized, the pH will continue to rise as more base is added and there are excess OH^- ions.

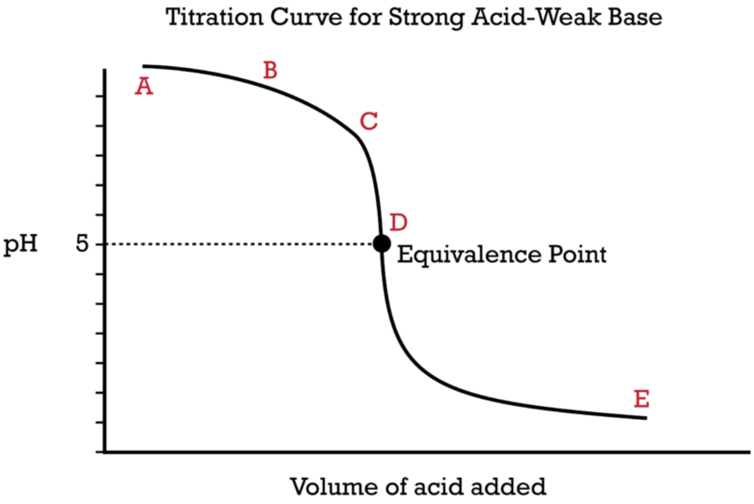
Now that we know what happens in a strong acid-strong base titration, what does the titration curve look like? The main points described above are shown in the titration curve below.



The points A through D sum up the description of the events that take place during the titration. Point A is the start of the titration. Point B is the midpoint, the point where half of the H^+ ions have been neutralized. Point D is the equivalence point.

(2) Strong Acid vs. Weak Base

What would happen if we were to titrate a strong acid with a weak base or vice versa? The titration curve for a weak base-strong acid titration is shown below. Try to determine what is happening in the titration just by looking at the graph.



As the acid (the titrant) is added, the pH decreases as the H_3O^+ ions begin to neutralize the OH^- ions. Point D is the equivalence point. Notice that for a weak base and a strong acid titration, the pH at equivalence point is acidic. The equation for the reaction between NH_3, a weak base, and HCl, a strong acid, is shown below:

$$\mathrm{NH}_{3(aq)} + \mathrm{HCl}_{(aq)} \rightarrow \mathrm{NH_4Cl}_{(aq)} + \mathrm{H_2O}_{(l)}$$

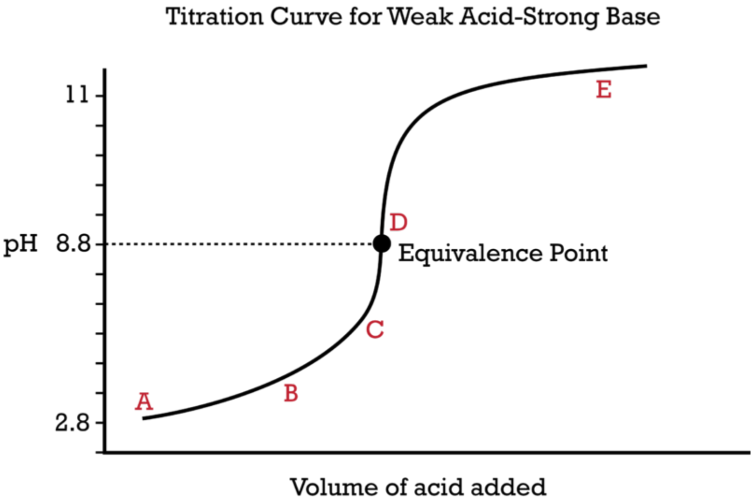
The ionic equation is:

$$\mathrm{NH}_{3(aq)} + \mathrm{H}^+_{(aq)} + \mathrm{Cl}^-_{(aq)} \rightarrow \mathrm{NH}^+_{4(aq)} + \mathrm{Cl}^-_{(aq)} + \mathrm{H_2O}_{(l)}$$

(3) Weak Acid vs. Strong Base

The third type of titration is that of a weak acid with a strong base. When we follow through with the same procedure as the previous two titrations, we can determine a great deal of information simply by looking at the pH curve. For example, let's consider the titration of a solution of acetic acid, $HC_2H_3O_2$, with a solution of potassium hydroxide, KOH. We can write the chemical reaction for this acid-base neutralization and begin to draw a rough sketch of a titration curve:

$$\mathrm{H}^+_{(aq)} + \mathrm{C_2H_3O}^-_{2(aq)} + \mathrm{K}^+_{(aq)} + \mathrm{OH}^-_{(aq)} \rightarrow \mathrm{K}^+_{(aq)} + \mathrm{C_2H_3O}^-_{2(aq)} + \mathrm{H_2O}_{(l)}$$

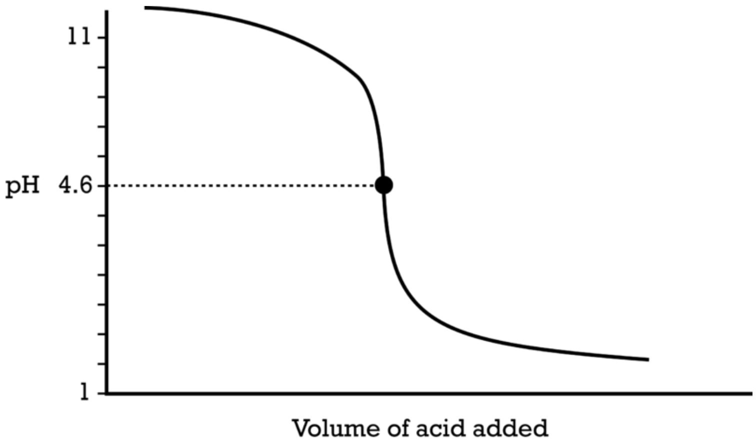


The points on the curve represent the same points as with the other two titration curves. Look, however, at the equivalence point. Notice how the pH for the equivalence point of the weak acid-strong base titration is above 7.0.

Example:

Draw a rough sketch of the titration curve between nitric acid and ethylamine, CH_3NH_2. Assume the acid is in the burette. What is the estimated pH at the equivalence point?

Solution:



The <H at the equivalence point is approximately 4.6 from this graph.

The titrant is the solution of known concentration. For accuracy reasons, this titrant is normally titrated to find its exact concentration before beginning the desired titration. The purpose of this initial titration is to determine, with as much accuracy as possible, the exact concentration of the solution in the burette. To determine the exact concentration of the titrant, we use a standard solution. A **standard solution** is a solution whose concentration is known exactly. Standard solutions have this property because these chemicals are normally found in pure, stable forms. Examples of chemicals used to prepare standard solutions are potassium hydrogen phthalate, $KHC_8H_4O_4$ (sometimes referred to as KHP), and sodium carbonate, Na_2CO_3.

When using a standard solution, the standard is first prepared by dissolving the solid in a known volume of water, adding a few drops of indicator, and titrating with the solution that you want to standardize.

Example:

What is the concentration of a sodium hydroxide solution if 32.34 mL is required to neutralize a solution prepared by dissolving 1.12 g of $KHC_8H_4O_{4(s)}$ in 25.00 mL of H_2O?

Solution:

Step 1: Find the moles of $KHC_8H_4O_4$.

$$\text{moles } KHC_8H_4O_4 = \frac{\text{mass}}{\text{molar mass}} = \frac{1.12\text{g}}{204.2 \text{ g/mol}} = 5.48 \times 10^{-3} \text{ mol}$$

Step 2: Use mole ratio from the reaction to find the moles of NaOH.

$$KHC_8H_4O_{4(aq)} + NaOH_{(aq)} \rightarrow KNaC_8H_4O_{4(aq)} + H_2O_{(l)}$$

Since the reaction is 1:1, 1 mole of KHP reacts with every mole of NaOH.

$$\text{mol NaOH} = 5.48 \times 10^{-3} \text{ mol}$$

Step 3: Determine the concentration of NaOH.

$$[\text{NaOH}] = \frac{5.48 \times 10^{-3} \text{ mol}}{0.03234 \text{ L}} = 0.170 \text{ M}$$

Therefore, the exact concentration of the sodium hydroxide solution used in the titration is 0.170 mol/L.

Choosing an Appropriate Indicator

To choose an appropriate indicator for a titration, a titration curve is useful. Knowing the pH at equivalence for the different types of titrations (see **Table** 16.13) is also necessary.

TABLE 16.13: pH at Equivalence for Titrations

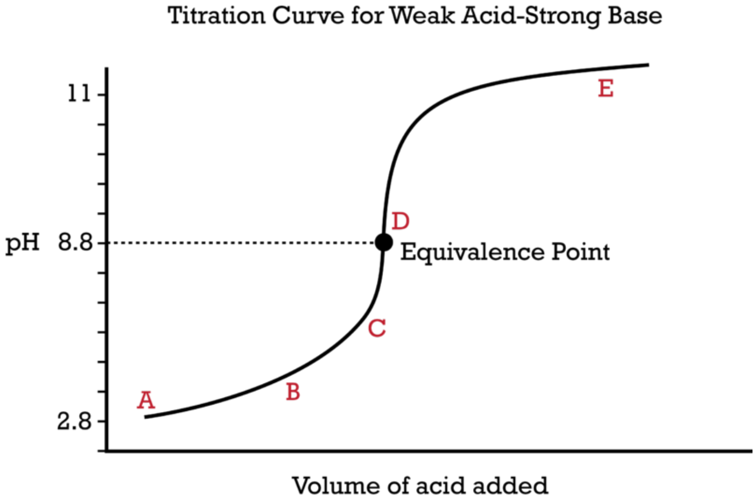
Type of Titration	pH at Equivalence
Strong Acid –Strong Base	pH = 7
Strong Acid –Weak Base	pH <7
Weak Acid –Strong Base	pH >7

Choosing an indicator close to the equivalence point is essential to see the point where all of the H^+ ions and OH^- ions have been neutralized. The color change should occur on or around the equivalence point. So, for example, with a strong acid-strong base titration, the pH at equivalence is 7.0. Indicators such as bromothymol blue (pH range = 6.0 - 7.6) and phenol red (pH range = 6.6 - 8.0) are common. Notice the midpoint color (green) for bromothymol blue would appear at a pH = 6.8, which is close to 7.0. For phenol red, the midpoint color (orange) would appear at pH = 7.3, again close to 7.0.

The same process is used for other titration types. For a strong acid-weak base titration where the pH at equivalence is less than 7, the indicators normally chosen are methyl red (pH range = 4.8 - 6.0) and chlorophenol red (pH range = 4.8 - 6.2). For a weak acid-strong base titration, where the pH at equivalence is greater than 7, the indicators normally chosen are phenolphthalein (pH range = 8.2 - 10) and thymol blue (pH range = 8.0 - 9.6). As with strong acid-strong base titrations, the visual observation of the indicator's midpoint color should signal close proximity to the equivalence point.

Example:

Look at the graph below and determine the appropriate indicator.



Solution:

We first look at the graph and mark the vertical stretch of the titration curve in order to find the half-way mark on this vertical stretch. Looking at the graph, when we follow this half-way mark over to the y-axis, we can see that the equivalence point occurs at approximately pH = 8.8. The indicator appropriate to use would be phenolphthalein pH range = 8.2 - 10). As soon as the pink color forms, we are at the equivalence point.

There is an interesting observation about the endpoint that has yet to be mentioned. The endpoint was defined earlier as the point where the indicator changes color. In an acid-base neutralization reaction, this point may not be the point where all of the H^+ ions have been neutralized by OH^- ions, or vice versa. The experimenter continues titration until the indicator changes color, that is, the endpoint has been reached. The equivalence point is the point where the moles of hydrogen ion and the moles of hydroxide ion are equal. It requires knowledge by the experimenter to select an indicator that will make the endpoint as close as possible to the equivalent point.

The Mathematics of Titration

For the calculations involved here, we will only use our acid and base examples where the stoichiometric ratio of H^+ and OH^- is 1:1. To determine the volume required to neutralize an acid or a base, or in other words, to reach the equivalence point, we will use a formula similar to the dilution formula:

$$M_a \times V_a = M_b \times V_b$$

where M_a is the molarity of the acid, V_a is the volume of the acid, M_b is the molarity of the base, and V_b is the volume of the base. Note that if the acid and base do not neutralize each other in a 1:1 ratio, this equation does not hold true.

Example:

When 10.0 mL of a 0.125 mol/L solution of hydrochloric acid, HCl, is titrated with a 0.100 mol/L solution of potassium hydroxide, KOH, what is the volume of the hydroxide solution required to neutralize the acid? What type of titration is this?

Solution:

Step 1: Write the balanced ionic chemical equation.

$$H^+ + Cl^- + K^+ + OH^- \rightarrow K^+ + Cl^- + H_2O$$

Step 2: Use the formula and fill in all of the given information.

$$M_a \times V_a = M_b \times V_b$$
$$M_a = 0.125 \text{ mol/L}$$
$$V_a = 10.0 \text{ mL}$$
$$M_b = 0.100 \text{ mol/L}$$
$$V_b = ?$$
$$M_a \times V_a = M_b \times V_b$$
$$V_b = \frac{M_a \times V_a}{M_b} = \frac{(0.125 \text{ mol/L})(10.0 \text{ mL}}{0.100 \text{ mol/L}} = 12.5 \text{ mL}$$

Therefore, for this strong acid-strong base titration, the volume of base required is 12.5 mL.

This video shows the technique for performing a titration using an indicator: http://www.youtube.com/watch?v=9DkB82xLvNE (5:03).

MEDIA

Click image to the left for more content.

Lesson Summary

- An indicator is a substance that changes color at a specific pH and is used to indicate the pH of the solution relative to that point.
- A natural indicator is an indicator that is a naturally occurring substance.
- Indicators are normally weak organic acids or bases with complicated structures.
- Universal indicator is a mixture of indicators that produces a different color for each pH from 0 –14.
- A titration is the addition of a known concentration of base (or acid) to a solution of acid (or base) of unknown concentration.
- The titrant is the solution of known concentration. This solution is normally in the burette.
- The endpoint is the point in the titration where the indicator changes color.
- The equivalence point is the point in the titration where the number of moles of acid equals the number of moles of base.
- The three types of titrations usually performed in the laboratory are: strong acid vs. strong base, strong acid vs. weak base, and weak acid vs. strong base.
- A titration curve is a graph of the pH versus the volume of titrant added.
- For a strong acid vs. strong base titration, the pH at equivalence is 7.0. For a strong acid vs. weak base titration, the pH at equivalence is less than 7.0. For a weak acid vs. strong base titration, the pH at equivalence is greater than 7.0.
- A standard solution is a solution whose concentration is known exactly and is used to find the exact concentration of the titrant.
- For titrations where the stoichiometric ratio of mol H^+ to mol OH^- is 1:1, the concentrations or volumes for the unknown acid or base can be calculated with the formula $Ma \times Va = Mb \times Vb$.

Further Reading / Supplemental Links

The following link is to a video about acid-base neutralization and titration.

- http://link.brightcove.com/services/player/bcpid9113583001?bctid=1405713919

The video at the link below shows the lab techniques needed for titration.

- http://chem-ilp.net/labTechniques/TitrationVideo.htm

This video is a ChemStudy film called “Acid Base Indicators.” The film is somewhat dated but the information is accurate.

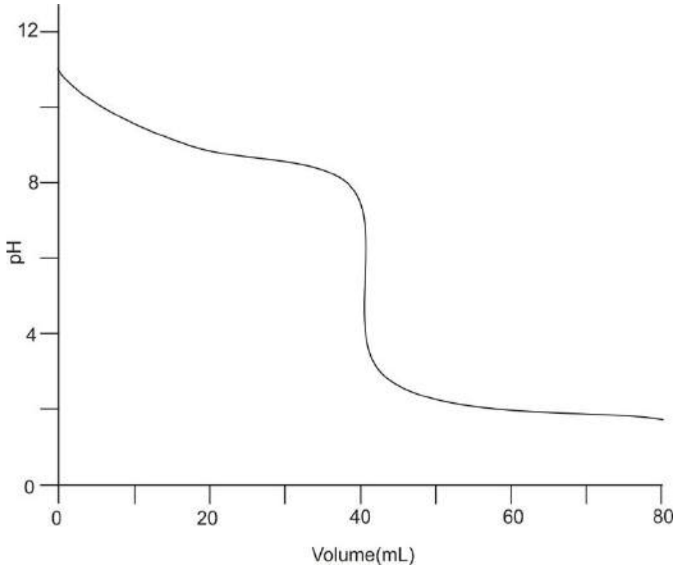
- http://www.youtube.com/watch?v=yi8QrjmV6Sw

Review Questions

1. Why do you think there would be more experimental error when using an indicator instead of a pH meter during a titration?
2. Which of the following definitions best suits that of an endpoint?
    a. The stoichiometric point where the number of moles of acid equals the number of moles of base.

  b. The visual stoichiometric point where the number of moles of acid equals the number of moles of base.
  c. The midpoint of the vertical stretch on the titration curve.
  d. None of the above

3. In the following titration curve, what pair of aqueous solutions would best represent what is shown to be happening in the curve?
   a. $\text{HCOOH}_{(aq)} + \text{NH}_{3(aq)}$
   b. $\text{HCOOH}_{(aq)} + \text{NaOH}_{(aq)}$
   c. $\text{H}_2\text{SO}_{4(aq)} + \text{Ba(OH)}_{2(aq)}$
   d. $\text{HClO}_{4(aq)} + \text{NH}_{3(aq)}$



4. What would be the best indicator to choose for the pH curve shown in question 3?
   a. Methyl red
   b. Litmus
   c. Phenolphthalein
   d. Phenol red
5. What is the best indicator to use in the titration of benzoic acid with barium hydroxide?
   a. Methyl violet, range = 0.0 – 1.6
   b. Bromothymol blue, range = 3.0 – 4.7
   c. Phenolphthalein, range = 8.2 – 10.0
   d. Methyl blue, range = 10.6 – 13.4
   e. Indigo carmine, range = 11.4 – 13.0
6. If 22.50 mL of a sodium hydroxide solution is necessary to neutralize 18.50 mL of a 0.1430 mol/LHNO_3 solution, what is the concentration of NaOH?
   a. 0.1176 mol/L
   b. 0.1430 mol/L
   c. 0.1740 mol/L
   d. 2.64 mol/L
7. Calculate the concentration of hypochlorous acid if 25.00 mL of HClO is neutralized by 32.34 mL of a 0.1320 mol/L solution of sodium hydroxide.

16.8 Calculating pH of Weak Acid and Base Solutions

- Perform calculations to determine the pH of a weak acid or base solution.

Ouch, that hurts!

Bees are beautiful creatures that help plants flourish. They carry pollen from one plant to another to facilitate plant growth and development. But, they can also be troublesome when they sting you. For people who are allergic to bee venom, this can be a serious, life-threatening problem. For the rest of us, it can be a painful experience. When stung by a bee, one first-aid treatment is to apply a paste of baking soda (sodium bicarbonate) to the stung area. This weak base helps with the itching and swelling that accompanies the bee sting.

Calculating pH of Weak Acid and Base Solutions

The K_a and K_b values have been determined for a great many acids and bases, as shown in Tables 21.5 and 21.6. These can be used to calculate the pH of any solution of a weak acid or base whose ionization constant is known.

Sample Problem: Calculating the pH of a Weak Acid

Calculate the pH of a 2.00 M solution of nitrous acid (HNO_2). The K_a for nitrous acid is 4.5×10^{-4}.

Step 1: List the known values and plan the problem.

Known

- initial $[HNO_2] = 2.00$ M

- $K_a = 4.5 \times 10^{-4}$

Unknown

- pH = ?

First, an ICE table is set up with the variable x used to signify the change in concentration of the substance due to ionization of the acid. Then the K_a expression is used to solve for x and calculate the pH.

Step 2: Solve.

TABLE 16.14:

Concentrations	**[HNO_2]**	**[H^+]**	**[$NO_2{}^-$]**
Initial	2.00	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$2.00-x$	x	x

The K_a expression and value is used to set up an equation to solve for x.

$$K_a = 4.5 \times 10^{-4} = \frac{(x)(x)}{2.00-x} = \frac{x^2}{2.00-x}$$

The quadratic equation is required to solve this equation for x. However, a simplification can be made because of the fact that the extent of ionization of weak acids is small. The value of x will be significantly less than 2.00, so the "$-x$" in the denominator can be dropped.

$$\begin{aligned} 4.5 \times 10^{-4} &= \frac{x^2}{2.00-x} \approx \frac{x^2}{2.00} \\ x &= \sqrt{4.5 \times 10^{-4}(2.00)} = 2.9 \times 10^{-2}\ \text{M} = \left[H^+\right] \end{aligned}$$

Since the variable x represents the hydrogen-ion concentration, the pH of the solution can now be calculated.

$$pH = -\log[\text{H}^+] = -\log[2.9 \times 10^{-2}] = 1.54$$

Step 3: Think about your result.

The pH of a 2.00 M solution of a strong acid would be equal to $-\log(2.00) = -0.30$. The higher pH of the 2.00 M nitrous acid is consistent with it being a weak acid and therefore not as acidic as a strong acid would be.

The procedure for calculating the pH of a solution of a weak base is similar to that of the weak acid in the sample problem. However, the variable x will represent the concentration of the hydroxide ion. The pH is found by taking the negative logarithm to get the pOH, followed by subtracting from 14 to get the pH.

Summary

- The procedure for calculating the pH of a weak acid or base is illustrated.

Practice

Perform the calculations at the site below:

http://www.sciencegeek.net/APchemistry/APtaters/pHcalculations.htm

Review

1. What does x stand for in the equation?
2. What simplifying assumption is made?
3. What would x stand for if we were calculating pOH?

16.9 References

1. (A) Becky Cortino; (B) Flickr:isox4. (A) http://www.flickr.com/photos/mediasavvy/8239231530/; (B) http://www.flickr.com/photos/isox4/5167980026/ . CC BY 2.0
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CHAPTER 17 Thermochemistry

Chapter Outline

Chm.2.2.1

- Explain collision theory –molecules must collide in order to react, and they must collide in the correct or appropriate orientation and
  with sufficient energy to equal or exceed the activation energy.
- Interpret potential energy diagrams for endothermic and exothermic reactions including reactants, products, and activated complex.

Chm.2.2.2

• Temperature change –Tie to endothermic/exothermic reaction. Express ΔH as (+) for endothermic and (–) for exothermic.

17.1 Spontaneous and Nonspontaneous Reactions

- Define spontaneous reaction.
- Define nonspontaneous reaction.
- Give examples of spontaneous and nonspontaneous reactions.

Watch that nitro!

Nitroglycerin is tricky stuff. An active ingredient in dynamite (where it is stabilized), “raw” nitroglycerin is very unstable. Physical shock will cause the material to explode. The reaction is shown below:

$$4C_3H_5(ONO_2)_3 \rightarrow 12CO_2 + 10H_2O + 6N_2 + O_2$$

The explosion of nitroglycerin releases large volumes of gases and is very exothermic.

Spontaneous Reactions

Reactions are favorable when they result in a decrease in enthalpy and an increase in entropy of the system. When both of these conditions are met, the reaction occurs naturally. A **spontaneous reaction** is a reaction that favors the formation of products at the conditions under which the reaction is occurring. A roaring bonfire is an example of a spontaneous reaction, since it is exothermic (there is a decrease in the energy of the system as energy is released to the surroundings as heat). The products of a fire are composed partly of gases such as carbon dioxide and water vapor. The entropy of the system increases during a combustion reaction. The combination of energy decrease and entropy increase dictates that combustion reactions are spontaneous reactions.

A **nonspontaneous reaction** is a reaction that does not favor the formation of products at the given set of conditions. In order for a reaction to be nonspontaneous, it must be endothermic, accompanied by a decrease in entropy, or both.

FIGURE 17.1

Bonfire.

Our atmosphere is composed primarily of a mixture of nitrogen and oxygen gases. One could write an equation showing these gases undergoing a chemical reaction to form nitrogen monoxide.

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

Fortunately, this reaction is nonspontaneous at normal temperatures and pressures. It is a highly endothermic reaction with a slightly positive entropy change (ΔS). Nitrogen monoxide is capable of being produced at very high temperatures and has been observed to form as a result of lightning strikes.

One must be careful not to confuse the term spontaneous with the notion that a reaction occurs rapidly. A spontaneous reaction is one in which product formation is favored, even if the reaction is extremely slow. A piece of paper will not suddenly burst into flames, although its combustion is a spontaneous reaction. What is missing is the required activation energy to get the reaction started. If the paper were to be heated to a high enough temperature, it would begin to burn, at which point the reaction would proceed spontaneously until completion.

In a reversible reaction, one reaction direction may be favored over the other. Carbonic acid is present in carbonated beverages. It decomposes spontaneously to carbon dioxide and water according to the following reaction.

$$H_2CO_3(aq) \rightleftarrows CO_2(g) + H_2O(l)$$

If you were to start with pure carbonic acid in water and allow the system to come to equilibrium, more than 99% of the carbonic acid would be converted into carbon dioxide and water. The forward reaction is spontaneous because the products of the forward reaction are favored at equilibrium. In the reverse reaction, carbon dioxide and water are the reactants and carbonic acid is the product. When carbon dioxide is bubbled into water, less than 1% is converted to carbonic acid when the reaction reaches equilibrium. The reverse reaction, as written above, is not spontaneous.

Summary

- Spontaneous and nonspontaneous reactions are defined.
- Examples of both types of reactions are given.

Practice

Read the material at the link below and answer the following questions:

http://chemistry.about.com/od/workedchemistryproblems/a/Entropy-And-Reaction-Spontaneity-Example-Problem.htm

1. Why is system I a spontaneous reaction?
2. Why is system II not spontaneous?
3. Why is system III spontaneous?

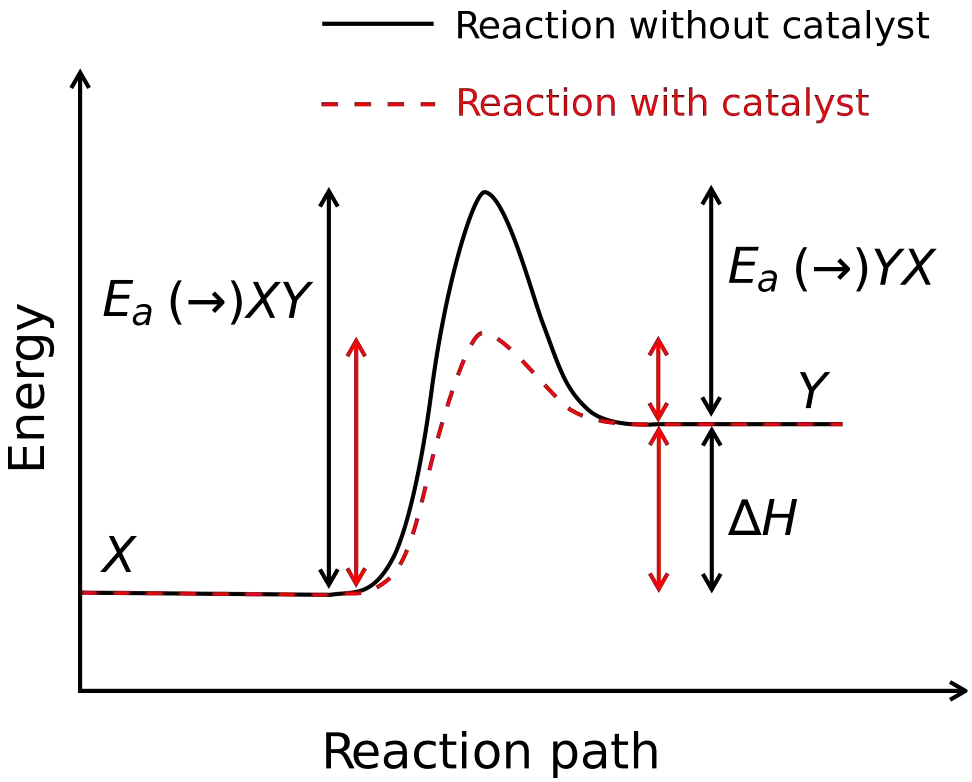
Review

1. Why is a combustion reaction spontaneous?
2. Is NO formation spontaneous at room temperature?
3. How do we know that the equilibrium between carbonic acid and CO_2 goes strongly to the right?

- **nonspontaneous reaction:** A reaction that does not favor the formation of products at the given set of conditions.
- **spontaneous reaction:** A reaction that favors the formation of products at the conditions under which the reaction is occurring.

17.2 Enthalpy

- Define enthalpy.
- List factors that influence enthalpy.



Does the catalyst affect enthalpy?

The factors influencing a reaction are complicated and varied. Since a catalyst affects activation energy, we might assume it would have some sort of impact on enthalpy, but it does not. The change in enthalpy of a reaction depends solely on the chemical compositions of the reactants and products, not on the path taken to get from one to the other.

Enthalpy

Heat changes in chemical reactions are most often measured in the laboratory under conditions in which the reacting system is open to the atmosphere. In that case, the system is at a constant pressure. **Enthalpy** (H) is the heat content of a system at constant pressure. Chemists routinely measure changes in enthalpy of chemical systems as reactants are converted into products. The heat that is absorbed or released by a reaction at constant pressure is the same as the enthalpy change, and is given the symbol ΔH. Unless otherwise specified, all reactions in this material are assumed to take place at constant pressure.

The change in enthalpy of a reaction is a measure of the differences in enthalpy of the reactants and products. The enthalpy of a system is determined by the energies needed to break chemical bonds and the energies needed to form chemical bonds. Energy needs to be put into the system in order to break chemical bonds –they do not come apart spontaneously in most cases. Bond formation to produce products will involve release of energy. The change in enthalpy shows the trade-offs made in these two processes. Does it take more energy to break bonds that that needed to form bonds? If so, the reaction is endothermic and the enthalpy change is positive. If more energy is produced in bond formation than that needed for bond breaking, the reaction is exothermic and the enthalpy is negative.

Several factors influence the enthalpy of a system. Enthalpy is an extensive property, determined in part by the amount of material we work with. The state of reactants and products (solid, liquid, or gas) influences the enthalpy value for a system. The direction of the reaction affects the enthalpy value. A reaction that takes place in the opposite direction has the same numerical enthalpy value, but the opposite sign.

Summary

- Enthalpy is related to the heat of a reaction.
- Factors influencing entropy are described.

Practice

Read the material at the link below and answer the following questions:

http://www.ausetute.com.au/heatreact.html

1. What is the heat of reaction?
2. What units are used?
3. During the experimental measurement of heat of reaction, is heat lost to the surroundings or gained from the surroundings?

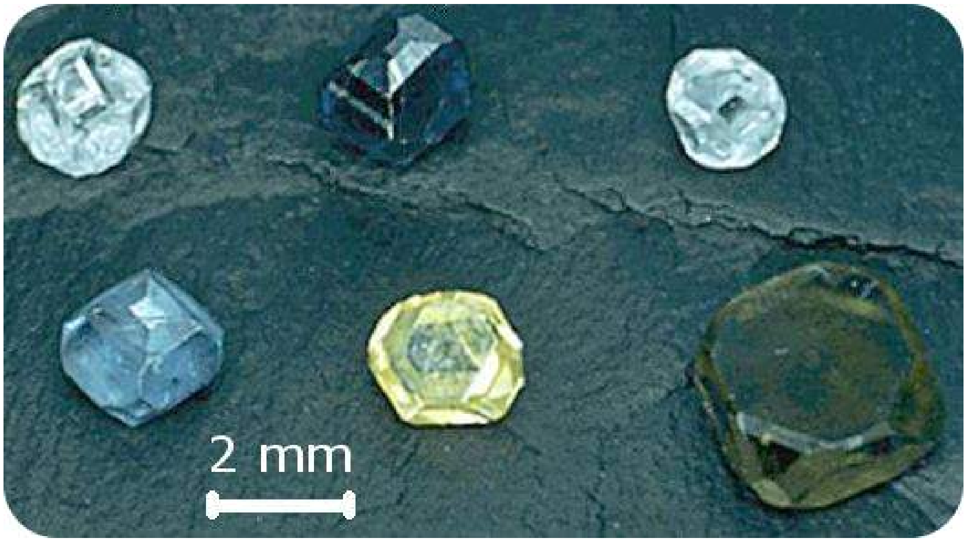
Review

1. What is enthalpy?
2. What is an extensive property?
3. Do the states of reactants and products influence enthalpy values?

- **enthalpy** (H) **:** The heat content of a system at constant pressure.

17.3 Calculating Heat of Reaction from Heat of Formation

- Calculate standard heat of reaction using standard heats of formation.



How are diamonds made?

Natural diamonds are mined from sites around the world. However, the price of natural diamonds is carefully controlled, so other sources for diamonds are being explored. Several different methods for producing synthetic diamonds are available, usually involving treating carbon at very high temperatures and pressures. The diamonds produced are now of high quality, but are primarily used in industrial applications. Diamonds are one of the hardest materials available and are widely used for cutting and grinding tools.

Calculating Heat of Reaction from Heat of Formation

An application of Hess's law allows us to use standard heats of formation to indirectly calculate the heat of reaction for any reaction that occurs at standard conditions. An enthalpy change that occurs specifically under standard conditions is called the **standard enthalpy (or heat) of reaction** and is given the symbol ΔH°. The standard heat of reaction can be calculated by using the following equation.

$$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum n\Delta H_f^\circ(\text{reactants})$$

The symbol Σ is the Greek letter sigma and means "the sum of". The standard heat of reaction is equal to the sum of all the standard heats of formation of the products minus the sum of all the standard heats of formation of the reactants. The symbol "n" signifies that each heat of formation must first be multiplied by its coefficient in the balanced equation.

TABLE 17.1: Standard Heats of Formation of Selected Substances

Substance	ΔH_f° (kJ/mol)	Substance	ΔH_f° (kJ/mol)

TABLE 17.1: (continued)

Al_2O_3(s)	-1669.8	H_2O_2(l)	-187.6
$BaCl_2$(s)	-860.1	KCl(s)	-435.87
Br_2(g)	30.91	NH_3(g)	-46.3
C (s, graphite)	0	NO(g)	90.4
C (s, diamond)	1.90	NO_2(g)	33.85
CH_4(g)	-74.85	NaCl	-411.0
C_2H_5OH(l)	-276.98	O_3(g)	142.2
CO(g)	-110.5	P(s, white)	0
CO_2(g)	-393.5	P(s, red)	-18.4
CaO(s)	-635.6	PbO(s)	-217.86
$CaCO_3$(s)	-1206.9	S(rhombic)	0
HCl(g)	-92.3	S(monoclinic)	0.30
CuO(s)	-155.2	SO_2(g)	-296.1
$CuSO_4$(s)	-769.86	SO_3(g)	-395.2
Fe_2O_3(s)	-822.2	H_2S(g)	-20.15
H_2O(g)	-241.8	SiO_2	-859.3
H_2O(l)	-285.8	$ZnCl_2$	-415.89

Sample Problem: Calculating Standard Heat of Reaction

Calculate the standard heat of reaction (ΔH°) for the reaction of nitrogen monoxide gas with oxygen to form nitrogen dioxide gas.

Step 1: List the known quantities and plan the problem.

Known

- $\Delta H_f{}^\circ$ for $\mathrm{NO(g)} = 90.4$ kJ/mol
- $\Delta H_f{}^\circ$ for $\mathrm{O_2(g)} = 0$ (element)
- $\Delta H_f{}^\circ$ for $\mathrm{NO_2(g)} = 33.85$ kJ/mol

Unknown

- $\Delta H^\circ = ?$ kJ

First write the balanced equation for the reaction. Then apply the equation to calculate the standard heat of reaction for the standard heats of formation.

Step 2: Solve.

The balanced equation is: $\mathrm{2NO(g) + O_2(g) \rightarrow 2NO_2(g)}$

Applying the equation form the text:

$$\begin{aligned}\Delta H^\circ &= [2 \text{ mol } \mathrm{NO_2}(33.85 \text{ kJ/mol})] - [2 \text{ mol NO}(90.4 \text{ kJ/mol}) + 1 \text{ mol } \mathrm{O_2}(0 \text{ kJ/mol})] \\ &= -113 \text{ kJ}\end{aligned}$$

The standard heat of reaction is -113 kJ.

Step 3: Think about your result.

The reaction is exothermic, which makes sense because it is a combustion reaction and combustion reactions always release heat.

Summary

- Standard heats of reaction can be calculated from standard heats of formation.

Practice

Do the practice exercises at the link below:

http://wps.prenhall.com/wps/media/objects/3311/3390593/blb0507.html

Review

1. Is a Hess' Law calculation a direct determination of a standard heat of reaction?
2. What is the purpose of the n in the equation?
3. What does $\sum$ stand for?

- **standard heat of reaction:** An enthalpy change that occurs specifically under standard conditions.

17.4 Stoichiometric Calculations and Enthalpy Changes

- Perform calculations of enthalpy equations.

What will it cost?

There is a growing concern about the damage to the environment from emissions from manufacturing plants. Many companies are taking steps to reduce these harmful emissions by adding equipment that will trap the pollutants. In order to know what equipment (and how many) to order, studies need to be done to measure the amount of product currently produced. The since pollution is often both particulate and thermal, energy changes need to be determined in addition to the amounts of products released.

Stoichiometric Calculations and Enthalpy Changes

Chemistry problems that involve enthalpy changes can be solved by techniques similar to stoichiometry problems. Refer again to the combustion reaction of methane. Since the reaction of 1 mol of methane released 890.4 kJ, the reaction of 2 mol of methane would release $2 \times 890.4 \text{ kJ} = 1781 \text{ kJ}$. The reaction of 0.5 mol of methane would release $\frac{890.4 \text{ kJ}}{2} = 445.2 \text{ kJ}$. As with other stoichiometry problems, the moles of a reactant or product can be linked to mass or volume.

Sample Problem: Calculating Enthalpy Changes

Sulfur dioxide gas reacts with oxygen to form sulfur trioxide in an exothermic reaction according to the following thermochemical equation.

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g) + 198 \text{ kJ}$$

Calculate the enthalpy change that occurs when 58.0 g of sulfur dioxide is reacted with excess oxygen.

Step 1: List the known quantities and plan the problem.

Known

- mass SO_2 = 58.0 g
- molar mass SO_2 = 64.07 g/mol
- $\Delta H = -198$ kJ for the reaction of 2 mol SO_2

Unknown

- $\Delta H =?$ kJ

The calculation requires two steps. The mass of SO_2 is converted to moles. Then the mol SO_2 is multiplied by the conversion factor of $\left(\frac{-198 \text{ kJ}}{2 \text{ mol SO}_2}\right)$.

Step 2: Solve.

$$\Delta H = 58.0 \text{ g SO}_2 \times \frac{1 \text{ mol SO}_2}{64.07 \text{ g SO}_2} \times \frac{-198 \text{ kJ}}{2 \text{ mol SO}_2} = -89.6 \text{ kJ}$$

Step 3: Think about your result.

The mass of sulfur dioxide is slightly less than 1 mol. Since 198 kJ is released for every 2 mol of SO_2 that reacts, the heat released when about 1 mol reacts is one half of 198. The 89.6 kJ is slightly less than half of 198. The sign of ΔH is negative because the reaction is exothermic.

Summary

- Calculations of energy changes in enthalpy equations are described.

Practice

Work the problems at the site below. No peaking at the answers.

http://msmcraesscience.weebly.com/uploads/5/1/4/4/5144948/enthalpy_stoichiometry_worksheet_1.pdf

Review

1. What do you need to determine to solve enthalpy stoichiometry problems?
2. If I react 1.75 moles of methane, how much energy will be involved?
3. I ran a reaction producing sulfur dioxide and releasing 267.3 kJ of energy. How many moles of sulfur dioxide were involved in the reaction?

17.5 References

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CHAPTER 18 The Mole and Stoichiometry

Chapter Outline

Chm.2.2.4 Analyze the stoichiometric relationships inherent in a chemical reaction.

Chm.2.2.5 Analyze quantitatively the composition of a substance (empirical formula, molecular formula, percent composition, and hydrates).

18.1 Avogadro's Number

- Define mole.
- Define Avogadro's number.

Is there an easier way to load this truck?

When the weather is nice, many people begin to work on their yards and homes. For many projects, sand is needed as a foundation for a walk or to add to other materials. You could order up twenty million grains of sand and have people really stare at you. You could order by the pound, but that takes a lot of time weighing out. The best bet is to order by the yard, meaning a cubic yard. The loader can easily scoop up what you need and put it directly in your truck.

Avogadro's Number

It certainly is easy to count bananas or to count elephants (as long as you stay out of their way). However, you would be counting grains of sugar from your sugar canister for a long, long time. Atoms and molecules are extremely small –far, far smaller than grains of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years without stopping you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists needed a name that can stand for a very large number of items. Amedeo Avogadro (1776 - 1856), an Italian scientist, provided just such a number. He is responsible for the counting unit of measure called the mole. A **mole** (mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally

FIGURE 18.1

Italian scientist Amedeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry.

exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as H_2O and CO_2 exist as molecules and so the molecule is their representative particle. For ionic compounds such as NaCl and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number (6.02×10^{23}) of representative particles.

FIGURE 18.2

The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year.

Summary

- A mole of any substance contains Avogadro's number (6.02×10^{23}) of representative particles.

Practice

Use the link below to answer the following questions:

http://www.scientificamerican.com/article.cfm?id=how-was-avogadros-number

1. What was Avogadro's hypothesis?
2. Who first calculated this number?
3. Who coined the term "Avogadro's number"?
4. What contribution did Robert Millikan make to the determination for the value for the number?

Review

1. What is the SI unit for amount of a substance?
2. What is the representative particle for an element?
3. The formula unit is the representative particle for what?

- **Avogadro's number:** The number of representative particles in a mole, 6.02×10^{23}.
- **mole** (mol): The amount of a substance that contains 6.02×10^{23} representative particles of that substance.
- **representative particle**: The smallest unit in which a substance naturally exists.

18.2 Percent Composition

- Define percent composition.
- Perform percent composition calculations.

Nutrition Facts

Serv. Size 2 tbsp (32g)
Servings about 16

Calories 180
Fat Cal. 130

*Percent Daily Values (DV) are based on a 2,000 calorie diet.

Amount Per Serving	% DV*	Amount Per Serving	% DV*
Total Fat 15g	23%	**Total Carbohydrate** 8g	3%
Saturated Fat 2g	10%	Dietary Fiber 2g	8%
Trans Fat 0g		Sugars 3g	
Cholesterol 0mg	0%	**Protein** 7g	
Sodium 140mg	6%		

Vitamin A 0% • Vitamin C 0% • Calcium 2% • Iron 4%

INGREDIENTS: PEANUTS, SUGAR, HYDROGENATED VEGETABLE OIL (RAPESEED, COTTONSEED, SOYBEAN), SALT, MOLASSES.

Is there anything healthy in this jar?

Packaged foods that you eat typically have nutritional information provided on the label. The label on a jar of peanut butter (shown above) reveals that one serving size is considered to be 32 g. The label also gives the masses of various types of compounds that are present in each serving. One serving contains 7 g of protein, 15 g of fat, and 3 g of sugar. By calculating the fraction of protein, fat, or sugar in one serving of size of peanut butter and converting to percent values, we can determine the composition of the peanut butter on a percent by mass basis.

Percent Composition

Chemists often need to know what elements are present in a compound and in what percentage. The **percent composition** is the percent by mass of each element in a compound. It is calculated in a similar way that we just indicated for the peanut butter.

$$\%\text{ by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

Percent Composition from Mass Data

The sample problem below shows the calculation of the percent composition of a compound based on mass data.

Sample Problem One: Percent Composition from Mass

A certain newly synthesized compound is known to contain the elements zinc and oxygen. When a 20.00 g sample of the sample is decomposed, 16.07 g of zinc remains. Determine the percent composition of the compound.

Step 1: List the known quantities and plan the problem.

Known

- mass of compound = 20.00 g
- mass of Zn = 16.07 g

Unknown

- percent Zn = ? %
- percent O = ? %

Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compound to find the percent by mass.

Step 2: Calculate.

$$\begin{aligned}\text{Mass of oxygen} &= 20.00 \text{ g} - 16.07 \text{ g} = 3.93 \text{ g O}\\ \% \text{ Zn} &= \frac{16.07 \text{ g Zn}}{20.00 \text{ g}} \times 100\% = 80.35\% \text{ Zn}\\ \% \text{ O} &= \frac{3.93 \text{ g O}}{20.00 \text{ g}} \times 100\% = 19.65\% \text{ O}\end{aligned}$$

Step 3: Think about your result.

The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Percent Composition from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\% \text{ by mass } = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Sample Problem Two: Percent Composition from Chemical Formula

Dichlorineheptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorineheptoxide.

Step 1: List the known quantities and plan the problem.

Known

- mass of Cl in 1 mol Cl_2O_7 = 70.90 g
- mass of O in 1 mol Cl_2O_7 = 112.00 g
- molar mass of Cl_2O_7 = 182.90 g/mol

Unknown

- percent Cl = ? %

- percent O = ? %

Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

Step 2: Calculate.

$$\% \text{ Cl} = \frac{70.90 \text{ g Cl}}{182.90 \text{ g}} \times 100\% = 38.76\% \text{ Cl}$$
$$\% \text{ O} = \frac{112.00 \text{ g O}}{182.90 \text{ g}} \times 100\% = 61.24\% \text{ O}$$

Step 3: Think about your result.

The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorineheptoxide is 38.76% Cl and 61.24% O. Suppose that you needed to know the masses of chlorine and oxygen present in a 12.50 g sample of dichlorineheptoxide. You can set up a conversion factor based on the percent by mass of each element.

$$12.50 \text{ g Cl}_2\text{O}_7 \times \frac{38.76 \text{ g Cl}}{100 \text{ g Cl}_2\text{O}_7} = 4.845 \text{ g Cl}$$
$$12.50 \text{ g Cl}_2\text{O}_7 \times \frac{61.24 \text{ g O}}{100 \text{ g Cl}_2\text{O}_7} = 7.655 \text{ g O}$$

The sum of the two masses is 12.50 g, the mass of the sample size.

Summary

- Processes are described for calculating the percent composition of a material based on mass or on chemical composition.

Practice

Use the link below to review material and do calculations. Read both parts of the lesson and do as many calculations as you have time for.

http://www.chemteam.info/Mole/Percent-Composition-Part1.html

Review

1. What is the formula for calculating percent composition?
2. What information do you need to calculate percent composition by mass?
3. What do subscripts in a chemical formula tell you?

- **percent composition:** The percent by mass of each element in a compound.

18.3 Conversions Between Moles and Atoms

- Perform calculations involving conversions between number of moles and number of atoms or molecules.

Big numbers or little numbers?

Do you hate to type subscripts and superscripts? Even with a good word-processing program, having to click on an icon to get a superscript and then remembering to click off after you type the number can be a real hassle. If we did not know about moles and just knew about numbers of atoms or molecules (those big numbers that require lots of superscripts), life would be much more complicated and we would make many more typing errors.

Conversions Between Moles and Atoms

Conversions Between Moles and Number of Particles

Using our unit conversion techniques, we can use the mole label to convert back and forth between the number of particles and moles.

Sample Problem 1: Converting Number of Particles to Moles

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Step 1: List the known quantities and plan the problem.

Known

- number of C atoms = 4.72×10^{24}

- 1 mole = 6.02×10^{23}atoms

Unknown

- 4.72×10^{24} = ? mol C

One conversion factor will allow us to convert from the number of C atoms to moles of C atoms.

Step 2: Calculate.

$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C}$$

Step 3: Think about your result.

The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calculation is rounded to three significant figures.

Suppose that you wanted to know how many hydrogen atoms were in a mole of water molecules. First, you would need to know the chemical formula for water, which is H_2O. There are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen would there be in two water molecules? There would be 2 × 2 = 4 hydrogen atoms. How about in a dozen? In that case a dozen is 12 so 12 × 2 = 24 hydrogen atoms in a dozen water molecules. To get the answers, (4 and 24) you had to multiply the given number of molecules by two atoms of hydrogen per molecule. So to find the number of hydrogen atoms in a mole of water molecules, the problem could be solved using conversion factors.

$$1 \text{ mol } H_2O \times \frac{6.02 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O} \times \frac{2 \text{ atoms H}}{1 \text{ molecule } H_2O} = 1.20 \times 10^{24} \text{ atoms H}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.

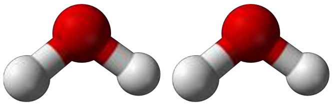

FIGURE 18.3

Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms.

Sample Problem 2: Atoms, Molecules, and Moles

Sulfuric acid has the chemical formula H_2SO_4. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid is the sample?

Step 1: List the known quantities and plan the problem.

Known

- 4.89×10^{25} = O atoms

- 1 mole = 6.02 × 10^{23} molecules H_2SO_4

Unknown

- mol of H_2SO_4 molecules

Two conversion factors will be used. First, convert atoms of oxygen to molecules of sulfuric acid. Then, convert molecules of sulfuric acid to moles of sulfuric acid.

Step 2: Calculate.

$$4.89 \times 10^{25} \text{ atoms O } \times \frac{1 \text{ molecule } H_2SO_4}{4 \text{ atoms O}} \times \frac{1 \text{ mol } H_2SO_4}{6.02 \times 10^{23} \text{ molecules } H_2SO4} = 20.3 \text{ mol } H_2SO_4$$

Step 3: Think about your result.

The original number of oxygen atoms was about 80 times larger than Avogadro's number. Since each sulfuric acid molecule contains 4 oxygen atoms, there are about 20 moles of sulfuric acid molecules.

Summary

- Methods are described for conversions between moles, atoms, and molecules.

Practice

Read the relevant portions of the following article and do problems 3, 5, 9, 13, and 18. Do not worry about the calculations involving conversions dealing with molar mass (that will come next).

http://faculty.rcc.edu/freitas/1AWorksheets/14GramsToMolesToMolecules.pdf

Review

1. What important number do we need to know to do these conversions?
2. I want to convert atoms to moles. My friend tells me to multiply the number of atoms by 6.02 × 10^{23} atoms/mole. Is this correct?
3. Why should I know the formula for a molecule in order to calculate the number of moles of one of the atoms?

18.4 Empirical and Molecular Formulas (8-30-13)

Lesson Objectives

The student will:

- reduce molecular formulas to empirical formulas.
- determine the empirical formula of a compound given either percent composition or masses.
- determine the molecular formula of a compound given percent composition and molar mass.

Introduction

The empirical formula is the simplest ratio of atoms in a compound. Formulas for **ionic** compounds are **always empirical formulas**, but for covalent compounds, the empirical formula is not always the actual formula for the molecule. Molecules such as benzene, C_6H_6, would have an empirical formula of CH.

Finding Empirical Formula from Experimental Data

Empirical formulas can be determined from percent composition (which is always based on mass).

Consider the following example.

Example:

We find that a 2.50 gram sample of a compound contains 0.900 grams of calcium and 1.60 grams of chlorine. The compound contains only these two elements. We can calculate the number of moles of calcium and chlorine atoms in the compound. We can then find the molar ratio of calcium atoms to chlorine atoms. From this, we can determine the empirical formula.

Solution:

First, we convert the mass of each element into moles.

$$\text{moles of Ca} = \frac{0.900\text{ g}}{40.1\text{ g/mol}} = 0.0224\text{ mol Ca}$$
$$\text{moles of Cl atoms} = \frac{1.60\text{ g}}{35.5\text{ g/mol}} = 0.0451\text{ mol Cl}$$

NOTE THE USE OF FLM!

At this point, we have the correct ratio for the atoms in the compound. The formula $Ca_{0.0224}Cl_{0.0451}$, however, isn't acceptable, because empirical formulas are always written with the simplest set of whole numbers.So, we need to find the simplest *whole* number ratio. To find a simple whole number ratio for these numbers, we divide each of them by the smaller number.

$$\text{Ca} = \frac{0.0224}{0.0224} = 1.00 \qquad \text{Cl} = \frac{0.0451}{0.0224} = 2.01$$

Now, we can see the correct empirical formula for this compound is $CaCl_2$.

Finding Empirical Formula from Percent Composition

When finding the empirical formula from percent composition, the first thing to do is to convert the percentages into masses. For example, suppose we are given that the percent composition of a compound as 40.0% carbon, 6.71% hydrogen, and 53.3% oxygen. Since every sample of this compound will have the same composition in terms of the ratio of atoms, we could choose a sample of any size. Suppose we choose a sample size of 100. g **(Can you explain why it would not make any difference to the solution if you had chosen 1.00 g instead?)**. The masses of each of the elements in this sample will be 40.0 grams of carbon, 6.71 grams of hydrogen, and 53.3 grams of oxygen. These masses can then be used to find the empirical formula. You could use any size sample, but choosing a sample size of 100. grams is usually most convenient because it makes the arithmetic simple.

Example:

Find the empirical formula of a compound whose percent composition is 40.0% carbon, 6.71% hydrogen, and 53.3% oxygen.

Solution:

We choose a sample size of 100. grams and multiply this 100. gram sample by each of the percentages to get masses for each element. This would yield 40.0 grams of carbon, 6.71 grams of hydrogen, and 53.3 grams of oxygen. The next step is to convert the mass of each element into moles.

$$\text{moles of C} = \frac{40.0\text{ g}}{12.0\text{ g/mol}} = 3.33\text{ moles C}$$
$$\text{moles of H} = \frac{6.71\text{ g}}{1.01\text{ g/mol}} = 6.64\text{ moles H}$$
$$\text{moles of O} = \frac{53.3\text{ g}}{16.0\text{ g/mol}} = 3.33\text{ mole Ca}$$

Then, we divide all three numbers by the smallest one to get simple whole number ratios:

$$\text{C} = \frac{3.33}{3.33} = 1 \qquad \text{H} = \frac{6.64}{3.33} = 2 \qquad \text{O} = \frac{3.33}{3.33} = 1$$

Finally, we can write the empirical formula CH_2O.

Sometimes, dividing each of the moles by the smallest number does not yield whole numbers. Whenever the subscript for any element in the empirical formula is 1, dividing each of the moles by the smallest will yield a simple whole number ratio, but if none of the elements in the empirical formula has a subscript of 1, then this technique will not yield a simple whole number ratio. In those cases, a little more work is required.

Example:

Determine the empirical formula for a compound that is 66.0% calcium and 34.0% phosphorus.

Solution:

We choose a sample size of 100. grams and multiply the 100. grams by the percentage of each element to get masses. This yields 66.0 grams of calcium and 34.0 grams of phosphorus. We then divide each of these masses by their molar mass to convert the masses into moles:

$$\text{moles of Ca} = \frac{66.0\text{ g}}{40.1\text{ g/mol}} = 1.65\text{ moles Ca}$$
$$\text{moles of P} = \frac{34.0\text{ g}}{31.0\text{ g/mol}} = 1.10\text{ moles P}$$

We then divide each of these moles by the smallest.

$$\text{Ca} = \frac{1.65}{1.10} = 1.50 \qquad \text{P} = \frac{1.10}{1.10} = 1.00$$

In this case, dividing each of the numbers by the smallest one does not yield a simple whole number ratio. In such a case, we must multiply both numbers by some factor that will produce a whole number ratio. If we multiply each of these by 2, we get a whole number ratio of 3 Ca to 2 P. Therefore, the empirical formula is Ca_3P_2.

Finding Molecular Formulas

Empirical formulas show the *simplest whole number ratio (SWNR)* of the atoms in a compound. Molecular formulas show the actual number of atoms of each element in a compound. When you find a empirical formula from either masses of elements or from percent composition, you are finding the empirical formula. For the compound N_2H_4, you will get an empirical formula of NH_2, and for C_3H_6, you will get CH_2. If we want to determine the actual molecular formula, we need one additional piece of information. The molecular formula is always a whole number multiple of the empirical formula. In order to get the molecular formula for N_2H_4, you must double each of the subscripts in the empirical formula. Since the molecular formula is a whole number multiple of the empirical formula, the molecular mass will be the same whole number multiple of the formula mass. The formula mass for NH_2 is 16 g/mol, and the molecular mass for N_2H_4 is 32 g/mol. When we have the empirical formula and the molecular mass for a compound, we can divide the formula mass into the molecular mass and find the whole number that we need to multiply each of the subscripts in the empirical formula.

Example:

Find the molecular formula for a compound with percent composition of 40.0% carbon, 67.1% hydrogen, and 53.3% oxygen. The molecular mass of the compound is 180 g/mol.

***problem solving note*: Resist the temptation use the given molar mass right off the bat. Put the molar mass in a lock box - you need a 'key' to unlock it. That key is the empircal formula which must first be obtained from the percent composition information!**

Solution:

This is the same as an earlier example, except now we also have the molecular mass of the compound. Earlier, we determined the empirical formula of this compound to be CH_2O. The empirical formula has a formula mass of 30.0 g/mol. In order to find the molecular formula for this compound, we divide the formula mass into the molecular mass (180 divided by 30) and find the multiplier for the empirical formula to be 6. As a result, the molecular formula for this compound will be $C_6H_{12}O_6$.

Example:

Find the molecular formula for a compound with percent composition of 85.6% carbon and 14.5% hydrogen. The molecular mass of the compound is 42.1 g/mol.

Solution:

We choose a sample size of 100. g and multiply each element percentage to get masses for the elements in this sample. This yields 85.6 g of C and 14.5 g of H. Dividing each of these by their atomic mass yields 7.13 moles of C and 14.4 moles of H. Dividing each of these by the smallest yields a whole number ratio of 1 carbon to 2 hydrogen. Thus, the empirical formula will be CH_2.

Now you have the key for the lock box!

Since a key goes 'into' the lock, you must divide the empirical formula mass into the molar mass!

The formula mass of CH_2 is 14 g/mol. Dividing 14 g/mol into the molecular mass of 42.1 g/mol yields a multiplier of 3. The molecular formula will be C_3H_6.

Lesson Summary

- The empirical formula of a compound indicates the simplest whole number ratio of atoms present in the compound.
- The empirical formula of a compound can be calculate from the masses of the elements in the compound or from the percent composition.
- The molecular formula of a compound is some whole number multiple of the empirical formula.

Further Reading / Supplemental Links

This website has solved example problems for a number of topics covered in this lesson, including the determination of empirical and molecular formulas.

- http://www.sciencejoywagon.com/chemzone/05chemical-reactions/

Review Questions

1. What is the empirical formula for C_8H_{18}?
2. What is the empirical formula for C_6H_6?
3. What is the empirical formula for WO_2?
4. A compound has the empirical formula C_2H_8N and a molar mass of 46 g/mol. What is the molecular formula of this compound?
5. A compound has the empirical formula C_2H_4NO. If its molar mass is 116.1 g/mol, what is the molecular formula of the compound?
6. A sample of pure indium chloride with a mass of 0.5000 grams is found to contain 0.2404 grams of chlorine. What is the empirical formula of this compound?
7. Determine the empirical formula of a compound that contains 63.0 grams of rubidium and 5.90 grams of oxygen.
8. Determine the empirical formula of a compound that contains 58.0%Rb, 9.50%N, and 32.5%O.
9. Determine the empirical formula of a compound that contains 33.3%Ca, 40.0%O, and 26.7%S.
10. Find the molecular formula of a compound with percent composition 26.7%P, 12.1%N, and 61.2%Cl and with a molecular mass of 695 g/mol.

molecular formula
a formula indicating the actual number of each kind of atom contained in a molecule

empirical formula
a formula indicating the simplest whole number ratio of each kind of atom contained in a molecule

18.5 Percent of Water in a Hydrate

- Define hydrate.
- Calculate the percent water in hydrate when give relevant data.

Why does the color change?

If you look at a typical bottle of copper sulfate, it will be a bluish-green. If someone tells you that copper sulfate is white, you won't believe them. You are both right; it just depends on the copper sulfate. Your blue-green copper sulfate has several water molecules attached to it while your friend's copper sulfate is anhydrous (no water attached). Why the difference? The water molecules interact with some of the d electrons in the copper ion and produce the color. When the water is removed, the electron configuration changes and the color disappears.

Percent of Water in a Hydrate

Many ionic compounds naturally contain water as part of the crystal lattice structure. A **hydrate** is a compound that has one or more water molecules bound to each formula unit. Ionic compounds that contain a transition metal are often highly colored. Interestingly, it is common for the hydrated form of a compound to be of a different color than the **anhydrous** form, which has no water in its structure. A hydrate can usually be converted to the anhydrous compound by heating. The **Figure** below shows the anhydrous compound cobalt(II) chloride to be blue while the hydrate is a distinctive magenta color.

The hydrated form of cobalt(II) chloride contains six water molecules in each formula unit. The name of the compound is cobalt(II) chloride hexahydrate and its formula is $CoCl_2•6H_2O$. The formula for water is set apart at the end of the formula with a dot, followed by a coefficient that represents the number of water molecules per formula unit.

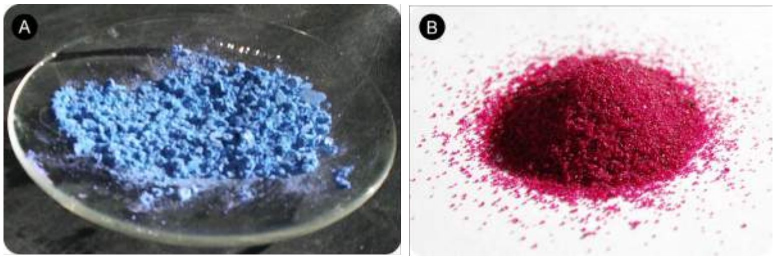

FIGURE 18.4

On the left is anhydrous cobalt(II) chloride, $CoCl_2$. On the right is the hydrated form of the compound called cobalt(II) chloride hexahydrate, $CoCl_2$•$6H_2O$.

It is useful to know the percent of water contained within a hydrate. The sample problem below demonstrates the procedure.

Sample Problem One: Percent of Water in a Hydrate

Find the percent water in cobalt(II) chloride hexahydrate, $CoCl_2$•$6H_2O$.

Step 1: List the known quantities and plan the problem.

The mass of water in the hydrate is the coefficient (6) multiplied by the molar mass of H_2O. The molar mass of the hydrate is the molar mass of the $CoCl_2$ plus the mass of water.

Known

- mass of H_2O in 1 mol hydrate = 108.12 g
- molar mass of hydrate = 237.95 g/mol

Unknown

- percent H_2O = ? %

Calculate the percent by mass of water by dividing the mass of H_2O in 1 mole of the hydrate by the molar mass of the hydrate and multiplying by 100%.

Step 2: Calculate.

$$\%\ H_2O = \frac{108.12\ \text{g}\ H_2O}{237.95\ \text{g}} \times 100\% = 45.44\%\ H_2O$$

Step 3: Think about your result.

Nearly half of the mass of the hydrate is composed of water molecules within the crystal.

Summary

- The process of calculating the percent water in a hydrate is described.

Practice

Use the following link to practice calculating percent water in a hydrate:

http://www.sd84.k12.id.us/shs/departments/science/martz/2007_ssem2/Chemistry/hydrate.htm

Review

1. What is a hydrate?
2. How can you convert a hydrate to an anhydrous compound?
3. What does hexahydrate mean?

- **Anhydrous:** Without water.
- **Hydrate:** A compound that has one or more water molecules bound to each formula unit

18.6 Everyday Stoichiometry

- Define stoichiometry.
- Describe everyday applications of the concept of stoichiometry.

How much equipment do you need for an experiment?

You are in charge of setting out the lab equipment for a chemistry experiment. If you have twenty students in the lab (and they will be working in teams of two) and the experiment calls for three beakers and two test tubes, how much glassware do you need to set out? Figuring this out involves a type of balanced equation and the sort of calculations you would do for a chemical reaction.

Everyday Stoichiometry

You have learned about chemical equations and the techniques used in order to balance them. Chemists use balanced equations to allow them to manipulate chemical reactions in a quantitative manner. Before we look at a chemical reaction, let's consider the equation for the ideal ham sandwich.

FIGURE 18.5

The ideal ham sandwich.

Our ham sandwich is composed of 2 slices of ham (H), a slice of cheese (C), a slice of tomato (T), 5 pickles (P), and 2 slices of bread (B). The equation for our sandwich is shown below:

$$2\text{H} + \text{C} + \text{T} + 5\text{P} + 2\text{B} \rightarrow \text{H}_2\text{CTP}_5\text{B}_2$$

Now let us suppose that you are having some friends over and need to make five ham sandwiches. How much of each sandwich ingredient do you need? You would take the number of each ingredient required for one sandwich (its coefficient in the above equation) and multiply by five. Using ham and cheese as examples and using a conversion factor, we can write:

$$5\text{H}_2\text{CTP}_5\text{B}_2 \times \frac{2\text{ H}}{1\text{ H}_2\text{CTP}_5\text{B}_2} = 10\text{ H}$$
$$5\text{H}_2\text{CTP}_5\text{B}_2 \times \frac{1\text{ C}}{1\text{ H}_2\text{CTP}_5\text{B}_2} = 5\text{ C}$$

The conversion factors contain the coefficient of each specific ingredient as the numerator and the formula of on sandwich as the denominator. The result is what you would expect. In order to make five ham sandwiches, you would need 10 slices of ham and 5 slices of cheese.

This type of calculation demonstrates the use of stoichiometry. **Stoichiometry** is the calculation of the amount of substances in a chemical reaction from the balanced equation. The sample problem below is another stoichiometry problem involving ingredients of the ideal ham sandwich.

Sample Problem: Ham Sandwich Stoichiometry

Kim looks in the refrigerator and finds that she has 8 slices of ham. In order to make as many sandwiches as possible, how many pickles does she need? Use the equation above.

Step 1: List the known quantities and plan the problem.

Known

- have 8 ham slices (H)

- 2 H = 5 P (conversion factor)

Unknown

- How many pickles (P) needed?

The coefficients for the two reactants (ingredients) are used to make a conversion factor between ham slices and pickles.

Step 2: Solve.

$$8\ H \times \frac{5\ P}{2\ H} = 20\ P$$

Since 5 pickles combine with 2 ham slices in each sandwich, 20 pickles are needed to fully combine with 8 ham slices.

Step 3: Think about your result.

The 8 ham slices will make 4 ham sandwiches. With 5 pickles per sandwich, the 20 pickles are used in the 4 sandwiches.

Summary

- An example of everyday stoichiometry is given.

Practice

Use the link below to answer the following questions:

http://www.chem4kids.com/files/react_stoichio.html

1. What does stoichiometry help you figure out?
2. What are all reactions dependent upon?
3. If I have ten hydrogen molecules and three oxygen molecules, how many molecules of water can I make?
4. What will be left over and how much?

Review

1. I don't like pickles. What would my ideal ham sandwich be?
2. How does that change the equation?
3. Will this change affect the amounts of other materials?

- **stoichiometry:** The calculation of amounts of substances in a chemical reaction from the balanced equation.

18.7 Mass-Mole and Mole-Mass Stoichiometry

- Perform calculations involving conversions of mass to moles.
- Perform calculations involving conversions of moles to mass.

Need nails?

When you are doing a large construction project, you have a good idea of how many nails you will need (lots!). When you go to the hardware store, you don't want to sit there and count out several hundred nails. You can buy nails by weight, so you determine how many nails are in a pound, calculate how many pounds you need, and you're on your way to begin building.

While the mole ratio is ever-present in all stoichiometry calculations, amounts of substances in the laboratory are most often measured by mass. Therefore, we need to use mole-mass calculations in combination with mole ratios to solve several different types of mass-based stoichiometry problems.

Mass to Moles Problems

In this type of problem, the mass of one substance is given, usually in grams. From this, you are to determine the amount in moles of another substance that will either react with or be produced from the given substance.

$$\text{mass of } \textit{given} \rightarrow \text{moles of } \textit{given} \rightarrow \text{moles of } \textit{unknown}$$

The mass of the given substance is converted into moles by use of the molar mass of that substance from the periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation.

Sample Problem: Mass-Mole Stoichiometry

Tin metal reacts with hydrogen fluoride to produce tin(II) fluoride and hydrogen gas according to the following balanced equation.

$$\text{Sn}(s) + 2\text{HF}(g) \rightarrow \text{SnF}_2(s) + \text{H}_2(g)$$

How many moles of hydrogen fluoride are required to react completely with 75.0 g of tin?

Step 1: List the known quantities and plan the problem.

Known

- given: 75.0 g Sn
- molar mass of Sn = 118.69 g/mol
- 1 mol Sn = 2 mol HF (mole ratio)

Unknown

- mol HF

Use the molar mass of Sn to convert the grams of Sn to moles. Then use the mole ratio to convert from mol Sn to mol HF. This will be done in a single two-step calculation.

g Sn → mol Sn → mol HF

Step 2: Solve.

$$75.0 \text{ g Sn} \times \frac{1 \text{ mol Sn}}{118.69 \text{ g Sn}} \times \frac{2 \text{ mol HF}}{1 \text{ mol Sn}} = 1.26 \text{ mol HF}$$

Step 3: Think about your result.

The mass of tin is less than one mole, but the 1:2 ratio means that more than one mole of HF is required for the reaction. The answer has three significant figures because the given mass has three significant figures.

Moles to Mass Problems

In this type of problem, the amount of one substance is given in moles. From this, you are to determine the mass of another substance that will either react with or be produced from the given substance.

$$\text{moles of } given \rightarrow \text{moles of } unknown \rightarrow \text{mass of } unknown$$

The moles of the given substance are first converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Then, the moles of the unknown are converted into mass in grams by use of the molar mass of that substance from the periodic table.

Sample Problem: Mole-Mass Stoichiometry

Hydrogen sulfide gas burns in oxygen to produce sulfur dioxide and water vapor.

$$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(g)$$

What mass of oxygen gas is consumed in a reaction that produces 4.60 mol SO_2?

Step 1: List the known quantities and plan the problem.

Known

- given: 4.60 mol SO_2
- 2 mol SO_2 = 3 mol O_2 (mole ratio)
- molar mass of O_2 = 32.00 g/mol

Unknown

- mass O_2 = ? g

Use the mole ratio to convert from mol SO_2 to mol O_2. Then convert mol O_2 to grams. This will be done in a single two-step calculation.

mol $SO_2 \rightarrow$ mol $O_2 \rightarrow$ g O_2

Step 2: Solve.

$$4.60 \text{ mol } SO_2 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } SO_2} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 221 \text{ g } O_2$$

Step 3: Think about your result.

According to the mole ratio, 6.90 mol O_2 is produced with a mass of 221 g. The answer has three significant figures because the given number of moles has three significant figures.

Summary

- Calculations involving conversions of mass to moles and moles to mass are described.

Practice

Work problems 11-20 at the link below:

http://myweb.astate.edu/mdraganj/Moles1.html

Review

1. In the first problem, what would happen if you multiply grams Sn by 118.69 grams/mole Sn?
2. Why is a balanced equation needed?
3. Does the physical form of the material matter for these calculations?

- **mass-mole calculations:** mass of given → moles of given → moles of unknown
- **mole-mass calculations:** moles of given → moles of unknown → mass of unknown

18.8 Stoichiometry Calculations

Lesson Objectives

- Explain the meaning of the term “stoichiometry.”
- Determine mole ratios in chemical equations.
- Calculate the number of moles of any reactant or product from a balanced equation, given the amount of one reactant or product.
- Calculate the mass of any reactant or product from a balanced equation, given the amount of one reactant or product.

Introduction

Chemical equations provide us with information about the types of compounds that react to form products. They also provide us with the ratios in which these compounds react to form products. In this chapter you will explore the quantitative relationships that exist between the reactants and products in a balanced equation. This is known as stoichiometry.

Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in a balanced chemical equation. The word stoichiometry comes from the Greek words stoikheion, which means element, and metron, which means measure.

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Interpreting Chemical Equations

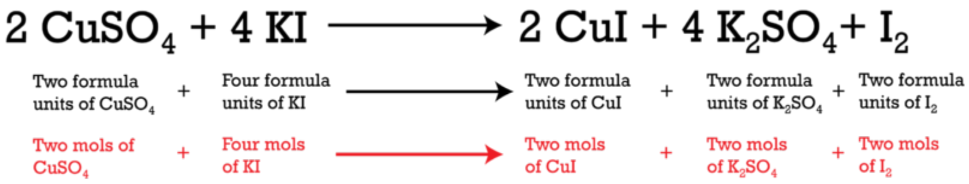
Consider the following reaction:

$$N_2O_3 + H_2O \rightarrow 2\,HNO_2$$

We have learned that the coefficients in a chemical equation tell us the relative amounts of each substance. Therefore, one way to describe the ratios involved in this reaction would be, “One molecule of dinitrogen trioxide plus one molecule of water yields two molecules of nitrous acid.” However, because these are only ratios, this statement would be equally valid using units other than molecules. Thus, we could also say, “One *mole* of dinitrogen trioxide plus one mole of water yields two moles of nitrous acid.” We can use moles instead, because a mole is simply an amount equal to Avogadro’s number, just like a dozen is an amount equal to 12.

It is important that you do not confuse this with units that describe properties other than amount. For example, it would NOT be correct to say that one gram of dinitrogen trioxide plus one gram of water yields two grams of nitrous acid, because each of these molecules has a different mass.

Now consider this reaction:



Here, we can say, "2 moles of copper (II) sulfate react with 4 moles of potassium iodide, yielding two moles of copper (I) iodide, 4 moles of potassium sulfate, and 1 mole of molecular iodine." However, because ionic substances exist as crystal lattices instead of discrete molecules, it is generally not correct to refer to "molecules" of something like KI. Instead, the term "formula unit" is used to describe one potassium ion plus one iodide ion.

Sample Question: Indicate the ratios of compounds involved in each of these balanced chemical equations in both moles and molecules/formula units.

(a) $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

(b) $KBrO_3 + 6\ KI + 5\ HBr \rightarrow 7\ KBr + 3\ I_2 + 3\ H_2O$

Solution:

(a) Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.

Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.

(b) Two formula units of $KBrO_3$ plus six formula units of KI plus six formula units of HBr yields seven formula units of KBr plus three molecules of I_2 and three molecules of H_2O.

Two moles of $KBrO_3$ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I_2 and three moles of H_2O.

Mole Ratios

A mole ratio is the relationship between the number of moles of the substances in a reaction.

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

For example, the following mole ratios can be obtained from the reaction above:

$$\frac{2\ mol\ H_2}{1\ mol\ O_2}\ or\ \frac{1\ mol\ O_2}{2\ mol\ H_2}\ or\ \frac{2\ mol\ H_2O}{2\ mol\ H_2}\ or\ \frac{2\ mol\ H_2O}{1\ mol\ O_2}\ or\ \frac{2\ mol\ H_2}{2\ mol\ H_2O}\ or\ \frac{1\ mol\ O_2}{2\ mol\ H_2O}$$

Using the coefficients of a balanced reaction, you can compare the molar ratios of any two substances in the reaction you are interested in, whether they are reactants or products. The chemical equation MUST always be balanced before the mole ratios are used for calculations.

Mole-Mole Calculations

We have already learned the process through which chemists solve many math problems, dimensional analysis. The mole-mole ratio we obtain from a balanced reaction can be used as a ratio in part of that process.

Example: If only 0.050 mol of magnesium hydroxide, $Mg(OH)_2$, is present, how many moles of phosphoric acid, H_3PO_4, would be required for the reaction?

$$2\ H_3PO_4 + 3\ Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6\ H_2O$$

Solution: We need to set up this problem using dimensional analysis.

Given: 0.050 mol $Mg(OH)_2$

Find: mol H_3PO_4

The ratio we need is one that compares mol $Mg(OH)_2$ to mol H_3PO_4. This is the ratio obtained in the balanced reaction. Note that there are other reactants and products in this reaction, but we don't need to use them to solve this problem.

$$0.050\ \cancel{mol\ Mg(OH)_2} \cdot \frac{2\ mol\ H_3PO_4}{3\ \cancel{mol\ Mg(OH)_2}} = 0.033\ mol\ H_3PO_4$$

Notice that if the equation was not balanced, the amount of H_3PO_4 would have been different. The reaction MUST be balanced to use the reaction in any calculations.

Mass-Mass Calculations

A mass-mass calculation would allow you to solve one of the following types of problems:

- Determine the mass (in grams, kg, etc.) of reactant necessary to product a given amount of product.
- Determine the mass of product that would be produced from a given amount of reactant.
- Determine the mass of reactant necessary to react completely with a second reactant.

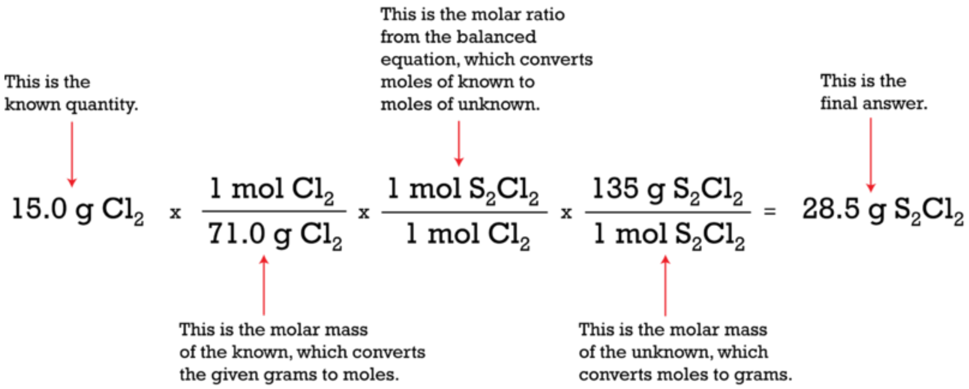
These types of problems are generally solved by using dimensional analysis. Let's look at an example.

Example: 15.0 g of chlorine gas is bubbled over liquid sulfur to produce disulfur dichloride. How much product is produced, in grams, according to the balanced equation:

$$Cl_2(g) + 2\ S(l) \rightarrow S_2Cl_2(l)$$

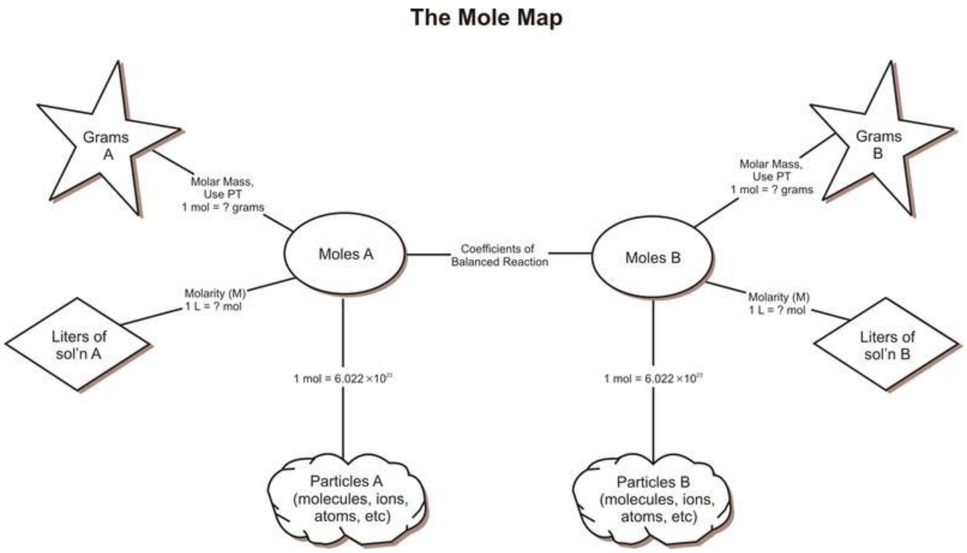
Solution:

a. Identify the **given:** 15.0 $g\ Cl_2$
b. Identify the **find:** $g\ S_2Cl_2$
c. Use the **ratios** that allow you to cancel the units you don't want and get to the unit you are calculating for.



Calculations Using a Mole Map

If we combine the mole-mole ratio with other ratios we have already learned, we have several tools we can use to solve a wide variety of problems. The mole map is a diagram we can use to figure out which ratios to use when solving problems.



You use this map much like you would use a road map. You must first find out where you are on the map (your given units) and where you would like to go (your "find" units). The map will then let you know which roads (ratios) to take to get there. Let's see how this works with a couple of example problems.

Example

The thermite reaction is a very exothermic reaction which produces liquid iron, given by the following balanced equation:

$$Fe_2O_3(s) + 2\,Al(s) \rightarrow 2\,Fe(l) + Al_2O_3(s)$$

If 5.00 g of iron is produced, how much iron(III) oxide was placed in the original container?

Solution

a. Identify the "given": 5.00 g iron. (Even though this is a product, it is still the measurement given to us in the problem. The given may be a reactant or a product.)
b. Identify the units of the "find": $g\ Fe_2O_3$ (remember, mass is measured in grams)
c. Ratios: This is where the map comes in handy. To start with, we are at 5.00 $g\ Fe$. For this problem, then, "A" on the map stands for Fe. We start at grams A.

We want to know $g\ Fe_2O_3$. For this problem, "B" stands for Fe_2O_3. We are heading to grams B.

Our map tells us this problem will take 3 ratios (3 roads from $g\ A$ to $g\ B$): molar mass of A, mol:mol ratio from a balanced reaction, and molar mass of B. To solve our problem, the work will look like:

$$5.00\ \cancel{g\ Fe} \cdot \frac{1\ \cancel{mol\ Fe}}{55.85\ \cancel{g\ Fe}} \cdot \frac{1\ \cancel{mol\ Fe_2O_3}}{2\ \cancel{mol\ Fe}} \cdot \frac{159.7\ g\ Fe_2O_3}{1\ \cancel{mol\ Fe_2O_3}} = 14.3\ g\ Fe_2O_3$$

Sample Question

Ibuprofen is a common painkiller used by many people around the globe. It has the formula $C_{13}H_{18}O_2$. If 200.g of Ibuprofen is combusted how much carbon dioxide is produced? The balanced reaction is:

$$2\ C_{13}H_{18}O_2 + 33\ O_2 \rightarrow 26\ CO_2 + 18\ H_2O$$

Solution

Given: 200.$g\ C_{13}H_{18}O_2$ (g A on the map)

Find: $g\ CO_2$ (g B on the map)

Ratios: The map says we need to use the molar mass of $C_{13}H_{18}O_2$, then the coefficients of the balanced reaction, then the molar mass of CO_2.

$$200.\cancel{g\ C_{13}H_{18}O_2} \cdot \frac{1\ \cancel{mol\ C_{13}H_{18}O_2}}{206.3\ \cancel{g\ C_{13}H_{18}O_2}} \cdot \frac{26\ \cancel{mol\ CO_2}}{2\ \cancel{mol\ C_{13}H_{18}O_2}} \cdot \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} = 555\ g\ CO_2$$

Sample Question

If sulfuric acid is mixed with sodium cyanide, the deadly gas hydrogen cyanide is produced. How many moles of sulfuric acid are required to produce 12.5 g of hydrogen cyanide? The balanced reaction is:

$$2\ NaCN + H_2SO_4 \rightarrow Na_2SO_4 + 2\ HCN$$

Solution

Given: 12.5 $g\ HCN$ (g A on map)

Find: mol H_2SO_4 (mol A on map)

Ratios: The mole map says we need the molar mass of HCN and the coefficients of the balanced reaction.

$$12.5\ g\ HCN \cdot \frac{1\ mol\ HCN}{27.0\ g\ HCN} \cdot \frac{1\ mol\ H_2SO_4}{2\ mol\ HCN} = 0.231\ mol\ H_2SO_4$$

Lesson Summary

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- When the moles of one substance in a reaction is known, the coefficients of the balanced equation can be used to determine the moles of all other substances in the reaction.
- Mass-mass calculations can be done using dimensional analysis, with help from a mole map if needed.

Vocabulary

Stoichiometry

The calculation of quantitative relationships of the reactants and products in a balanced chemical equation.

Mole ratio

The ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation.

Further Reading / Supplemental Links

- Stoichiometry: http://www.lsua.us/chem1001/stoichiometry/stoichiometry.html

Review Questions

1. Distinguish between formula unit, molecule, and mole. Use an example in your answer.
2. Given the reaction between ammonia and oxygen to produce nitrogen monoxide, how many moles of water vapor can be produced from 2 mol of ammonia? The balanced reaction is: $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$
3. When properly balanced, how many moles of bismuth(III) oxide can be produced from 0.625 mol of bismuth? The unbalanced reaction is: $Bi(s) + O_2(g) \rightarrow Bi_2O_3(s)$
4. For the following reaction, balance the equation and then determine the ratio of moles of $B(OH)_3$ to moles of water? $B_2O_3(s) + H_2O(l) \rightarrow B(OH)_3(s)$

Write the balanced chemical equation for the reactions below. When written, find the mol ratios indicated.

1. Gaseous propane, C_3H_8, combusts to form gaseous carbon dioxide and water; find ratio of mol O_2 to mol CO_2.
2. Solid lithium reacts with an aqueous solution of aluminum chloride to produce aqueous lithium chloride and solid aluminum; find ratio of mol $AlCl_3(aq)$ to mol $LiCl(aq)$.
3. An aqueous solution of ammonium hydroxide reacts with an aqueous solution of phosphoric acid to produce aqueous ammonium phosphate and water; find ratio of mol $H_3PO_4(aq)$ to mol $H_2O(l)$.
4. Solid rubidium reacts with solid phosphorous to produce solid rubidium phosphide; find ratio of mol $Rb(s)$ to mol $P(s)$.
5. For the given reaction (balanced): $Ca_3(PO_4)_2 + 3\ SiO_2 + 5\ C \rightarrow 3\ CaSiO_3 + 5\ CO + 2\ P$
   a. how many moles of silicon dioxide are required to react with 0.35 mol of carbon?
   b. how many moles of calcium phosphate are required to produce 0.45 mol of calcium silicate?
6. For the given reaction (balanced), $4\ FeS + 7\ O_2 \rightarrow 2\ Fe_2O_3 + 4\ SO_2$
   a. how many moles of iron(III) oxide are produced from 1.27 moles of oxygen?
   b. how many moles of iron(II) sulfide are required to produce 3.28 moles of sulfur dioxide?
7. Given the reaction between copper (II) sulfide and nitric acid, how many grams of nitric acid will react with 2.00 g of copper(II) sulfide?

$$3\ CuS(s) + 8\ HNO_3(aq) \rightarrow 3\ Cu(NO_3)_2(aq) + 2\ NO(g) + 4\ H_2O(l) + 3\ S(s)$$

8. When properly balanced, what mass of iodine is needed to produce 2.5 g of sodium iodide in the equation below? $I_2(aq) + Na_2S_2O_3(aq) \rightarrow Na_2S_4O_6(aq) + NaI(aq)$
9. Determine the mass of lithium hydroxide produced when 0.38 grams of lithium nitride reacts with water according to the following equation: $Li_3N + 3\ H_2O \rightarrow NH_3 + 3\ LiOH$
10. If 3.01×10^{23} formula units of cesium hydroxide are produced according to this reaction: $2\ Cs + 2\ H_2O \rightarrow 2\ CsOH + H_2$, how many grams of cesium reacted?
11. How many liters of oxygen are necessary for the combustion of 425 g of sulfur, assuming that the reaction occurs at STP? The balanced reaction is: $S + O_2 \rightarrow SO_2$ (hint: one mole of oxygen is 22.4 Liters at STP)
12. If I have 2.0 grams of carbon monoxide, how many molecules of carbon monoxide are there?
13. What mass of oxygen is needed to burn 3.5 g of propane (C_3H_8) is burned according to the following equation: $C_3H_8 + 5\ O_2 \rightarrow 4\ H_2O + 3\ CO_2$
14. How many grams of water are produced if 5 moles of oxygen react according to the following reaction? $2\ H_2 + O_2 \rightarrow 2\ H_2O$

18.9 Mass-Mass Stoichiometry

- Perform calculations involving the determination of the mass of product based on the given mass of the reactant.

How many walnuts are needed to equal 250 grams?

I want to send 250 grams of shelled walnuts to a friend (don't ask why –just go with the question). How many walnuts in shells do I need to buy? To figure this out, I need to know how much the shell of a walnut weighs (about 40% of the total weight of the unshelled walnut). I can then calculate the mass of walnuts that will give me 250 grams of shelled walnuts and then determine how many walnuts I need to buy.

Mass to Mass Problems

Mass-mass calculations are the most practical of all mass-based stoichiometry problems. Moles cannot be measured directly, while the mass of any substance can generally be easily measured in the lab. This type of problem is three steps and is a combination of the two previous types.

$$\text{mass of } given \rightarrow \text{moles of } given \rightarrow \text{moles of } unknown \rightarrow \text{mass of } unknown$$

The mass of the given substance is converted into moles by use of the molar mass of that substance from the periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Finally, the moles of the unknown are converted to mass by use of its molar mass.

Sample Problem: Mass-Mass Stoichiometry

Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$\mathrm{NH_4NO_3}(s) \rightarrow \mathrm{N_2O}(g) + 2\mathrm{H_2O}(l)$$

In a certain experiment, 45.7 g of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Step 1: List the known quantities and plan the problem.

Known

- given: 45.7 g NH_4NO_3
- 1 mol NH_4NO_3 = 1 mol N_2O = 2 mol H_2O (mole ratios)
- molar mass of NH_4NO_3 = 80.06 g/mol
- molar mass of N_2O = 44.02 g/mol
- molar mass of H_2O = 18.02 g/mol

Unknown

- mass N_2O = ? g
- mass H_2O = ? g

Perform two separate three-step mass-mass calculations as shown below.

$$\text{g NH}_4\text{NO}_3 \rightarrow \text{mol NH}_4\text{NO}_3 \rightarrow \text{mol N}_2\text{O} \rightarrow \text{g N}_2\text{O}$$
$$\text{g NH}_4\text{NO}_3 \rightarrow \text{mol NH}_4\text{NO}_3 \rightarrow \text{mol H}_2\text{O} \rightarrow \text{g H}_2\text{O}$$

Step 2: Solve.

$$45.7 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.06 \text{ g NH}_4\text{NO}_3} \times \frac{1 \text{ mol N}_2\text{O}}{1 \text{ mol NH}_4\text{NO}_3} \times \frac{44.02 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}} = 25.1 \text{ g N}_2\text{O}$$
$$45.7 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.06 \text{ g NH}_4 \text{ NO}_3} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol NH}_4\text{NO}_3} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 20.6 \text{ g N}_2\text{O}$$

Step 3: Think about your result.

The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

Summary

- Mass-mass calculations involve converting the mass of a reactant to moles of reactant, then using mole ratios to determine moles of product which can then be converted to mass of product.

Practice

Read the material at the link below, then do the mass-mass problems at the link found at the bottom of the page:

http://www.chemteam.info/Stoichiometry/Mass-Mass.html

Review

1. If matter is neither created nor destroyed, why can't we just go directly from grams of reactant to grams of product?
2. Why is it important to get the subscripts correct in the formulas?
3. Why do the coefficients need to be correct?

- **mass-mass calculations:** mass of given → moles of given → moles of unknown → mass of unknown

18.10 Gas Stoichiometry

- Use the ideal gas law to calculate stoichiometry problems for gases.

The Haber cycle reaction of gaseous nitrogen and hydrogen to form ammonia is a critical step in the production of fertilizer from ammonia. It is important to have an excess of the starting materials so that a maximum yield of ammonia can be achieved. By knowing how much ammonia is needed for manufacture of a batch of fertilizer, the proper amounts of nitrogen and hydrogen gases can be incorporated into the process.

Gas Stoichiometry

You have learned how to use molar volume to solve stoichiometry problems for chemical reactions involving one or more gases at STP. Now, we can use the ideal gas law to expand our treatment of chemical reactions to solve stoichiometry problems for reactions that occur at any temperature and pressure.

Sample Problem: Gas Stoichiometry and the Ideal Gas Law

What volume of carbon dioxide is produced by the combustion of 25.21 g of ethanol (C_2H_5OH) at 54°C and 728 mmHg? Assume the gas is ideal.

Before using the ideal gas law, it is necessary to write and balance the chemical equation. Recall that most combustion reactions, the given substance reacts with O_2 to form CO_2 and H_2O. Here is the balanced equation for the combustion of ethanol.

$$\mathrm{C_2H_5OH}(l) + 3\mathrm{O_2}(g) \rightarrow 2\mathrm{CO_2}(g) + 3\mathrm{H_2O}(l)$$

Step 1: List the known quantities and solve the problem.

Known

- mass $C_2H_5OH = 25.21$ g
- molar mass $C_2H_5OH = 46.08$ g/mol
- $P = 728$ mmHg
- $T = 54°\text{C} = 327$ K

Unknown

- *Volume* $CO_2 = ?$ *L*

The number of moles of carbon dioxide gas is first calculated by stoichiometry. Then the ideal gas law is used to calculate the volume of CO_2 produced.

Step 2: Solve.

$$25.21 \text{ g } \mathrm{C_2H_5OH} \times \frac{1 \text{ mol } \mathrm{C_2H_5OH}}{46.08 \text{ g } \mathrm{C_2H_5OH}} \times \frac{2 \text{ mol } \mathrm{CO_2}}{1 \text{ mol } \mathrm{C_2H_5OH}} = 1.094 \text{ mol } \mathrm{C_2H_5OH}$$

The moles of ethanol (n) is now substituted into $PV = nRT$ to solve for the volume.

$$V = \frac{nRT}{P} = \frac{1.094 \text{ mol} \times 62.36 \text{ L} \cdot \text{mmHg/K} \cdot \text{mol} \times 327 \text{ K}}{728 \text{ mmHg}} = 30.6 \text{ L}$$

Step 3: Think about your result.

The mass of ethanol is slightly more than one half mole, meaning that the mole ratio results in slightly more than one mole of carbon dioxide being produced. Because of the elevated temperature and reduced pressure compared to STP, the resulting volume is larger than 22.4 L.

Summary

- The ideal gas law is used to calculate stoichiometry problems for gases.

Practice

Solve the problems on the worksheet at this site:

http://misterguch.brinkster.net/PRA036.pdf

Review

1. Do we need gas conditions to be at STP to calculate stoichiometry problems?
2. Why do we want to determine the stoichiometry of these reactions?
3. What assumption are we making about the gases involved?

18.11 References

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