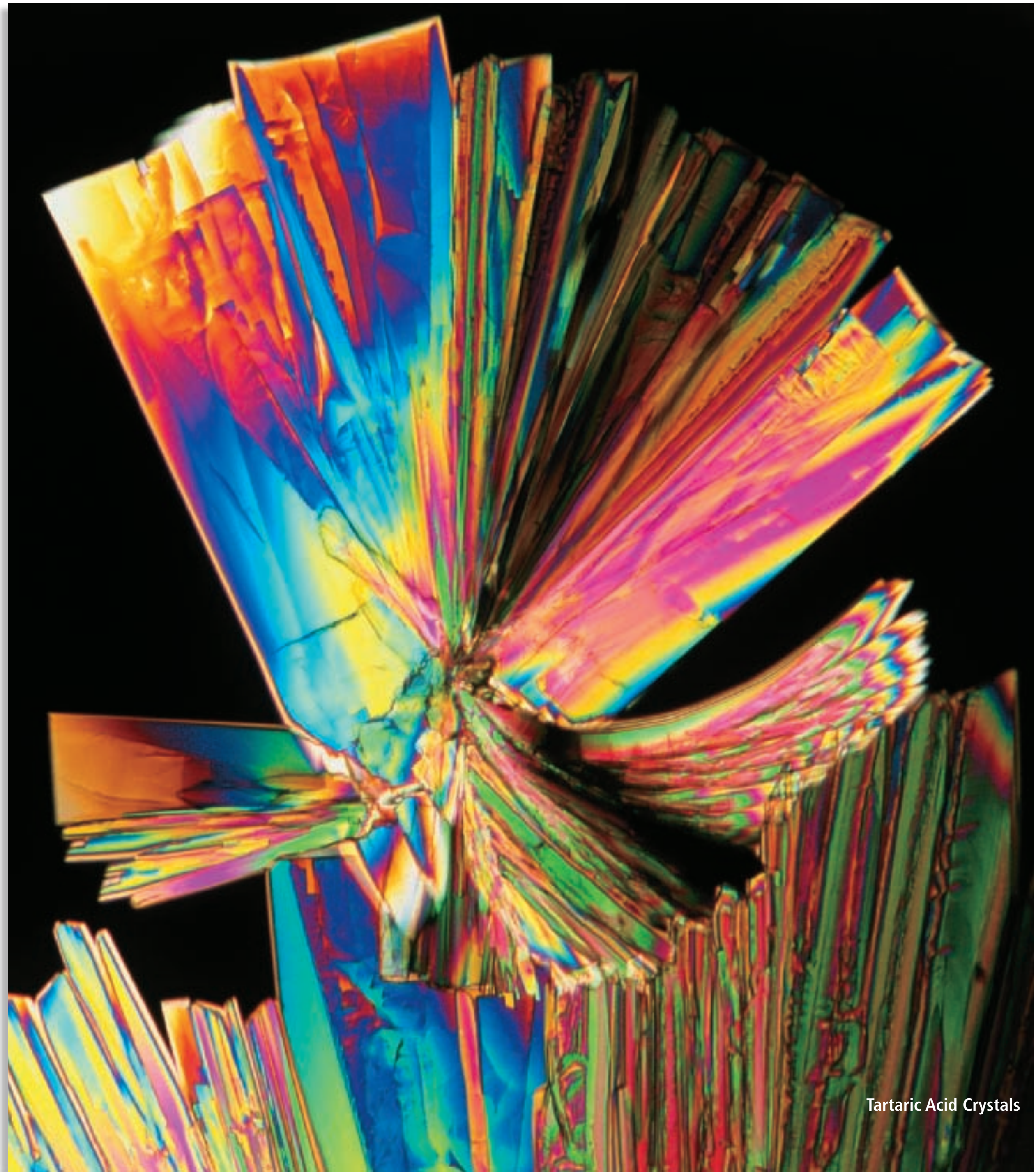
CHAPTER 1

Matter and Change

Chemistry is central to all of the sciences.



Tartaric Acid Crystals

Chemistry Is a Physical Science

SECTION 1

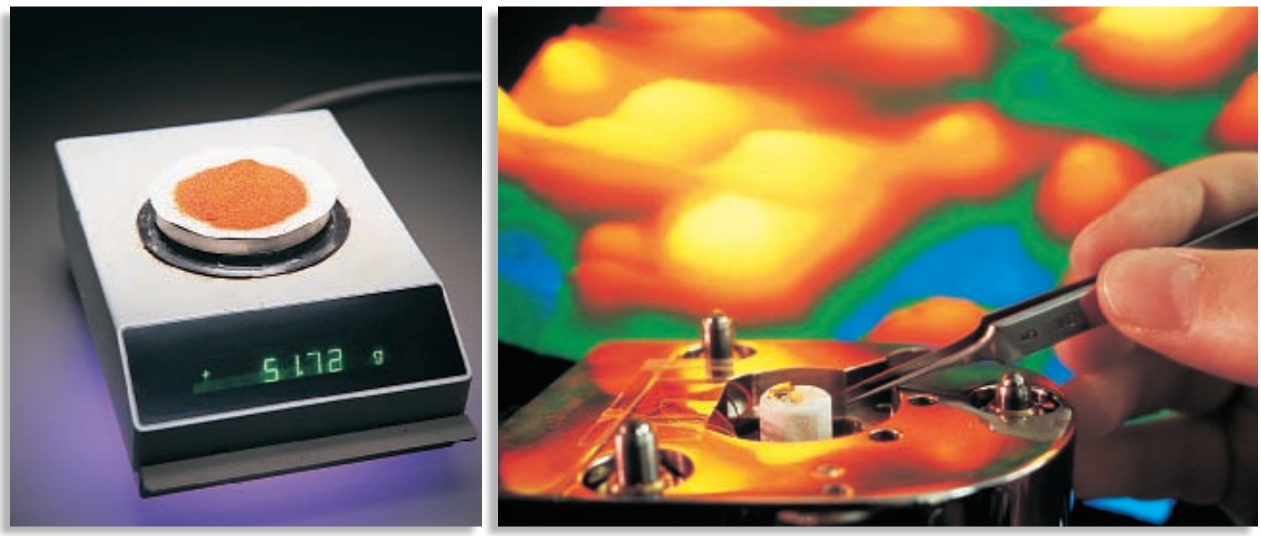
Objectives

- Define *chemistry*.
- List examples of the branches of chemistry.
- Compare and contrast basic research, applied research, and technological development.

The natural sciences were once divided into two broad categories: the biological sciences and the physical sciences. Living things are the main focus of the biological sciences. The physical sciences focus mainly on nonliving things. However, because we now know that both living and nonliving matter consist of chemical structures, chemistry is central to all the sciences, and there are no longer distinct divisions between the biological and physical sciences.

Chemistry *is the study of the composition, structure, and properties of matter, the processes that matter undergoes, and the energy changes that accompany these processes*. Chemistry deals with questions such as, What is a material's makeup? How does a material change when heated, cooled, or mixed with other materials and why does this behavior occur? Chemists answer these kinds of questions during their work.

Instruments are routinely used in chemistry to extend our ability to observe and make measurements. Instruments make it possible, for example, to look at microstructures—things too tiny to be seen with the unaided eye. The scanning electron microscope reveals tiny structures by beaming particles called electrons at materials. When the electrons hit a material, they scatter and produce a pattern that shows the material's microstructure. Invisible rays called X rays can also be used to



(a) (b)

FIGURE 1 A balance (a) is an instrument used to measure the mass of materials. A sample of DNA placed in a scanning tunneling microscope produces an image (b) showing the contours of the DNA's surface.

determine microstructures. The patterns that appear, called X-ray diffraction patterns, can be analyzed to reveal the arrangement of atoms, molecules, or other particles that make up the material. By learning about microstructures, chemists can explain the behavior of macrostructures—the visible things all around you.

Branches of Chemistry

Chemistry includes many different branches of study and research. The following are six main areas, or branches, of study. But like the biological and physical sciences, these branches often overlap.

1. *Organic chemistry*—the study of most carbon-containing compounds
2. *Inorganic chemistry*—the study of non-organic substances, many of which have organic fragments bonded to metals (organometallics)
3. *Physical chemistry*—the study of the properties and changes of matter and their relation to energy
4. *Analytical chemistry*—the identification of the components and composition of materials
5. *Biochemistry*—the study of substances and processes occurring in living things
6. *Theoretical chemistry*—the use of mathematics and computers to understand the principles behind observed chemical behavior and to design and predict the properties of new compounds

In all areas of chemistry, scientists work with chemicals. *A* **chemical** *is any substance that has a definite composition.* For example, consider the material called sucrose, or cane sugar. It has a definite composition in terms of the atoms that compose it. It is produced by certain plants in the chemical process of photosynthesis. Sucrose is a chemical. Carbon dioxide, water, and countless other substances are chemicals as well.

Knowing the properties of chemicals allows chemists to find suitable uses for them. For example, researchers have synthesized new substances, such as artificial sweeteners and synthetic fibers. The reactions used to make these chemicals can often be carried out on a large scale to make new consumer products such as flavor enhancers and fabrics.

Basic Research

Basic research is carried out for the sake of increasing knowledge, such as how and why a specific reaction occurs and what the properties of a substance are. Chance discoveries can be the result of basic research. The properties of Teflon™, for example, were first discovered by accident. A researcher named Roy Plunkett was puzzled by the fact that a gas cylinder used for an experiment appeared to be empty even though the measured mass of the cylinder clearly indicated there was something inside. Plunkett cut the cylinder open and found a white solid. Through basic research, Plunkett's research team determined the nonstick properties, molecular structure, and chemical composition of the new material.

Applied Research

Applied research is generally carried out to solve a problem. For example, when certain refrigerants escape into the upper atmosphere, they damage the ozone layer, which helps block harmful ultraviolet rays from reaching the surface of Earth. In response to concerns that this atmospheric damage could pose health problems, chemists have developed new refrigerants. In applied research, researchers are driven not by curiosity or a desire to know but by a desire to solve a specific problem.

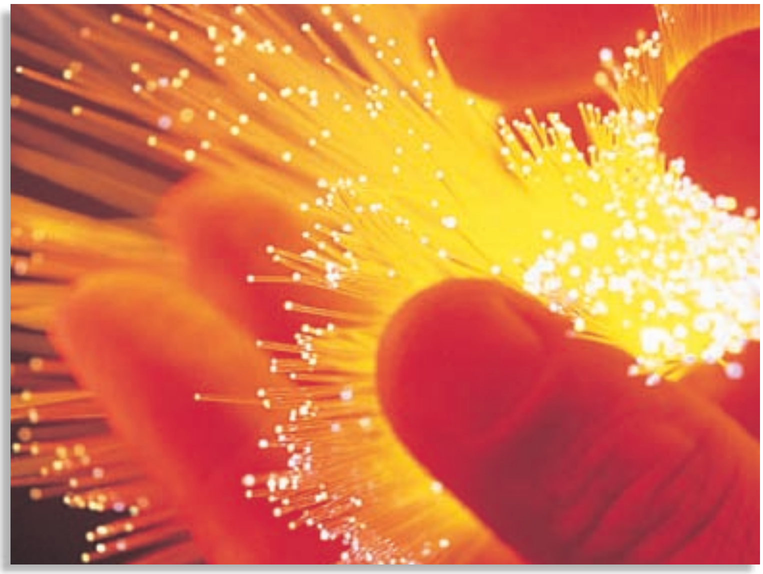


FIGURE 2 The chemical structure of the material in an optical fiber gives it the property of total internal reflection. This property, which allows these fibers to carry light, was discovered through basic and applied research. The use of this property to build networks by sending data on light pulses is the technological development of fiber optics.

Technological Development

Technological development typically involves the production and use of products that improve our quality of life. Examples include computers, catalytic converters for cars, and biodegradable materials.

Technological applications often lag far behind the discoveries that are eventually used in technologies. For example, nonstick cookware, a technological application, was developed well after the accidental discovery of Teflon. When it was later discovered that the Teflon coating on cookware often peeled off, a new challenge arose. Using applied research, scientists were then able to improve the bond between the Teflon and the metal surface of the cookware so that it did not peel.

Basic research, applied research, and technological development often overlap. Discoveries made in basic research may lead to applications that can result in new technologies. For example, knowledge of crystals and light that was gained from basic research was used to develop lasers. It was then discovered that pulses of light from lasers can be sent through optical fibers. Today, telephone messages and cable television signals are carried quickly over long distances using fiber optics.

SECTION REVIEW

1. Define *chemistry*.
2. Name six branches of study in chemistry.
3. Compare and contrast basic research, applied research, and technological development.

Critical Thinking

4. **INFERRING RELATIONSHIPS** Scientific and technological advances are constantly changing how people live and work. Discuss a change that you have observed in your lifetime and that has made life easier or more enjoyable for you.

SECTION 2

Matter and Its Properties

OBJECTIVES

- Distinguish between the physical properties and chemical properties of matter.
- Classify changes of matter as physical or chemical.
- Explain the gas, liquid, and solid states in terms of particles.
- Explain how the law of conservation of energy applies to changes of matter.
- Distinguish between a mixture and a pure substance.

Look around you. You can see a variety of objects—books, desks, chairs, and perhaps trees or buildings outside. All those things are made up of matter, but exactly what is matter? What characteristics, or properties, make matter what it is? In this section, you will learn the answers to these questions.

Explaining what matter is involves finding properties that all matter has in common. That may seem difficult, given that matter takes so many different forms. For the moment, just consider one example of matter—a rock. The first thing you might notice is that the rock takes up space. In other words, it has *volume*. Volume is the amount of three-dimensional space an object occupies. All matter has volume. All matter also has a property called mass. **Mass** *is a measure of the amount of matter*. Mass is the measurement you make using a balance. **Matter** can thus be defined as *anything that has mass and takes up space*. These two properties are the general properties of all matter.

Basic Building Blocks of Matter

Matter comes in many forms. The fundamental building blocks of matter are atoms and molecules. These particles make up elements and compounds. *An* **atom** *is the smallest unit of an element that maintains the chemical identity of that element. An* **element** *is a pure substance that cannot be broken down into simpler, stable substances and is made of one type of atom.* Carbon is an element and contains one kind of atom.

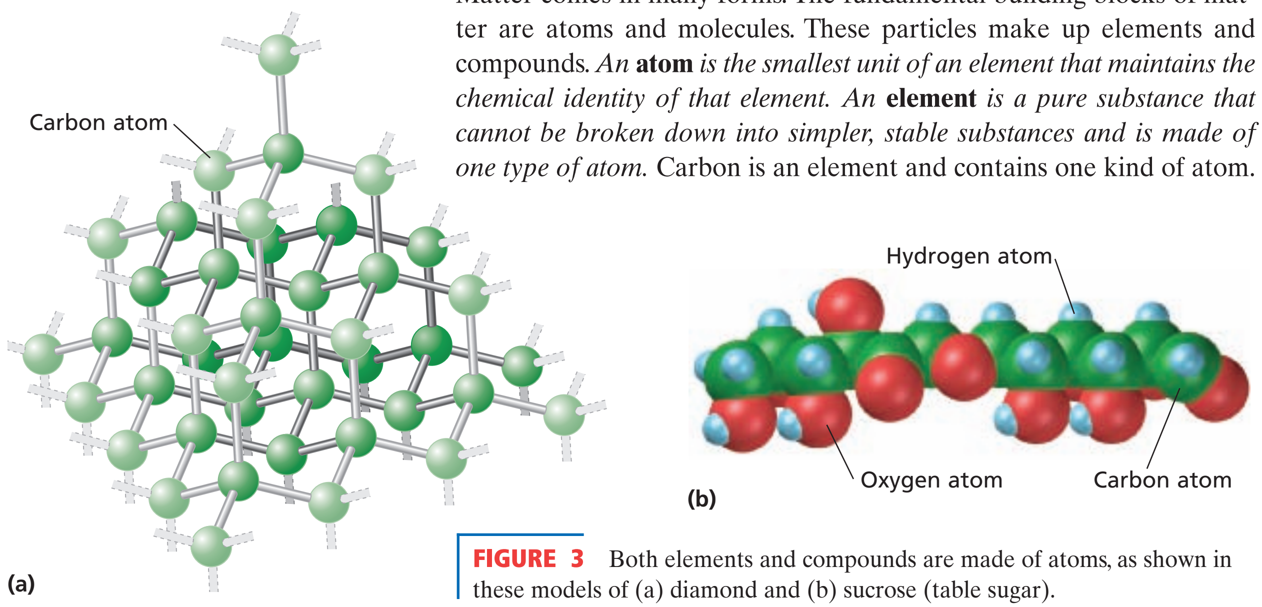


FIGURE 3 Both elements and compounds are made of atoms, as shown in these models of (a) diamond and (b) sucrose (table sugar).

A **compound** *is a substance that can be broken down into simple stable substances. Each compound is made from the atoms of two or more elements that are chemically bonded.* Water is an example of a compound. It is made of two elements, hydrogen and oxygen. The atoms of hydrogen and oxygen are chemically bonded to form a water molecule. You will learn more about the particles that make up compounds when you study chemical bonding in Chapter 6. For now, you can think of a *molecule* as the smallest unit of an element or compound that retains all of the properties of that element or compound.

Properties and Changes in Matter

Every substance, whether it is an element or a compound, has characteristic properties. Chemists use properties to distinguish between substances and to separate them. Most chemical investigations are related to or depend on the properties of substances.

A property may be a characteristic that defines an entire set of substances. That property can be used to classify an unknown substance as a member of that group. For example, many elements are classified as metals. The distinguishing property of metals is that they conduct electricity well. Therefore, if an unknown element is tested and found to conduct electricity well, it is a metal.

Properties can help reveal the identity of an unknown substance. However, conclusive identification usually cannot be made based on only one property. Comparisons of several properties can be used together to establish the identity of an unknown. Properties are either intensive or extensive. **Extensive properties** *depend on the amount of matter that is present*. Such properties include volume, mass, and the amount of energy in a substance. In contrast, **intensive properties** *do not depend on the amount of matter present*. Such properties include the melting point, boiling point, density, and ability to conduct electricity and to transfer energy as heat. Intensive properties are the same for a given substance regardless of how much of the substance is present. Properties can also be grouped into two general types: physical properties and chemical properties.

Physical Properties and Physical Changes

A **physical property** *is a characteristic that can be observed or measured without changing the identity of the substance*. Physical properties describe the substance itself, rather than describing how it can change into other substances. Examples of physical properties are melting point and boiling point. Those points are, respectively, the temperature at which a substance melts from solid to liquid and the temperature at which it boils from liquid to gas. For example, water melts from ice to liquid at 0°C (273 K or 32°F). Liquid water boils to vapor at 100°C (373 K or 212°F).

A change in a substance that does not involve a change in the identity of the substance is called a **physical change.** Examples of physical

FIGURE 4 Water boils at 100°C no matter how much water is in the container. Boiling point is an intensive property.

changes include grinding, cutting, melting, and boiling a material. These types of changes do not change the identity of the substance present.

Melting and boiling are part of an important class of physical changes called changes of state. As the name suggests, *a* **change of state** *is a physical change of a substance from one state to another*. The three common states of matter are solid, liquid, and gas.

Matter in the **solid** *state has definite volume and definite shape.* For example, a piece of quartz or coal keeps its size and its shape, regardless of the container it is in. Solids have this characteristic because the particles in them are packed together in relatively fixed positions. The particles are held close together by the strong attractive forces between them, and only vibrate about fixed points.

Matter in the **liquid** *state has a definite volume but an indefinite shape;* a liquid assumes the shape of its container. For example, a given quantity of liquid water takes up a definite amount of space, but the water takes the shape of its container. Liquids have this characteristic because the particles in them are close together but can move past one another. The particles in a liquid move more rapidly than those in a solid. This causes them to overcome temporarily the strong attractive forces between them, allowing the liquid to flow.

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Matter in the **gas** *state has neither definite volume nor definite shape.* For example, a given quantity of helium expands to fill any size container and takes the shape of the container. All gases have this characteristic because they are composed of particles that move very rapidly and are at great distances from one another compared with the particles of liquids and solids. At these great distances, the attractive forces between gas particles have less of an effect than they do at the small distances between particles of liquids and solids.

An important fourth state of matter is **plasma.** Plasma is a *high-temperature physical state of matter in which atoms lose most of their electrons, particles that make up atoms.* Plasma is found in fluorescent bulbs.

Melting, the change from solid to liquid, is an example of a change of state. Boiling is a change of state from liquid to gas. Freezing, the opposite of melting, is the change from a liquid to a solid. A change of state does not affect the identity of the substance. For example, when ice melts to liquid water or when liquid water boils to form water vapor, the same substance, water, is still present, as shown in **Figure 6.** The water has simply changed state, but it has not turned into a different compound. Only the distances and interactions between the particles that make up water have changed.

FIGURE 5 Because it possesses certain chemical properties, a test strip containing Benedict's solution is used to test for the presence of sugar in urine. The test strip is dipped into the sample. The test strip is then matched to a color scale to determine the sugar level in the urine.

Chemical Properties and Chemical Changes

Physical properties can be observed without changing the identity of the substance, but properties of the second type—chemical properties—cannot. *A* **chemical property** *relates to a substance's ability to undergo changes that transform it into different substances.* Chemical properties are easiest to see when

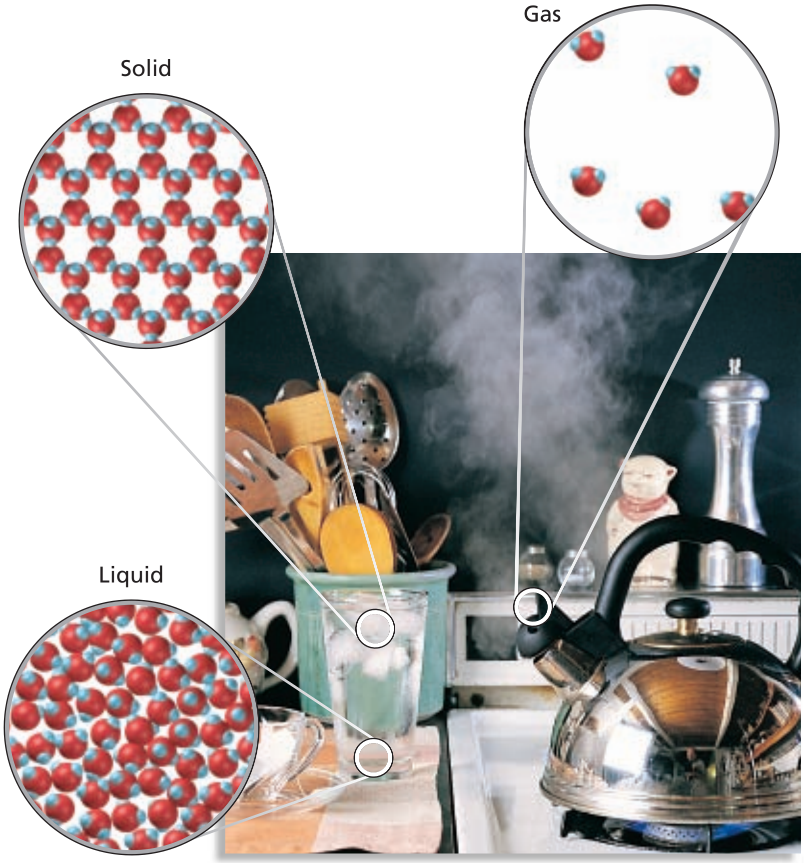

FIGURE 6 Models for water in three states. The molecules are close together in the solid and liquid states but far apart in the gas state. The molecules in the solid state are relatively fixed in position, but those in the liquid and gas states can flow around each other.

substances react to form new substances. For example, the ability of charcoal (carbon) to burn in air is a chemical property. When charcoal burns, it combines with oxygen in air to become a new substance, carbon dioxide gas. After the chemical change, the amounts of the original substances, carbon and oxygen, are less than before. A different substance with different properties has been formed. Other examples of chemical properties include the ability of iron to rust by combining with oxygen in air and the ability of silver to tarnish by combining with sulfur.

A change in which one or more substances are converted into different substances is called a **chemical change** *or* **chemical reaction.** *The substances that react in a chemical change are called the* **reactants.** *The substances that are formed by the chemical change are called the* **products.** In the case of burning charcoal, carbon and oxygen are the reactants in a combustion, or burning, reaction. Carbon dioxide is the product. The chemical change can be described as follows:

Carbon plus oxygen yields (or forms) carbon dioxide.

Arrows and plus signs can be substituted for the words *yields* and *plus,* respectively:

$$\text{carbon} + \text{oxygen} \longrightarrow \text{carbon dioxide}$$

extension

Historical Chemistry

Go to **go.hrw.com** for a full-length article on the chemical reactions of noble gases.

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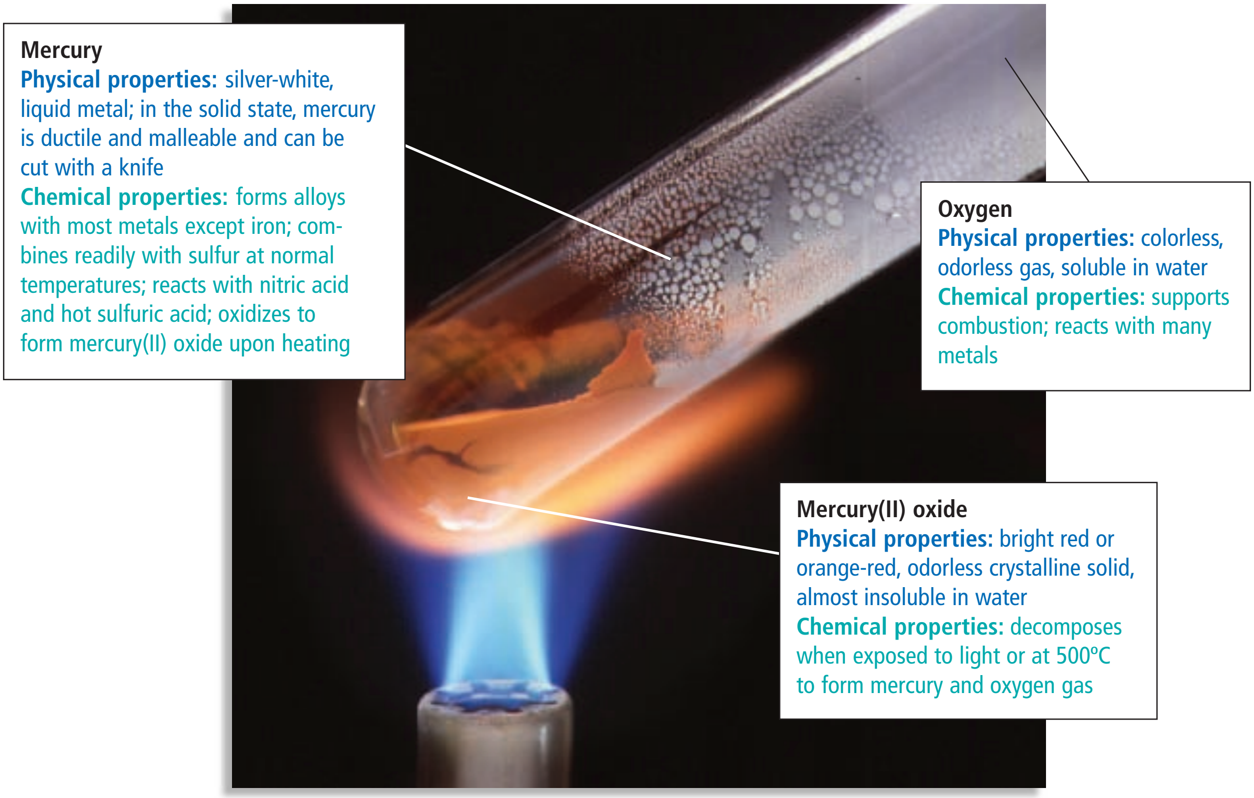


FIGURE 7 When mercury(II) oxide is heated, it decomposes to form oxygen gas and mercury (which can be seen on the side of the test tube). Decomposition is a chemical change that can be observed by comparing the properties of mercury(II) oxide, mercury, and oxygen.

The decomposition of the mercury compound shown in **Figure 7** can be expressed as follows:

$$\text{mercury(II) oxide} \longrightarrow \text{mercury} + \text{oxygen}$$

Chemical changes and reactions, such as combustion and decomposition, form products whose properties differ greatly from those of the reactants. However, chemical changes do not affect the total amount of matter present before and after a reaction. The amount of matter, and therefore the total mass, remains the same.

Energy and Changes in Matter

When physical or chemical changes occur, energy is always involved. The energy can take several different forms, such as heat or light. Sometimes heat provides enough energy to cause a physical change, as in the melting of ice, and sometimes heat provides enough energy to cause a chemical change, as in the decomposition of water vapor to form oxygen gas and hydrogen gas. But the boundary between physical and chemical changes isn't always so clear. For example, while most chemists would consider the dissolving of sucrose in water to be a physical change, many chemists would consider the dissolving of table salt in water to be a chemical change. As you learn more about the structure of matter, you will better understand why the boundaries between chemical and physical changes can be confusing.

Accounting for all the energy present before and after a change is not a simple process. But scientists who have done such experimentation are confident that the total amount of energy remains the same. Although energy can be absorbed or released in a change, it is not destroyed or created. It simply assumes a different form. This is the law of conservation of energy.

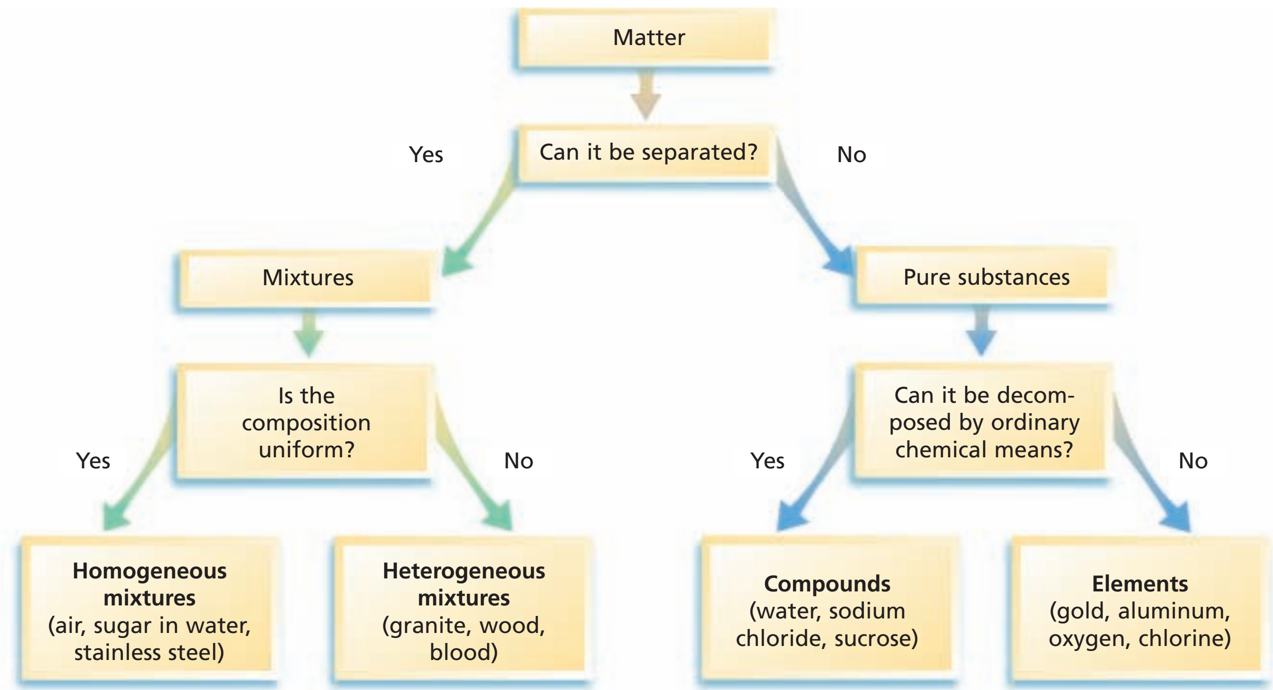
Classification of Matter

Matter exists in an enormous variety of forms. Any sample of matter, however, can be classified either as a pure substance or as a mixture. The composition of a pure substance is the same throughout and does not vary from sample to sample. A pure substance can be an element or a compound. Mixtures, in contrast, contain more than one substance. They can vary in composition and properties from sample to sample and sometimes from one part of a sample to another part of the same sample. All matter, whether it is a pure substance or a mixture, can be classified in terms of uniformity of composition and properties of a given sample. **Figure 8** illustrates the overall classification of matter into elements, compounds, and mixtures.

Mixtures

You deal with mixtures every day. Nearly every object around you, including most things you eat and drink and even the air you breathe, is a mixture. *A* **mixture** *is a blend of two or more kinds of matter, each*

FIGURE 8 This classification scheme for matter shows the relationships among mixtures, compounds, and elements.



(a) (b) (c)

FIGURE 9 (a) Barium chromate can be separated from the solution in the beaker using filtration. (b) A centrifuge can be used to separate certain solid components. The centrifuge spins rapidly, which causes the solids to settle to the bottom of the test tube. (c) The components of an ink can be separated using paper chromatography.

of which retains its own identity and properties. The parts, or components, of a mixture are simply mixed together physically and can usually be separated. As a result, the properties of a mixture are a combination of the properties of its components. Because mixtures can contain various amounts of different substances, a mixture's composition must be specified. This is often done in terms of percentage by mass or by volume. For example, a mixture might be 5% sodium chloride and 95% water by mass.

Some mixtures are *uniform in composition;* that is, they are said to be **homogeneous.** They have the same proportion of components throughout. *Homogeneous mixtures are also called* **solutions.** A salt-water solution is an example of such a mixture. Other mixtures are *not uniform throughout;* that is, they are **heterogeneous.** For example, in a mixture of clay and water, heavier clay particles concentrate near the bottom of the container.

Some mixtures can be separated by filtration or vaporized to separate the different components. Filtration can be used to separate a mixture of solid barium chromate from the other substances, as shown in the beaker in **Figure 9a.** The yellow barium compound is trapped by the filter paper, but the solution passes through. If the solid in a liquid-solid mixture settles to the bottom of the container, the liquid can be carefully poured off (decanted). A centrifuge **(Figure 9b)** can be used to separate some solid-liquid mixtures, such as those in blood. Another technique, called paper chromatography, can be used to separate mixtures of dyes or pigments because the different substances move at different rates on the paper **(Figure 9c).**

Pure Substances

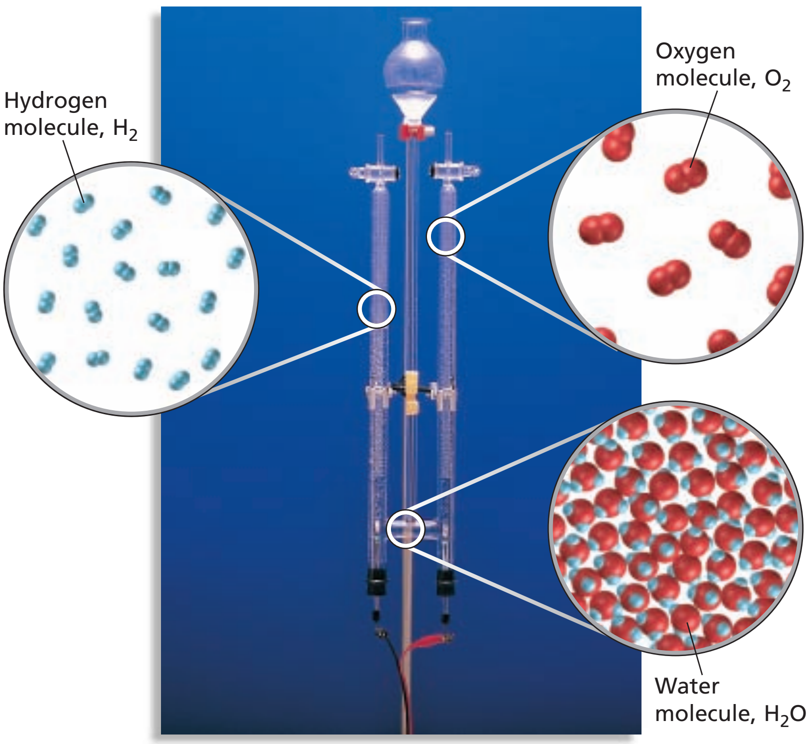
Any sample of a pure substance is homogeneous. *A* **pure substance** *has a fixed composition and differs from a mixture in the following ways:*

1. *Every sample of a given pure substance has exactly the same characteristic properties.* All samples of a pure substance have the same characteristic physical and chemical properties. These properties are so specific that they can be used to identify the substance. In contrast, the properties of a mixture depend on the relative amounts of the mixture's components.
2. *Every sample of a given pure substance has exactly the same composition.* Unlike mixtures, all samples of a pure substance have the same makeup. For example, pure water is always 11.2% hydrogen and 88.8% oxygen by mass.

Pure substances are either compounds or elements. A compound can be decomposed, or broken down, into two or more simpler compounds or elements by a chemical change. Water is a compound made of hydrogen and oxygen chemically bonded to form a single substance. Water can be broken down into hydrogen and oxygen through a chemical reaction called electrolysis, as shown in **Figure 10a.**

Sucrose is made of carbon, hydrogen, and oxygen. Sucrose breaks down to form the other substances shown in **Figure 10b.** Under intense heating, sucrose breaks down to produce carbon and water.

FIGURE 10 (a) Passing an electric current through water causes the compound to break down into the elements hydrogen and oxygen, which differ in composition from water. (b) When sucrose is heated, it caramelizes. When it is heated to a high enough temperature, it breaks down completely into carbon and water.



(a)

(b)

FIGURE 11 The labeling on a reagent bottle lists the grade of the reagent and the percentages of impurities for that grade. What grade is this chemical?

TABLE 1 *Some Grades of Chemical Purity*

Increasing purity (↑)	
	Primary standard reagents
	ACS (American Chemical Society–specified reagents)
	USP (United States Pharmacopoeia standards)
	CP (chemically pure; purer than technical grade)
	NF (National Formulary specifications)
	FCC (Food Chemical Code specifications)
	Technical (industrial chemicals)

Laboratory Chemicals and Purity

The chemicals in laboratories are generally treated as if they are pure. However, all chemicals have some impurities. Chemical grades of purity are listed in **Table 1.** The purity ranking of the grades can vary when agencies differ in their standards. For some chemicals, the USP grade may specify higher purity than the CP grade. For other chemicals, the opposite may be true. However, the primary standard reagent grade is always purer than the technical grade for the same chemical. Chemists need to be aware of the kinds of impurities in a reagent because these impurities could affect the results of a reaction. For example, the chemical label shown in **Figure 11** shows the impurities for that grade. The chemical manufacturer must ensure that the standards set for that reagent by the American Chemical Society are met.

SECTION REVIEW

1. **a.** What is the main difference between physical properties and chemical properties?
   b. Give an example of each.
2. Classify each of the following as either a physical change or a chemical change.
   a. tearing a sheet of paper
   b. melting a piece of wax
   c. burning a log
3. How do you decide whether a sample of matter is a solid, liquid, or gas?
4. Contrast mixtures with pure substances.

Critical Thinking

5. **ANALYZING INFORMATION** Compare the composition of sucrose purified from sugar cane with the composition of sucrose purified from sugar beets. Explain your answer.

SECTION 3

Elements

OBJECTIVES

- Use a periodic table to name elements, given their symbols.
- Use a periodic table to write the symbols of elements, given their names.
- Describe the arrangement of the periodic table.
- List the characteristics that distinguish metals, nonmetals, and metalloids.

As you have read, elements are pure substances that cannot be decomposed by chemical changes. The elements serve as the building blocks of matter. Each element has characteristic properties. The elements are organized into groups based on similar chemical properties. This organization of elements is the *periodic table,* which is shown in **Figure 12** on the next page.

Introduction to the Periodic Table

Each small square of the periodic table shows the symbol for the element and the atomic number. For example, the first square, at the upper left, represents element 1, hydrogen, which has the symbol H. As you look through the table, you will see many familiar elements, including iron, sodium, neon, silver, copper, aluminum, sulfur, and lead. You can often relate the symbols to the English names of the elements. Some symbols are derived from the element's older name, which was often in Latin. Still others come from German. For example, wolfram comes from the German name for tungsten. **Table 2** lists some elements and their older names.

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TABLE 2 ***Elements with Symbols Based on Older Names***

Modern name	Symbol	Older name
Antimony	Sb	stibium
Copper	Cu	cuprum
Gold	Au	aurum
Iron	Fe	ferrum
Lead	Pb	plumbum
Mercury	Hg	hydrargyrum
Potassium	K	kalium
Silver	Ag	argentum
Sodium	Na	natrium
Tin	Sn	stannum
Tungsten	W	wolfram

Periodic Table

- Metals
- Metalloids
- Nonmetals

Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
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 La	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 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg							

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
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 12 The periodic table of elements. The names of the elements can be found on Table A-6 in the appendix.

The vertical columns of the periodic table are called **groups,** or **families.** Notice that they are numbered from 1 to 18 from left to right. Each group contains elements with similar chemical properties. For example, the elements in Group 2 are beryllium, magnesium, calcium, strontium, barium, and radium. All of these elements are reactive metals with similar abilities to bond to other kinds of atoms. The two major categories of elements are metals and nonmetals. Metalloids have properties intermediate between those of metals and nonmetals.

The horizontal rows of elements in the periodic table are called **periods.** Physical and chemical properties change somewhat regularly across a period. Elements that are close to each other in the same period tend to be more similar than elements that are farther apart. For example, in Period 2, the elements lithium and beryllium, in Groups 1 and 2, respectively, are somewhat similar in properties. However, their properties are very different from the properties of fluorine, the Period-2 element in Group 17.

The two sets of elements placed below the periodic table make up what are called the lanthanide series and the actinide series. These metallic elements fit into the table just after elements 57 and 89. They are placed below the table to keep the table from being too wide.

There is a section in the back of this book called the *Elements Handbook* which covers some elements in greater detail. You will use information from the handbook to complete the questions in the Using the Handbook sections in the chapter reviews.

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Chemistry in Action

Superconductors

Any metal becomes a better conductor of electrical energy as its temperature decreases. In 1911, scientists discovered that when mercury is cooled to about –269°C, it loses all resistance and becomes a superconductor. Scientists have long tried to find a material that would superconduct at temperatures above –196°C, the boiling point of liquid nitrogen. In 1987, scientists discovered ceramic materials that became superconductors when cooled only to –183°C. These "high-temperature" superconductors are used to build very powerful electromagnets. Ceramic electromagnets are used in medical magnetic resonance imaging (MRI) machines and in high-efficiency electric motors and generators.

Types of Elements

The periodic table is broadly divided into two main sections: metals and nonmetals. As you can see in **Figure 12,** the metals are at the left and in the center of the table. The nonmetals are toward the right. Some elements, such as boron and silicon, show characteristics of both metals and nonmetals.

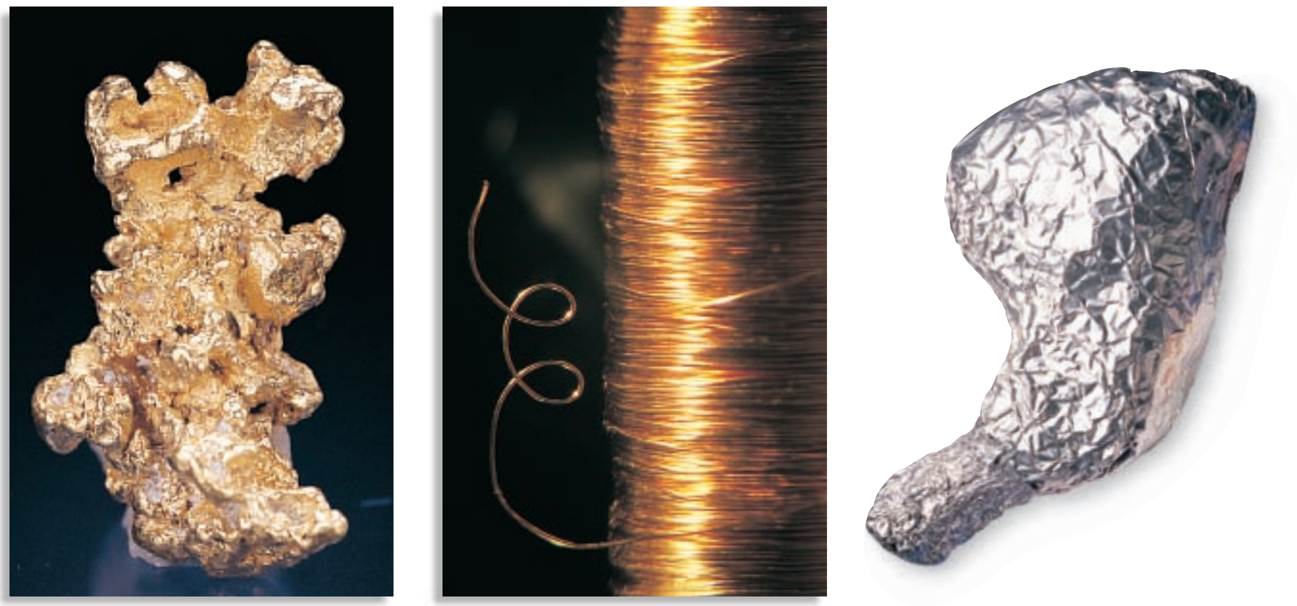
Metals

Some of the properties of metals may be familiar to you. For example, you can recognize metals by their shininess, or metallic luster. Perhaps the most important characteristic property of metals is the ease with which they conduct electricity and transfer energy. Thus, *a* **metal** *is an element that is a good electrical conductor and a good heat conductor*.

At room temperature, most metals are solids. Most metals also have the property of *malleability,* that is, they can be hammered or rolled into thin sheets. Metals also tend to be *ductile,* which means that they can be drawn into a fine wire. Metals behave this way because they have high *tensile strength,* the ability to resist breaking when pulled.

Although all metals conduct electricity well, metals also have very diverse properties. Mercury is a liquid at room temperature, whereas tungsten has the highest melting point of any element. The metals in Group 1 are so soft that they can be cut with a knife, yet others, such as chromium, are very hard. Some metals, such as manganese and bismuth, are very brittle, yet others, such as iron and copper, are very malleable and ductile. Most metals have a silvery or grayish white *luster*. Two exceptions are gold and copper, which are yellow and reddish brown, respectively. **Figure 13** shows examples of metals.

FIGURE 13 (a) Gold has a low reactivity, which is why it may be found in nature in relatively pure form. (b) Copper is used in wiring because it is ductile and conducts electrical energy (c) Aluminum is malleable. It can be rolled into foil that is used for wrapping food.



(a) (b) (c)

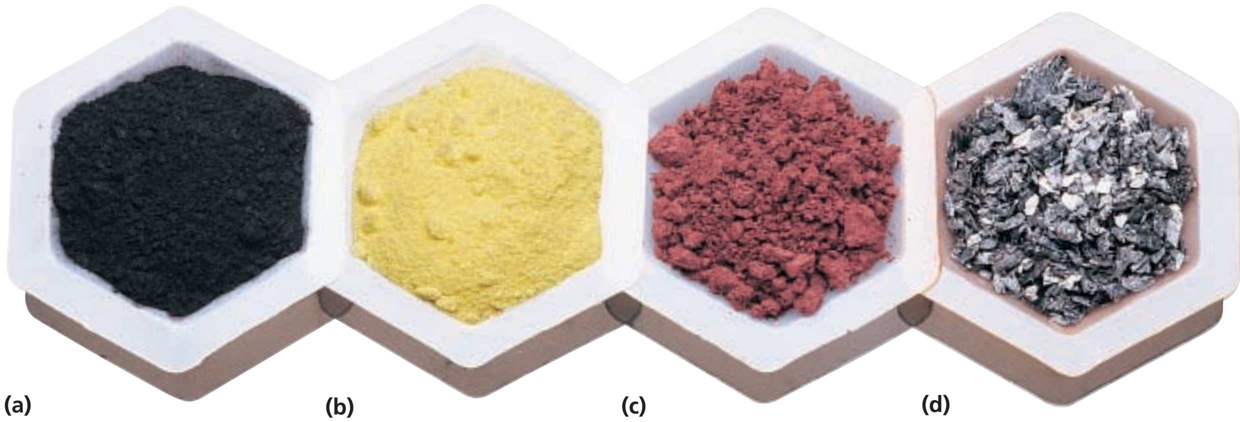


FIGURE 14 Various nonmetallic elements: (a) carbon, (b) sulfur, (c) phosphorus, and (d) iodine

Copper: A Typical Metal

Copper has a characteristic reddish color and a metallic luster. It is found naturally in minerals such as chalcopyrite and malachite. Pure copper melts at 1083°C and boils at 2567°C. It can be readily drawn into fine wire, pressed into thin sheets, and formed into tubing. Copper conducts electricity with little loss of energy.

Copper remains unchanged in pure, dry air at room temperature. When heated, it reacts with oxygen in air. It also reacts with sulfur and the elements in Group 17 of the periodic table. The green coating on a piece of weathered copper comes from the reaction of copper with oxygen, carbon dioxide, and sulfur compounds. Copper is an essential mineral in the human diet.

Nonmetals

Many nonmetals are gases at room temperature. These include nitrogen, oxygen, fluorine, and chlorine. One nonmetal, bromine, is a liquid. The solid nonmetals include carbon, phosphorus, selenium, sulfur, and iodine. These solids tend to be brittle rather than malleable and ductile. Some nonmetals are illustrated in **Figure 14.**

Low conductivity can be used to define nonmetals. *A* **nonmetal** *is an element that is a poor conductor of heat and electricity.* If you look at **Figure 12,** you will see that there are fewer nonmetals than metals.

Phosphorus: A Typical Nonmetal

Phosphorus is one of five solid nonmetals. Pure phosphorus is known in two common forms. Red phosphorus is a dark red powder that melts at 597°C. White phosphorus is a waxy solid that melts at 44°C. Because it ignites in air at room temperature, white phosphorus is stored under water.

Phosphorus is too reactive to exist in pure form in nature. It is present in huge quantities in phosphate rock, where it is combined with oxygen and calcium. All living things contain phosphorus.

FIGURE 15 Selenium is a nonmetal, though it looks metallic.

Metalloids

As you look from left to right on the periodic table, you can see that the metalloids are found between the metals and the nonmetals. *A* **metalloid**

FIGURE 16 Some noble gases are used to make lighted signs of various colors.

is an element that has some characteristics of metals and some characteristics of nonmetals. All metalloids are solids at room temperature. They tend to be less malleable than metals but not as brittle as nonmetals. Some metalloids, such as antimony, have a somewhat metallic luster.

Metalloids tend to be semiconductors of electricity. That is, their ability to conduct electricity is intermediate between that of metals and that of nonmetals. Metalloids are used in the solid state circuitry found in desktop computers, digital watches, televisions, and radios.

Noble Gases

The elements in Group 18 of the periodic table are the noble gases. These elements are generally unreactive. In fact, it was not until 1962 that the first noble gas compound, xenon hexafluoroplatinate, was prepared. Low reactivity makes the noble gases very different from the other families of elements. Group 18 elements are gases at room temperature. Neon, argon, krypton, and xenon are all used in lighting. Helium is used in party balloons and weather balloons because it is less dense than air.

SECTION REVIEW

1. Use the periodic table to write the names for the following elements: O, S, Cu, Ag.
2. Use the periodic table to write the symbols for the following elements: iron, nitrogen, calcium, mercury.
3. Which elements are most likely to undergo the same kinds of reactions, those in a group or those in a period?
4. Describe the main differences between metals, nonmetals, and metalloids.

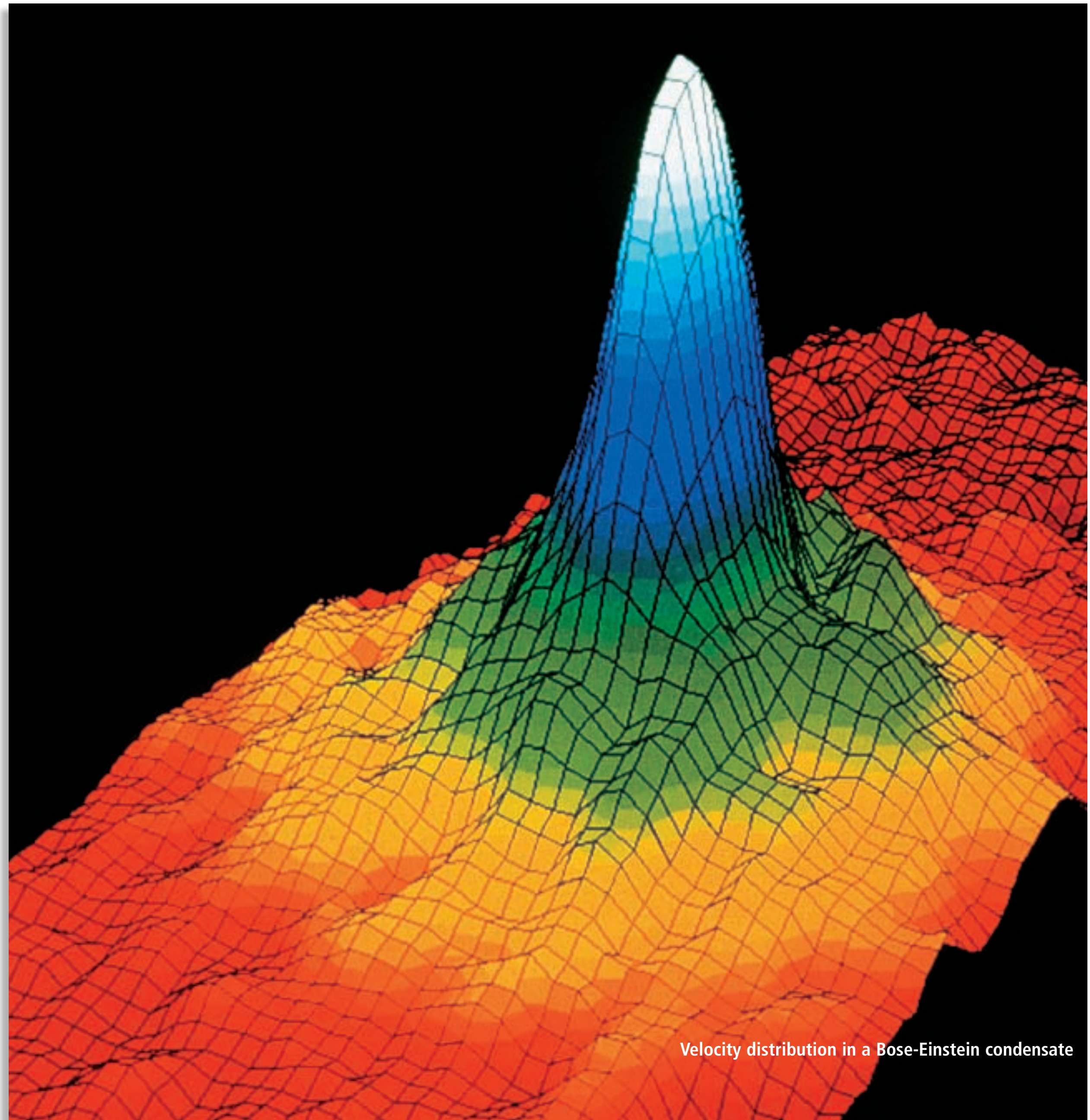
Critical Thinking

5. **INFERRING CONCLUSIONS** If you find an element in nature in its pure elemental state, what can you infer about the element's chemical reactivity? How can you tell whether that element is a metal or a nonmetal?

CHAPTER 2

Measurements and Calculations

Quantitative measurements are fundamental to chemistry.



Velocity distribution in a Bose-Einstein condensate

Scientific Method

SECTION 1

OBJECTIVES

- Describe the purpose of the scientific method.
- Distinguish between qualitative and quantitative observations.
- Describe the differences between hypotheses, theories, and models.

Sometimes progress in science comes about through accidental discoveries. Most scientific advances, however, result from carefully planned investigations. The process researchers use to carry out their investigations is often called the scientific method. *The* **scientific method** *is a logical approach to solving problems by observing and collecting data, formulating hypotheses, testing hypotheses, and formulating theories that are supported by data.*

Observing and Collecting Data

Observing is the use of the senses to obtain information. Observation often involves making measurements and collecting data. The data may be descriptive (qualitative) or numerical (quantitative) in nature. Numerical information, such as the fact that a sample of copper ore has a mass of 25.7 grams, is *quantitative*. Non-numerical information, such as the fact that the sky is blue, is *qualitative*.

Experimenting involves carrying out a procedure under controlled conditions to make observations and collect data. To learn more about matter, chemists study systems. *A* **system** *is a specific portion of matter in a given region of space that has been selected for study during an experiment or observation.* When you observe a reaction in a test tube, the test tube and its contents form a system.

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FIGURE 1 These students have designed an experiment to determine how to get the largest volume of popped corn from a fixed number of kernels. They think that the volume is likely to increase as the moisture in the kernels increases. Their experiment will involve soaking some kernels in water and observing whether the volume of the popped corn is greater than that of corn popped from kernels that have not been soaked.

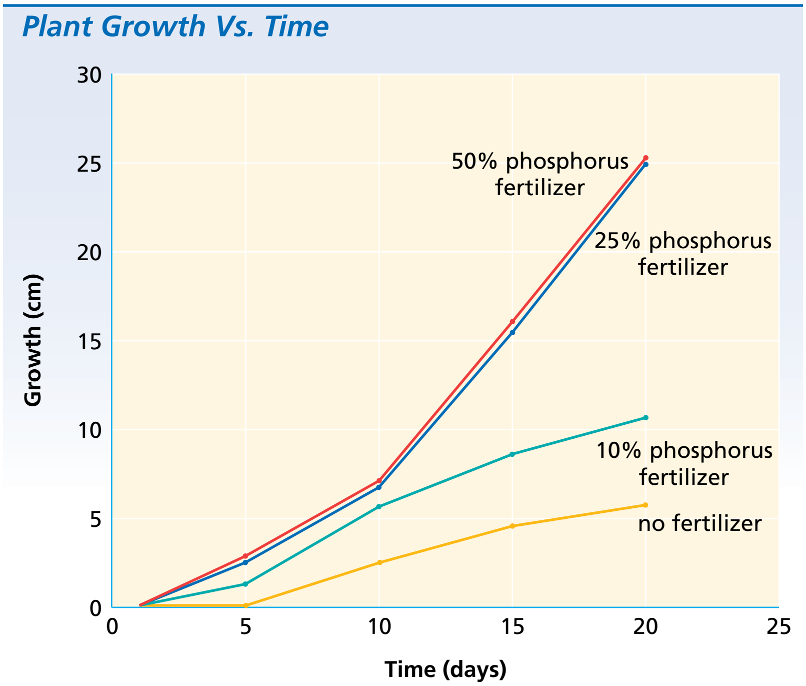


FIGURE 2 A graph of data can show relationships between two variables. In this case the graph shows data collected during an experiment to determine the effect of phosphorus fertilizer compounds on plant growth. The following is one possible hypothesis: *If* phosphorus stimulates corn-plant growth, *then* corn plants treated with a soluble phosphorus compound should grow faster, under the same conditions, than corn plants that are not treated.

Formulating Hypotheses

As scientists examine and compare the data from their own experiments, they attempt to find relationships and patterns—in other words, they make generalizations based on the data. Generalizations are statements that apply to a range of information. To make generalizations, data are sometimes organized in tables and analyzed using statistics or other mathematical techniques, often with the aid of graphs and a computer.

Scientists use generalizations about the data to formulate a **hypothesis,** *or testable statement*. The hypothesis serves as a basis for making predictions and for carrying out further experiments. Hypotheses are often drafted as "if-then" statements. The "then" part of the hypothesis is a prediction that is the basis for testing by experiment. **Figure 2** shows data collected to test a hypothesis.

Testing Hypotheses

Testing a hypothesis requires experimentation that provides data to support or refute a hypothesis or theory. During testing, the experimental conditions that remain constant are called *controls,* and any condition that changes is called a *variable.* Any change observed is usually due to the effects of the variable. If testing reveals that the predictions were not correct, the hypothesis on which the predictions were based must be discarded or modified.

STAGES IN THE SCIENTIFIC METHOD

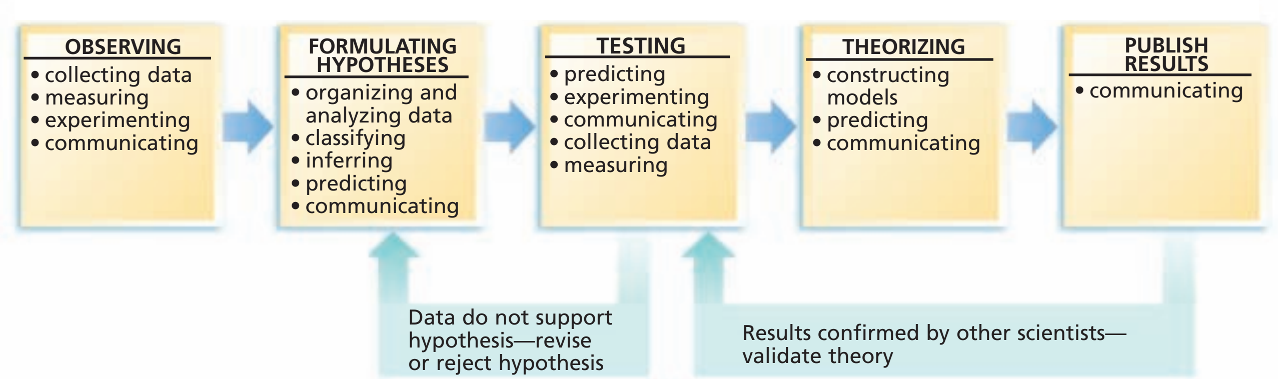


FIGURE 3 The scientific method is not a single, fixed process. Scientists may repeat steps many times before there is sufficient evidence to formulate a theory. You can see that each stage represents a number of different activities.

Theorizing

When the data from experiments show that the predictions of the hypothesis are successful, scientists typically try to explain the phenomena they are studying by constructing a model. *A* **model** *in science is more than a physical object; it is often an explanation of how phenomena occur and how data or events are related*. Models may be visual, verbal, or mathematical. One important model in chemistry is the atomic model of matter, which states that matter is composed of tiny particles called atoms.

If a model successfully explains many phenomena, it may become part of a theory. The atomic model is a part of the atomic theory, which you will study in Chapter 3. *A* **theory** *is a broad generalization that explains a body of facts or phenomena*. Theories are considered successful if they can predict the results of many new experiments. Examples of the important theories you will study in chemistry are kinetic-molecular theory and collision theory. **Figure 3** shows where theory fits in the scheme of the scientific method.

SECTION REVIEW

1. What is the scientific method?
2. Which of the following are quantitative?
   a. the liquid floats on water
   b. the metal is malleable
   c. the liquid has a temperature of 55.6°C
3. How do hypotheses and theories differ?
4. How are models related to theories and hypotheses?

Critical Thinking

5. **INTERPRETING CONCEPTS** Suppose you had to test how well two types of soap work. Describe your experiment by using the terms *control* and *variable.*

Chemistry in Action



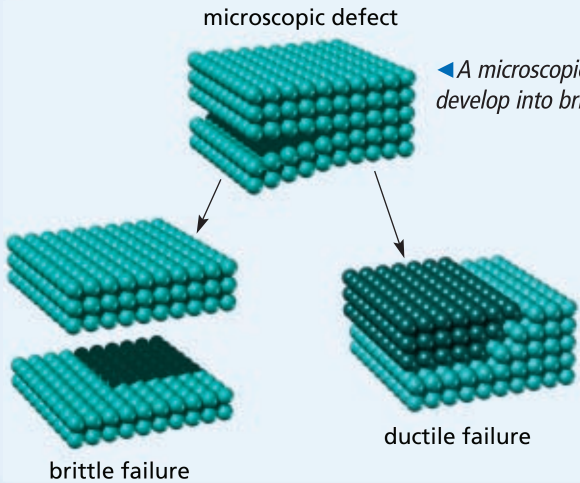
Breaking Up Is Easy To Do

It may seem obvious that chemistry is important in the making of materials, but chemistry is also vital to the study of how materials break. Everyday items have to be made to withstand various types of force and pressure or they cannot be used. For example, scientists and engineers work to ensure that highway bridges do not collapse.

When excessive force is applied to an object, the material that the object is made of will break. The object breaks because the force creates stress on the bonds between the atoms of the material and causes the bonds to break. This creates microscopic cracks in the material. When a material breaks, it is said to have undergone *failure.* Materials typically break in one of two ways: *ductile failure* and *brittle failure.* Both types of failure start with microscopic cracks in the material. However, the way a material eventually breaks depends how its atoms are organized.

Shattering glass undergoes brittle failure. Glass shatters when the bonds between the two layers of atoms that are along the initial crack break. This breakage causes the layers to pull apart, which separates the material into pieces. This type of failure is common in materials that do not have a very orderly arrangement of atoms.

When a car bumper crumples, ductile failure happens. This type of failure tends to happen in materials such as metals, that have a regular, ordered arrangement of atoms. This arrangement of atoms is known as a *crystal structure.* Ductile failure happens when the bonds in the material break across many layers of atoms that are not in the same plane as the original crack. Rather than splitting apart, the layers slip past each other into new positions. The atoms form new chemical bonds, between them and the material stays in one piece; only the shape has changed.



◀*A microscopic crack in a material can develop into brittle or ductile failure.*

In addition to the type of material influencing breakage, the quality of the material also influences breakage. All objects contain microscopic defects, such as bubbles in plastic pieces. A material will tend to undergo failure at its defect sites first. Careful fabrication procedures can minimize, but not completely eliminate, defects in materials.

Even though materials are designed to withstand a certain amount of force, the normal wear and tear that materials experience over their lifetimes creates defects in the material. This process is referred to as *fatigue.* If fatigue were to go undetected, the microscopic cracks that form could then undergo brittle or ductile failure. It would be catastrophic if the materials in certain products, such as airplane parts, failed. To avoid such a failure, people monitor materials that are exposed to constant stress for signs of fatigue. The defects in the metal parts of airplanes can be detected with nondestructive techniques, such as electromagnetic analysis.

Questions

1. Can you name some ways in which metal or plastic parts might obtain defects caused by chemical reactions?
2. Does a ceramic dinner plate undergo brittle or ductile failure when it is dropped and breaks?

Units of Measurement

SECTION 2

OBJECTIVES

- Distinguish between a quantity, a unit, and a measurement standard.
- Name and use SI units for length, mass, time, volume, and density.
- Distinguish between mass and weight.
- Perform density calculations.
- Transform a statement of equality into a conversion factor.

Measurements are quantitative information. A measurement is more than just a number, even in everyday life. Suppose a chef were to write a recipe listing quantities such as 1 salt, 3 sugar, and 2 flour. The cooks could not use the recipe without more information. They would need to know whether the number 3 represented teaspoons, tablespoons, cups, ounces, grams, or some other unit for sugar.

Measurements *represent* quantities. *A* **quantity** *is something that has magnitude, size, or amount.* A quantity is not the same as a measurement. For example, the quantity represented by a teaspoon is volume. The teaspoon is a unit of measurement, while volume is a quantity. A teaspoon is a measurement standard in this country. Units of measurement compare what is to be measured with a previously defined size. Nearly every measurement is a number plus a unit. The choice of unit depends on the quantity being measured.

Many centuries ago, people sometimes marked off distances in the number of foot lengths it took to cover the distance. But this system was unsatisfactory because the number of foot lengths used to express a distance varied with the size of the measurer's foot. Once there was agreement on a standard for foot length, confusion as to the real length was eliminated. It no longer mattered who made the measurement, as long as the standard measuring unit was correctly applied.

SI Measurement

Scientists all over the world have agreed on a single measurement system called *Le Système International d'Unités*, abbreviated **SI.** This system was adopted in 1960 by the General Conference on Weights and Measures. SI now has seven base units, and most other units are derived from these seven. Some non-SI units are still commonly used by chemists and are also used in this book.

SI units are defined in terms of standards of measurement. The standards are objects or natural phenomena that are of constant value, easy to preserve and reproduce, and practical in size. International organizations monitor the defining process. In the United States, the National Institute of Standards and Technology (NIST) plays the main role in maintaining standards and setting style conventions. For example, numbers are written in a form that is agreed upon internationally. The number seventy-five thousand is written 75 000, not 75,000, because the comma is used in other countries to represent a decimal point.

TABLE 1 SI Base Units

Quantity	Quantity symbol	Unit name	Unit abbreviation	Defined standard
Length	l	meter	m	the length of the path traveled by light in a vacuum during a time interval of 1/299 792 458 of a second
Mass	m	kilogram	kg	the unit of mass equal to the mass of the international prototype of the kilogram
Time	t	second	s	the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom
Temperature	T	kelvin	K	the fraction 1/273.16 of the thermodynamic temperature of the triple point of water
Amount of substance	n	mole	mol	the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12
Electric current	I	ampere	A	the constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross section, and placed 1 meter apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per meter of length
Luminous intensity	I_v	candela	cd	the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian

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SI Base Units

The seven SI base units and their standard abbreviated symbols are listed in **Table 1.** All the other SI units can be derived from the fundamental units.

Prefixes added to the names of SI base units are used to represent quantities that are larger or smaller than the base units. **Table 2** lists SI prefixes using units of length as examples. For example, the prefix *centi-*, abbreviated c, represents an exponential factor of 10^{-2}, which equals 1/100. Thus, 1 centimeter, 1 cm, equals 0.01 m, or 1/100 of a meter.

Mass

As you learned in Chapter 1, mass is a measure of the quantity of matter. The SI standard unit for mass is the kilogram. The standard for mass defined in **Table 1** is used to calibrate balances all over the world.

TABLE 2 *SI Prefixes*

Prefix	Unit abbreviation	Exponential factor	Meaning	Example
tera	T	10^{12}	1 000 000 000 000	1 terameter (Tm) = 1×10^{12} m
giga	G	10^{9}	1 000 000 000	1 gigameter (Gm) = 1×10^{9} m
mega	M	10^{6}	1 000 000	1 megameter (Mm) = 1×10^{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
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) = 1×10^{-6} m
nano	n	10^{-9}	1/1 000 000 000	1 nanometer (nm) = 1×10^{-9} m
pico	p	10^{-12}	1/1 000 000 000 000	1 picometer (pm) = 1×10^{-12} m
femto	f	10^{-15}	1/1 000 000 000 000 000	1 femtometer (fm) = 1×10^{-15} m
atto	a	10^{-18}	1/1 000 000 000 000 000 000	1 attometer (am) = 1×10^{-18} m

The mass of a typical textbook is about 1 kg. The gram, g, which is 1/1000 of a kilogram, is more useful for measuring masses of small objects, such as flasks and beakers. For even smaller objects, such as tiny quantities of chemicals, the milligram, mg, is often used. One milligram is 1/1000 of a gram, or 1/1 000 000 of a kilogram.

Mass is often confused with weight because people often express the weight of an object in grams. Mass is determined by comparing the mass of an object with a set of standard masses that are part of the balance. **Weight** *is a measure of the gravitational pull on matter.* Unlike weight, mass does not depend on gravity. Mass is measured on instruments such as a balance, and weight is typically measured on a spring scale. Taking weight measurements involves reading the amount that an object pulls down on a spring. As the force of Earth's gravity on an object increases, the object's weight increases. The weight of an object on the moon is about one-sixth of its weight on Earth.

Length

The SI standard unit for length is the meter. A distance of 1 m is about the width of an average doorway. To express longer distances, the kilometer, km, is used. One kilometer equals 1000 m. Road signs in the United States sometimes show distances in kilometers as well as miles. The kilometer is the unit used to express highway distances in most other countries of the world. To express shorter distances, the centimeter

CROSS-DISCIPLINARY

Some Handy Comparisons of Units

To become comfortable with units in the SI system, try relating some common measurements to your experience.

A meter stick is a little longer than a yardstick. A millimeter is about the diameter of a paper clip wire, and a centimeter is a little more than the width of a paper clip.

One gram is about the mass of a paper clip. A kilogram is about 2.2 pounds (think of two pounds plus one stick of butter). And there are about five milliliters in a teaspoon.

FIGURE 4 The meter is the SI unit of length, but the centimeter is often used to measure smaller distances. What is the length in cm of the rectangular piece of aluminum foil shown?

is often used. From **Table 2,** you can see that one centimeter equals 1/100 of a meter. The width of this book is just over 20 cm.

Derived SI Units

Many SI units are combinations of the quantities shown in **Table 1.** *Combinations of SI base units form* **derived units.** Some derived units are shown in **Table 3.**

Derived units are produced by multiplying or dividing standard units. For example, area, a derived unit, is length times width. If both length and width are expressed in meters, the area unit equals meters times meters, or square meters, abbreviated m^2. The last column of

TABLE 3 ***Derived SI Units***

Quantity	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	$\frac{kg}{m^3}$	$\frac{\text{mass}}{\text{volume}}$
Molar mass	M	kilograms per mole	$\frac{kg}{mol}$	$\frac{\text{mass}}{\text{amount of substance}}$
Molar volume	V_m	cubic meters per mole	$\frac{m^3}{mol}$	$\frac{\text{volume}}{\text{amount of substance}}$
Energy	E	joule	J	force × length

Table 3 shows the combination of fundamental units used to obtain derived units.

Some combination units are given their own names. For example, pressure expressed in base units is the following.

$$kg/m{\bullet}s^2$$

The name *pascal,* Pa, is given to this combination. You will learn more about pressure in Chapter 11. Prefixes can also be added to express derived units. Area can be expressed in cm^2, square centimeters, or mm^2, square millimeters.

FIGURE 5 The speed that registers on a speedometer represents distance traveled per hour and is expressed in the derived units kilometers per hour or miles per hour.

Volume

Volume *is the amount of space occupied by an object.* The derived SI unit of volume is cubic meters, m^3. One cubic meter is equal to the volume of a cube whose edges are 1 m long. Such a large unit is inconvenient for expressing the volume of materials in a chemistry laboratory. Instead, a smaller unit, the cubic centimeter, cm^3, is often used. There are 100 centimeters in a meter, so a cubic meter contains 1 000 000 cm^3.

$$1\ m^3 \times \frac{100\ cm}{1\ m} \times \frac{100\ cm}{1\ m} \times \frac{100\ cm}{1\ m} = 1\ 000\ 000\ cm^3$$

When chemists measure the volumes of liquids and gases, they often use a non-SI unit called the liter. The liter is equivalent to one cubic decimeter. Thus, a liter, L, is also equivalent to 1000 cm^3. Another non-SI unit, the milliliter, mL, is used for smaller volumes. There are 1000 mL in 1 L. Because there are also 1000 cm^3 in a liter, the two units—milliliter and cubic centimeter—are interchangeable.

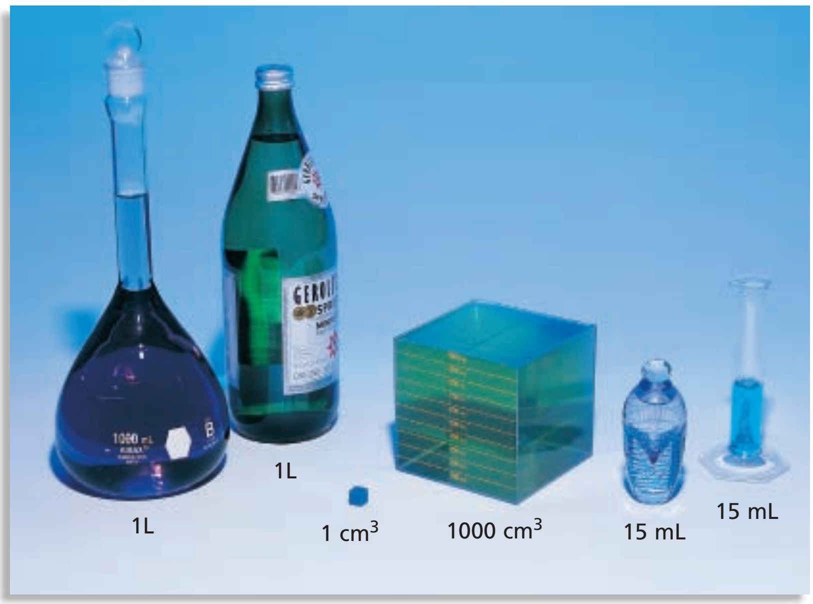

FIGURE 6 The relationships between various volumes are shown here. One liter contains 1000 mL of liquid, and 1 mL is equivalent to 1 cm^3. A small perfume bottle contains about 15 mL of liquid. The volumetric flask (far left) and graduated cylinder (far right) are used for measuring liquid volumes in the lab.

FIGURE 7 Density is the ratio of mass to volume. Both water and copper shot float on mercury because mercury is more dense.

Density

An object made of cork feels lighter than a lead object of the same size. What you are actually comparing in such cases is how massive objects are compared with their size. This property is called density. **Density** *is the ratio of mass to volume, or mass divided by volume*. Mathematically, the relationship for density can be written in the following way.

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

The quantity m is mass, V is volume, and D is density.

The SI unit for density is derived from the base units for mass and volume—the kilogram and the cubic meter, respectively—and can be expressed as kilograms per cubic meter, kg/m^3. This unit is inconveniently large for the density measurements you will make in the laboratory. You will often see density expressed in grams per cubic centimeter, g/cm^3, or grams per milliliter, g/mL. The densities of gases are generally reported either in kilograms per cubic meter, kg/m^3, or in grams per liter, g/L.

Density is a characteristic physical property of a substance. It does not depend on the size of the sample because as the sample's mass increases, its volume increases proportionately, and the ratio of mass to volume is constant. Therefore, density can be used as one property to help identify a substance. **Table 4** shows the densities of some common materials. As you can see, cork has a density of only 0.24 g/cm^3, which is less than the density of liquid water. Because cork is less dense than water, it floats on water. Lead, on the other hand, has a density of 11.35 g/cm^3. The density of lead is greater than that of water, so lead sinks in water.

Note that **Table 4** specifies the temperatures at which the densities were measured. That is because density varies with temperature. Most objects expand as temperature increases, thereby increasing in volume. Because density is mass divided by volume, density usually decreases with increasing temperature.

TABLE 4 ***Densities of Some Familiar Materials***

Solids	Density at 20°C (g/cm^3)	Liquids	Density at 20°C (g/mL)
cork	0.24*	gasoline	0.67*
butter	0.86	ethyl alcohol	0.791
ice	0.92†	kerosene	0.82
sucrose	1.59	turpentine	0.87
bone	1.85*	water	0.998
diamond	3.26*	sea water	1.025**
copper	8.92	milk	1.031*
lead	11.35	mercury	13.6

† measured at 0°C
* typical density

** measured at 15°C

PRACTICE *Answers in Appendix E*

1. What is the density of a block of marble that occupies 310. cm^3 and has a mass of 853 g?
2. Diamond has a density of 3.26 g/cm^3. What is the mass of a diamond that has a volume of 0.351 cm^3?
3. What is the volume of a sample of liquid mercury that has a mass of 76.2 g, given that the density of mercury is 13.6 g/mL?

extension

Go to **go.hrw.com** for more practice problems that ask you to calculate density.

Keyword: HC6MEAX

Conversion Factors

A **conversion factor** *is a ratio derived from the equality between two different units that can be used to convert from one unit to the other.* For example, suppose you want to know how many quarters there are in a certain number of dollars. To figure out the answer, you need to know how quarters and dollars are related. There are four quarters per dollar and one dollar for every four quarters. Those facts can be expressed as ratios in four conversion factors.

$$\frac{4 \text{ quarters}}{1 \text{ dollar}} = 1 \qquad \frac{1 \text{ dollar}}{4 \text{ quarters}} = 1 \qquad \frac{0.25 \text{ dollar}}{1 \text{ quarter}} = 1 \qquad \frac{1 \text{ quarter}}{0.25 \text{ dollar}} = 1$$

Notice that each conversion factor equals 1. That is because the two quantities divided in any conversion factor are equivalent to each other—as in this case, where 4 quarters equal 1 dollar. Because conversion factors are equal to 1, they can be multiplied by other factors in equations without changing the validity of the equations. You can use conversion factors to solve problems through dimensional analysis. **Dimensional analysis** *is a mathematical technique that allows you to use units to solve problems involving measurements.* When you want to use a conversion factor to change a unit in a problem, you can set up the problem in the following way.

$$\text{quantity sought} = \text{quantity given} \times \text{conversion factor}$$

For example, to determine the number of quarters in 12 dollars, you would carry out the unit conversion that allows you to change from dollars to quarters.

$$\text{number of quarters} = 12 \text{ dollars} \times \text{conversion factor}$$

Next you would have to decide which conversion factor gives you an answer in the desired unit. In this case, you have dollars and you want quarters. To eliminate dollars, you must divide the quantity by dollars. Therefore, the conversion factor in this case must have dollars in the denominator and quarters in the numerator. That factor is 4 quarters/1 dollar. Thus, you would set up the calculation as follows.

$$? \text{ quarters} = 12 \text{ dollars} \times \text{conversion factor}$$

$$= 12 \cancel{\text{dollars}} \times \frac{4 \text{ quarters}}{1 \cancel{\text{dollar}}} = 48 \text{ quarters}$$

Notice that the dollars have divided out, leaving an answer in the desired unit—quarters.

Suppose you had guessed wrong and used 1 dollar/4 quarters when choosing which of the two conversion factors to use. You would have an answer with entirely inappropriate units.

$$? \text{ quarters} = 12 \text{ dollars} \times \frac{1 \text{ dollar}}{4 \text{ quarters}} = \frac{3 \text{ dollars}^2}{\text{quarter}}$$

It is always best to begin with an idea of the units you will need in your final answer. When working through the Sample Problems, keep track of the units needed for the unknown quantity. Check your final answer against what you've written as the unknown quantity.

Deriving Conversion Factors

You can derive conversion factors if you know the relationship between the unit you have and the unit you want. For example, from the fact that *deci-* means "1/10," you know that there is 1/10 of a meter per decimeter and that each meter must have 10 decimeters. Thus, from the equality (1 m = 10 dm), you can write the following conversion factors relating meters and decimeters. In this book, when there is no digit shown in the denominator, you can assume the value is 1.

$$\frac{1 \text{ m}}{10 \text{ dm}} \text{ and } \frac{0.1 \text{ m}}{\text{dm}} \text{ and } \frac{10 \text{ dm}}{\text{m}}$$

The following sample problem illustrates an example of deriving conversion factors to make a unit conversion.

SAMPLE PROBLEM B

Express a mass of 5.712 grams in milligrams and in kilograms.

SOLUTION

Given: 5.712 g

Unknown: mass in mg and kg

The expression that relates grams to milligrams is

$$1 \text{ g} = 1000 \text{ mg}$$

The possible conversion factors that can be written from this expression are

$$\frac{1000 \text{ mg}}{\text{g}} \text{ and } \frac{1 \text{ g}}{1000 \text{ mg}}$$

To derive an answer in mg, you'll need to multiply 5.712 g by 1000 mg/g.

$$5.712\ \cancel{g} \times \frac{1000\ \text{mg}}{\cancel{g}} = 5712\ \text{mg}$$

This answer makes sense because milligrams is a smaller unit than grams and, therefore, there should be more of them.

The kilogram problem is solved similarly.

$$1\ \text{kg} = 1000\ \text{g}$$

Conversion factors representing this expression are

$$\frac{1\ \text{kg}}{1000\ \text{g}} \text{ and } \frac{1000\ \text{g}}{\text{kg}}$$

To derive an answer in kg, you'll need to multiply 5.712 g by 1 kg/1000 g.

$$5.712\ \cancel{g} \times \frac{1\ \text{kg}}{1000\ \cancel{g}} = 0.005712\ \text{kg}$$

The answer makes sense because kilograms is a larger unit than grams and, therefore, there should be fewer of them.

PRACTICE

Answers in Appendix E

1. Express a length of 16.45 m in centimeters and in kilometers.
2. Express a mass of 0.014 mg in grams.

extension

Go to **go.hrw.com** for more practice problems that ask you to perform unit conversions.

Keyword: HC6MEAX

SECTION REVIEW

1. Why are standards needed for measured quantities?
2. Label each of the following measurements by the quantity each represents. For instance, a measurement of 10.6 kg/m^3 represents density.
   - **a.** 5.0 g/mL
   - **b.** 37 s
   - **c.** 47 J
   - **d.** 39.56 g
   - **e.** 25.3 cm^3
   - **f.** 325 ms
   - **g.** 500 m^2
   - **h.** 30.23 mL
   - **i.** 2.7 mg
   - **j.** 0.005 L
3. Complete the following conversions.
   - **a.** 10.5 g = ____ kg
   - **b.** 1.57 km = ____ m
   - **c.** 3.54 μg = ____ g
   - **d.** 3.5 mol = ____ μmol
   - **e.** 1.2 L = ____ mL
   - **f.** 358 cm^3 = ____ m^3
   - **g.** 548.6 mL = ____ cm^3
4. Write conversion factors for each equality.
   - **a.** 1 m^3 = 1 000 000 cm^3
   - **b.** 1 in. = 2.54 cm
   - **c.** 1 μg = 0.000 001 g
   - **d.** 1 Mm = 1 000 000 m
5. **a.** What is the density of an 84.7 g sample of an unknown substance if the sample occupies 49.6 cm^3?

   b. What volume would be occupied by 7.75 g of this same substance?

Critical Thinking

6. **INFERRING CONCLUSIONS** A student converts grams to milligrams by multiplying by the conversion factor $\frac{1\ \text{g}}{1000\ \text{mg}}$. Is the student performing this calculation correctly?