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## 2 MATTER AND PROPERTIES OF MATTER

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### 2.1 WHAT IS MATTER?

In Chapter 1, chemistry was defined as the science of **matter**, *anything that has mass and occupies space*. This chapter deals specifically with matter—what it is, how it acts, what its properties are. Most of this book is concerned with chemical processes—those in which chemical bonds are broken and formed to produce different substances—and the *chemical properties* that they define. However, before going into chemical phenomena in any detail, it is helpful to consider matter in its bulk form, aside from its chemical behavior. Non-chemical aspects include physical

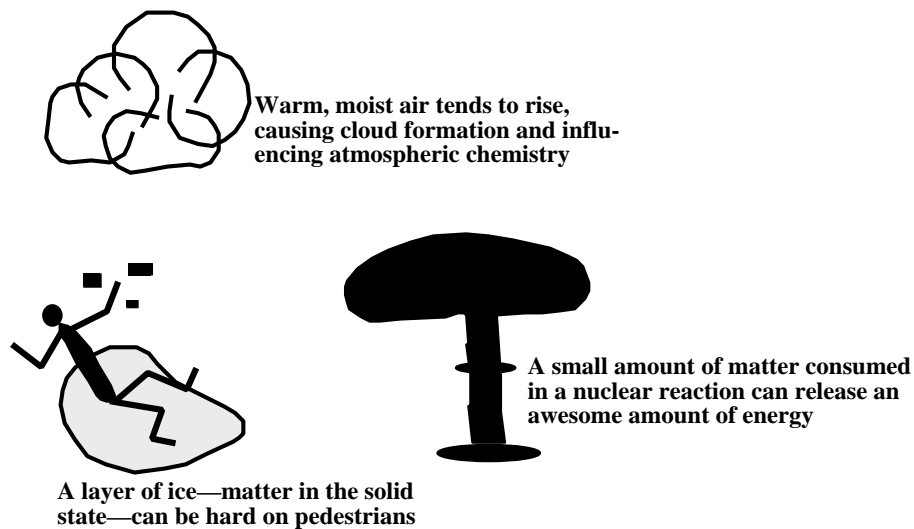


Figure 2.1 Different kinds of matter have a vast variety of properties that determine what matter does and how it is used.

state—whether a substance is a solid, liquid, or gas—color, hardness, extent to which matter dissolves in water, melting temperature, and boiling temperature. These kinds of *physical properties* are essential in describing the nature of matter and the chemical changes that it undergoes. They are also essential in understanding and describing its environmental chemical and biological behavior. For example, as discussed in Chapter 11, the temperature-density relationships of water result in stratification of bodies of water such that much different chemical and biochemical processes occur at different depths in a pond, lake, or reservoir. An analogous stratification of the atmosphere profoundly influences atmospheric chemistry.

## 2.2 CLASSIFICATION OF MATTER

Matter exists in the form of either *elements* or *compounds* (see Sections 1.3 and 1.4, respectively). Recall that compounds consist of atoms of two or more elements bonded together. Chemical changes in which chemical bonds are broken or formed are involved when a compound is produced from two or more elements or from other compounds, or when one or more elements are isolated from a compound. These kinds of changes, which are illustrated by the examples in Table 2.1, profoundly affect the properties of matter. For example, the first reaction in the table shows that sulfur (S), a yellow, crumbly solid, reacts with oxygen, a colorless, odorless gas that is essential for life, to produce sulfur dioxide, a toxic gas with a choking odor. In the second reaction in Table 2.1, the hydrogen atoms in methane, CH<sub>4</sub>, a highly flammable, light gas, are replaced by chlorine atoms to produce carbon tetrachloride, CCl<sub>4</sub>, a dense liquid so nonflammable that it has been used as a fire extinguisher.

**Table 2.1 Chemical Processes Involving Elements and Compounds**

Type of Process	Chemical reaction
Two elements combining to form a chemical compound, sulfur plus oxygen yields sulfur dioxide	$S + O_2 \rightarrow SO_2$
A compound reacting with an element to form two different compounds, methane plus chlorine yields carbon tetrachloride plus hydrogen chloride	$CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$
Two compounds reacting to form a different compound, calcium oxide plus water yields calcium hydroxide	$CaO + H_2O \rightarrow Ca(OH)_2$
A compound plus an element reacting to produce another compound and another pure element, iron oxide plus carbon yields carbon monoxide plus elemental iron (Fe)	$Fe_2O_3 + 3C \rightarrow 3CO + 2Fe$
A compound breaking down to its constituent elements, passing an electrical current through water yields hydrogen and oxygen	$2H_2O \rightarrow 2H_2 + O_2$

## Some General Types of Matter

In discussing matter, some general terms are employed that are encountered frequently enough that they should be mentioned here. More-exact definitions are given later in the text.

Elements are divided between metals and nonmetals; several elements with properties of both metals and nonmetals are called **metalloids**. **Metals** are elements that are generally solid, shiny in appearance, electrically conducting, and malleable—that is, they can be pounded into flat sheets without disintegrating. Examples of metals are iron, copper, and silver. Most metallic objects that are commonly encountered are not composed of just one kind of elemental metal, but are alloys consisting of homogeneous mixtures of two or more metals. **Nonmetals** often have a dull appearance and are not at all malleable. In contrast to metals, all of which except mercury exist as solids at room temperatures, nonmetals frequently occur as gases or liquids. Colorless oxygen gas, green chlorine gas (transported and stored as a liquid under pressure), and brown bromine liquid are common nonmetals. In a very general sense, we tend to regard as nonmetals substances that actually contain metal in a chemically combined form. One would classify table salt, sodium chloride, as a nonmetal even though it contains chemically bound sodium metal.

Another general classification of matter is in the categories of inorganic and organic substances. **Organic substances** consist of virtually all compounds that contain carbon, including substances made by life processes (wood, flesh, cotton, wool), petroleum, natural gas (methane), solvents (dry cleaning fluids), synthetic fibers, and plastics. All of the rest of the chemical kingdom is composed of **inorganic substances** made up of virtually all substances that do not contain carbon. These include metals, rocks, table salt, water, sand, and concrete.

## Mixtures and Pure Substances

Matter consisting of only one compound or of only one form of an element is a **pure substance**; all other matter is a **mixture** of two or more substances. The compound water with nothing dissolved in it is a pure substance. Highly purified helium gas is an elemental pure substance. The composition of a pure substance is defined and constant and its properties are always the same under specified conditions. Therefore, pure water at  $-1^{\circ}\text{C}$  is always a solid (ice), whereas a mixture of water and salt might be either a liquid or a solid at that temperature, depending upon the amount of salt dissolved in the water. Air is a mixture of elemental gases and compounds, predominantly nitrogen, oxygen, argon, carbon dioxide, and water vapor. Drinking water is a mixture containing calcium ion ( $\text{Ca}^{2+}$ , see ions in Section 1.4), hydrogen carbonate ion, (bicarbonate,  $\text{HCO}_3^-$ ), nitrogen gas, carbon dioxide gas, and other substances dissolved in the water. Mixtures can be separated into their constituent pure substances by **physical processes**. Liquified air is distilled to isolate pure oxygen (used in welding, industrial processes, for some specialized wastewater treatment processes, and for breathing by people with emphysema), liquid nitrogen (used for quick-freezing frozen foods), and argon (used in specialized types of welding because of its chemically non-reactive nature).

## *Homogeneous and Heterogeneous Mixtures*

A **heterogeneous mixture** is one that is not uniform throughout and possesses readily distinguishable constituents (matter in different phases, see Section 2.5) that can be isolated by means as simple as mechanical separation. Concrete (see [Figure 2.2](#)) is such a mixture; individual grains of sand and pieces of gravel can be separated from concrete with a pick or other tool. **Homogeneous mixtures** are uniform throughout; to observe different constituents would require going down to molecular levels. Homogeneous mixtures may have varying compositions and consist of two or more chemically distinct constituents, so they are not pure substances. Air, well filtered to remove particles of dust, pollen, or smoke, is a homogeneous mixture. Homogeneous mixtures are also called **solutions**, a term that is usually applied to mixtures composed of gases, solids, and other liquids dissolved in a liquid. A hazardous-waste leachate is a solution containing—in addition to water—contaminants such as dissolved acids, iron, heavy metals, and toxic organic compounds. Whereas mechanical means, including centrifugation and filtration, can be used to separate the solids from the liquids in a heterogeneous mixture of solids and liquids, the isolation of components of homogeneous mixtures requires processes such as distillation and freezing.

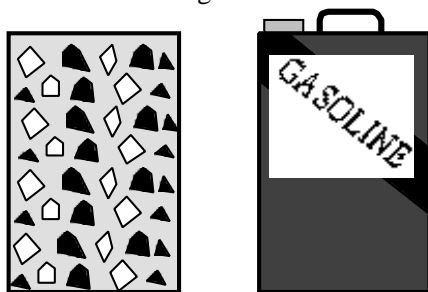


Figure 2.2. Concrete (left) is a heterogeneous mixture with readily visible grains of sand, pieces of gravel, and cement dust. Gasoline is a homogeneous mixture consisting of a solution of petroleum hydrocarbon liquids, dye, and additives such as ethyl alcohol to improve its fuel qualities.

### *Summary of the Classification of Matter*

As discussed above, all matter consists of compounds or elements. These may exist as pure substances or mixtures; the latter may be either homogeneous or heterogeneous. These relationships are summarized in [Figure 2.3](#).

## **2.3 QUANTITY OF MATTER: THE MOLE**

One of the most fundamental characteristics of a specific body of matter is the quantity of it. In discussing the quantitative chemical characteristics of matter it is essential to have a way of expressing quantity in a way that is proportional to the number of individual entities of the substance—that is, atoms, molecules, or ions—in numbers that are readily related to the properties of the atoms or molecules of the substance. The simplest way to do this would be as individual atoms, molecules, or ions, but for laboratory quantities, these would number the order of  $10^{23}$ , far too large

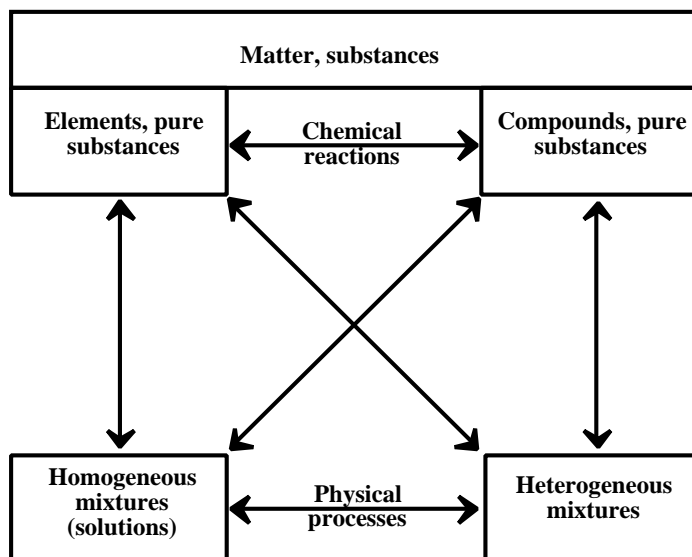


Figure 2.3 Classification of matter

to be used routinely in expressing quantities of matter. Instead, such quantities are readily expressed as moles of substance. A mole is defined in terms of specific entities, such as atoms of Ar, molecules of  $\text{H}_2\text{O}$ , or  $\text{Na}^+$  and  $\text{Cl}^-$  ions, each pair of which composes a “molecule” of NaCl. A **mole** is defined as *the quantity of substance that contains the same number of specified entities as there are atoms of C in exactly 0.012 kg (12 g) of carbon-12*. It is easier to specify the quantity of a substance equivalent to its number of moles than it is to define the mole. To do so, simply state the atomic mass (of an element) or the molecular mass (of a compound) and affix “mole” to it as shown by the examples below:

- *A mole of argon, which always exists as individual Ar atoms:* The atomic mass of Ar is 40.0. Therefore exactly one mole of Ar is 40.0 grams of argon.
- *A mole of molecular elemental hydrogen,  $\text{H}_2$ :* The atomic mass of H is 1.0, the molecular mass of  $\text{H}_2$  is, therefore, 2.0, and a mole of  $\text{H}_2$  is 2.0 g of  $\text{H}_2$ .
- *A mole of methane,  $\text{CH}_4$ :* The atomic mass of H is 1.0 and that of C is 12.0, so the molecular mass of  $\text{CH}_4$  is 16.0. Therefore a mole of methane has a mass of 16.0 g.

## The Mole and Avogadro’s Number

In Section 1.6, Avogadro’s number was mentioned as an example of a huge number,  $6.02 \times 10^{23}$ . **Avogadro’s number** is *the number of specified entities in a mole of substance*. The “specified entities” may consist of atoms or molecules or they may be groups of ions making up the smallest possible unit of an ionic

compound, such as 2 Na<sup>+</sup> ions and one S<sup>2-</sup> ion in Na<sub>2</sub>S. (It is not really correct to refer to Na<sub>2</sub>S as a molecule because the compound consists of ions arranged in a crystalline structure such that there are 2 Na<sup>+</sup> ions for each S<sup>2-</sup> ion.) A general term that covers all these possibilities is the **formula unit**. The average mass of a formula unit is called the **formula mass**. Examples of the terms defined in this section are given in [Table 2.2](#).

## 2.4 PHYSICAL PROPERTIES OF MATTER

**Physical properties** of matter are those that can be measured without altering the chemical composition of the matter. A typical physical property is color, which can be observed without changing matter at all. Malleability of metals, the degree to which they can be pounded into thin sheets, certainly alters the shape of an object but does not change it chemically. On the other hand, observation of a *chemical property*, such as whether sugar burns when ignited in air, potentially involves a complete change in the chemical composition of the substance tested.

Physical properties are important in describing and identifying particular kinds matter. For example, if a substance is liquid at room temperature, has a lustrous metallic color, conducts electricity well, and has a very high density (mass per unit volume) of 13.6 g/cm<sup>3</sup>, it is doubtless elemental mercury metal. Physical properties are very useful in assessing the hazards and predicting the fates of environmental pollutants. An organic substance that readily forms a vapor (volatile organic compound) will tend to enter the atmosphere or to pose an inhalation hazard. A brightly colored water soluble pollutant may cause deterioration of water quality by adding water “color.” Much of the health hazard of asbestos is its tendency to form extremely small-diameter fibers that are readily carried far into the lungs and that can puncture individual cells.

Several important physical properties of matter are discussed in this section. The most commonly considered of these are density, color, and solubility. Thermal properties are addressed separately in Section 2.9.

### Density

**Density** (d) is defined as mass per unit volume and is expressed by the formula

$$d = \frac{\text{mass}}{\text{volume}} \quad (2.4.1)$$

Density is useful for identifying and characterizing pure substances and mixtures. For example, the density of a mixture of automobile system antifreeze and water, used to prevent the engine coolant from freezing in winter or boiling in summer, varies with the composition of the mixture. Its composition and, therefore, the degree of protection that it offers against freezing can be estimated by measuring its density and relating that through a table to the freezing temperature of the mixture.

Density may be expressed in any units of mass or volume. The densities of liquids and solids are normally given in units of grams per cubic centimeter (g/cm<sup>3</sup>, the same as grams per milliliter, g/mL). These values are convenient because they are of the order of 1 g/cm<sup>3</sup> for common liquids and solids.

**Table 2.2 Relationships Involving Moles of Substance**

Substance	Formula unit	Formula mass	Mass of 1 mole	Numbers and kinds of individual entities in 1 mole
Helium	He atom	4.003 <sup>a</sup>	4.003 g	$6.02 \times 10^{23}$ (Avogadro's number) of He atoms
Fluorine gas	F <sub>2</sub> molecule	38.00 <sup>b</sup>	38.00 g	$6.02 \times 10^{23}$ F <sub>2</sub> molecules $2 \times 6.02 \times 10^{23}$ F atoms
Methane	CH <sub>4</sub> molecules	16.04 <sup>b</sup>	16.04 g	$6.02 \times 10^{23}$ CH <sub>4</sub> molecules $6.02 \times 10^{23}$ C atoms $4 \times 6.02 \times 10^{23}$ H atoms
Sodium oxide	Na <sub>2</sub> O	62.00 <sup>c</sup>	62.00 g	$6.02 \times 10^{23}$ Na <sub>2</sub> O formula units $2 \times 6.02 \times 10^{23}$ Na <sup>+</sup> ions $6.02 \times 10^{23}$ O <sup>2-</sup> atoms

<sup>a</sup> Specifically, atomic mass

<sup>b</sup> Specifically, molecular mass

<sup>c</sup> Reference should be made to formula mass for this compound because it consists of Na<sup>+</sup> and S<sup>2-</sup> ions in a ratio of 2/1.

The volume of a given mass of substance varies with temperature, so the density is a function of temperature. This variation is relatively small for solids, greater for liquids, and very high for gases. Densities of gases vary a great deal with pressure, as well. The temperature-dependent variation of density is an important property of water and results in stratification of bodies of water, which greatly affects the environmental chemistry that occurs in lakes and reservoirs. The density of liquid water has a maximum value of 1.0000 g/mL at 4°C, is 0.9998 g/mL at 0°C, and is 0.9970 g/mL at 25°C. The combined temperature/pressure relationship for the density of air causes air to become stratified into layers, particularly the troposphere near the surface and the stratosphere from about 13 to 50 kilometers altitude.

Whereas liquids and solids have densities of the order of 1 to several g/cm<sup>3</sup>, gases at atmospheric temperature and pressure have densities of only about one one thousandth as much (see [Table 2.3](#) and [Figure 2.4](#)).

**Table 2.3 Densities of Some Solids, Liquids, and Gases**

Solids	d, g/cm <sup>3</sup>	Liquids	d, g/cm <sup>3</sup>	Gases	d, g/L
Sugar	1.59	Benzene	0.879	Helium	0.178
Sodium chloride	2.16	Water (4°C)	1.0000	Methane	0.714
Iron	7.86	Carbon		Air	1.293
Lead	11.34	tetrachloride	1.595	Chlorine	3.17
		Sulfuric acid	1.84		



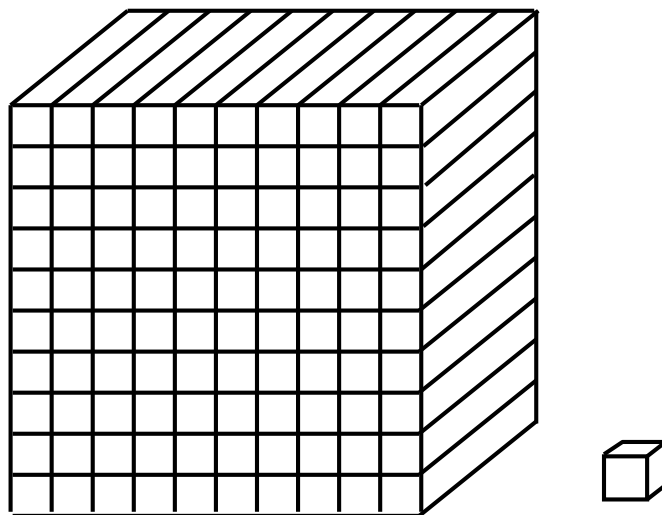


Figure 2.4 A quantity of air that occupies a volume of 1000 cm<sup>3</sup> of air at room temperature and atmospheric pressure (left) has about the same mass as 1 cm<sup>3</sup> of liquid water (right).

## Specific Gravity

Often densities are expressed by means of **specific gravity** defined as the ratio of the density of a substance to that of a standard substance. For solids and liquids, the standard substance is usually water; for gases it is usually air. For example, the density of ethanol (ethyl alcohol) at 20°C is 0.7895 g/mL. The specific gravity of ethanol at 20°C referred to water at 4°C is given by

$$\text{Specific gravity} = \frac{\text{density of ethanol}}{\text{density of water}} = \frac{0.7895 \text{ g/mL}}{1.0000 \text{ g/mL}} = 0.7895 \quad (2.4.2)$$

For an exact value of specific gravity, the temperatures of the substances should be specified. In this case the notation of “specific gravity of ethanol at 20°/4° C” shows that the specific gravity is the ratio of the density of ethanol at 20°C to that of water at 4°C.

## Color

Color is one of the more useful properties for identifying substances without doing any chemical or physical tests. A violet vapor, for example, is characteristic of iodine. A red/brown gas could well be bromine or nitrogen dioxide (NO<sub>2</sub>); a practiced eye can distinguish between the two. Just a small amount of potassium permanganate, KMnO<sub>4</sub>, in solution provides an intense purple color. A characteristic yellow/brown color in water may be indicative of organically bound iron.

The human eye responds to colors of electromagnetic radiation ranging in wavelength from about 400 nanometers (nm) to somewhat over 700 nm. Within this wavelength range humans see **light**; immediately below 400 nm is **ultraviolet**

**radiation**, and somewhat above 700 nm is **infrared radiation** (see Figure 2.5). Light with a mixture of wavelengths throughout the visible region, such as sunlight, appears white to the eye. Light over narrower wavelength regions has the following colors: 400–450 nm, blue; 490–550 nm, green; 550–580 nm, yellow; 580–650 nm, orange, 650 nm–upper limit of visible region, red. Solutions are colored because of the light they absorb. Red, orange, and yellow solutions absorb violet and blue light; purple solutions absorb green and yellow light; and blue and green solutions absorb orange and red light. Solutions that do not absorb light are colorless (clear); solids that do not absorb light are white.

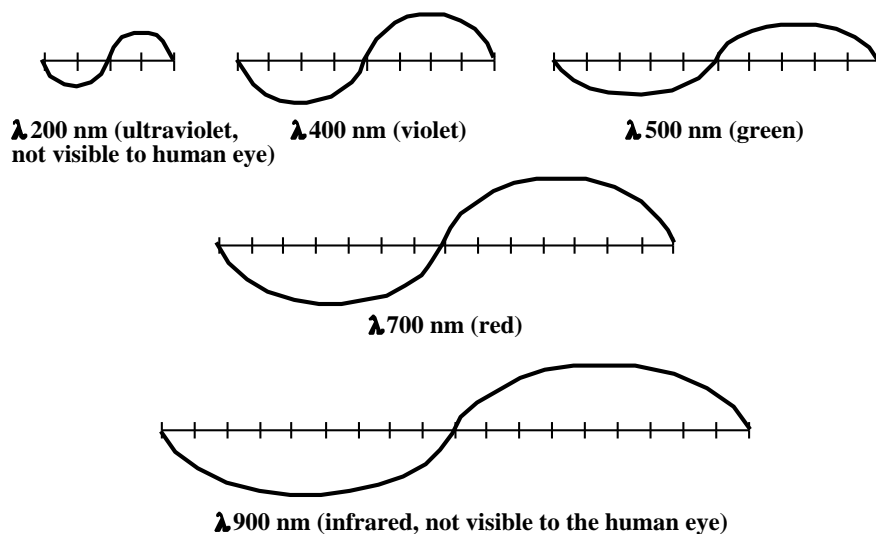


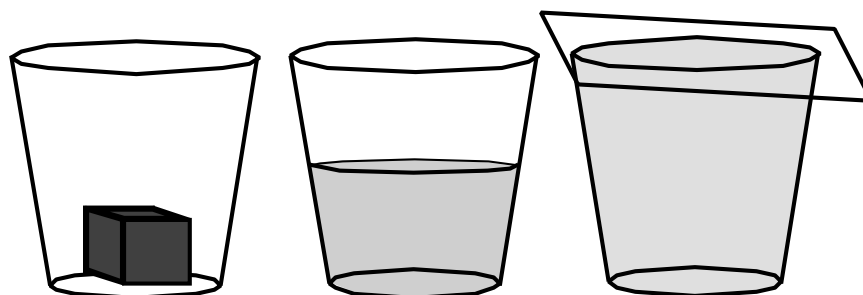
Figure 2.5 Electromagnetic radiation of different wavelengths. Each division represents 50 nm.

## 2.5 STATES OF MATTER

Figure 2.6 illustrates the three **states** in which matter can exist. **Solids** have a definite shape and volume. **Liquids** have an indefinite shape and take on the shape of the container in which they are contained. Solids and liquids are not significantly compressible, which means that a specific quantity of a substance has a definite volume and cannot be squeezed into a significantly smaller volume. **Gases** take on both the shape and volume of their containers. A quantity of gas can be compressed to a very small volume and will expand to occupy the volume of any container into which it is introduced.

Each of these separate phases is discussed in separate sections in this chapter. Everyone is familiar with the three states of matter for water. These are the following:

- Gas: Water vapor in a humid atmosphere, steam
- Liquid: Water in a lake, groundwater
- Solid: Ice in polar ice caps, snow in snowpack



**A solid has a definite shape and volume regardless of the container into which it is placed.**

**A quantity of liquid has a definite volume, but takes on the shape of its container.**

**A quantity of gas has the shape and volume of the container it occupies.**

Figure 2.6 Representations of the three states of matter.

Changes in matter from one phase to another are very important in the environment. For example, water vapor changing from the gas phase to liquid results in cloud formation or precipitation. Water is desalinated by producing water vapor from sea water, leaving the solid salt behind, and recondensing the pure water vapor as a salt-free liquid. Some organic pollutants are extracted from water for chemical analysis by transferring them from the water to another organic phase that is immiscible with water. The condensation of organic pollutants from gaseous materials to solid products in the atmosphere is responsible for the visibility-reducing particulate matter in photochemical smog. Additional examples of phase changes are discussed in Section 2.10 and some of the energy relationships involved are covered in Section 2.9.

## 2.6 GASES

We live at the bottom of a “sea” of gas—the earth’s atmosphere. It is composed of a mixture of gases, the most abundant of which are nitrogen, oxygen, argon, carbon dioxide, and water vapor. Although these gases have neither color nor odor, so that we are not aware of their presence, they are crucial to the well-being of life on Earth. Deprived of oxygen, an animal loses consciousness and dies within a short time. Nitrogen extracted from the air is converted to chemically bound forms that are crucial to plant growth. Plants require carbon dioxide and water vapor condenses to produce rain.

Gases and the atmosphere are addressed in more detail in Chapter 14. At this point, however, it is important to have a basic understanding of the nature and behavior of gases. Physically, gases are the “loosest” form of matter. A quantity of gas has neither a definite shape nor a definite volume so that it takes on the shape and volume of the container in which it is held. The reason for this behavior is that gas molecules move independently and at random, bouncing off each other as they do so. They move very rapidly; at 0°C the average molecule of hydrogen gas moves at 3600 miles per hour, 6 to 7 times as fast as a passenger jet airplane. Gas molecules colliding with container walls exert **pressure**. An underinflated automobile tire has a low pressure because there are relatively few gas molecules in the tire to collide with

the walls of the tire, which feel relatively soft when pushed. As more air is added the pressure is increased, the tire feels harder, and the pressure within the tire may become so high that it can no longer be contained by the tire structure so that the tire wall bursts.

A quantity of a gas is mostly empty space. That this is so is rather dramatically illustrated by comparison of the high volume of gas compared with the same amount of material in a liquid or solid. For example, at 100°C a mole (18 g) of liquid water occupies a little more than 18 mL of volume, equivalent to just a few teaspoonfuls. When enough heat energy is added to the water to convert it all to gaseous steam at 100°C, the volume becomes 30,600 mL, which is 1,700 times the original volume. This huge increase in volume occurs because of the large distances separating gas molecules compared with their own diameters. It is this great distance that allows gas to be compressed (pushed together). Therefore, air with a volume of more than twice that of a tire can be forced into the tire when it becomes pressurized because the air molecules can be forced closer together under the pressure in the tire when it is inflated.

The rapid, constant motion of gas molecules explains the phenomenon of **diffusion** in which gases move large distances from their sources. Diffusion is responsible for much of the hazard of volatile, flammable liquids, such as gasoline. Molecules of gasoline evaporated from an open container of this liquid can spread from their source. If the gaseous gasoline reaches an ignition source, such as an open flame, it may ignite and cause a fire or explosion.

## The Gas Laws

To describe the physical and chemical behavior of gases, it is essential to have a rudimentary understanding of the relationships among the following: Quantity of gas in numbers of moles ( $n$ , see Section 3.2), volume ( $V$ ), temperature ( $T$ ), and pressure ( $P$ ). These relationships are expressed by the **gas laws**. The gas laws as discussed here apply to **ideal gases**, for which it is assumed that the molecules of gas have negligible volume, that there are no forces of attraction between them, and that they collide with each other in a perfectly elastic manner. Real gases such as  $H_2$ ,  $O_2$ , He, and Ar behave in a nearly ideal fashion at moderate pressures and all but very low temperatures. The gas laws were derived from observations of the effects on gas volume of changing pressure, temperature, and quantity of gas, as stated by Boyle's law, Charles' law, and Avogadro's law, respectively. These three laws are defined below:

### *Boyle's Law*

As illustrated in [Figure 2.7](#), doubling the pressure on a quantity of a gas halves its volume. This illustrates **Boyle's law**, according to which *at constant temperature, the volume of a fixed quantity of gas is inversely proportional to the pressure of the gas*. Boyle's law may be stated mathematically as

$$V = (\text{a constant}) \times \frac{1}{P} \quad (2.6.1)$$

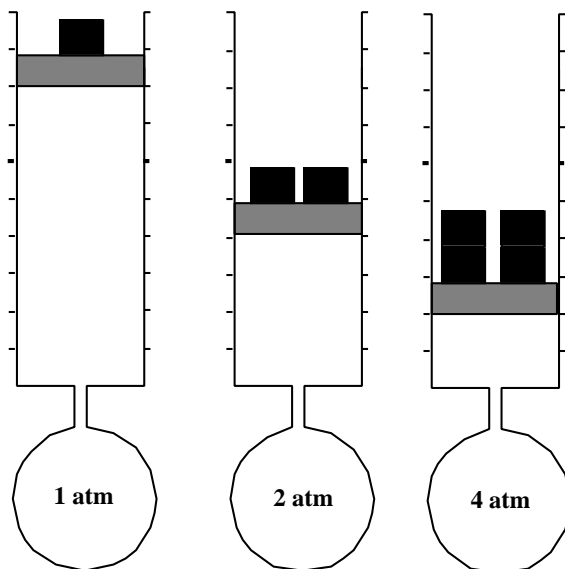


Figure 2.7 In accordance with Boyle's law, doubling the pressure of a specific quantity of gas at a constant temperature makes its volume half as great.

This inverse relationship between the volume of a gas and its pressure is a consequence of the compressibility of matter in the gas phase (recall that solids and liquids are not significantly compressible).

### *Charles' Law*

**Charles' law** gives the relationship between gas volume and temperature and states that *the volume of a fixed quantity of gas is directly proportional to the absolute temperature ( $^{\circ}\text{C} + 273$ ) at constant pressure*. This law may be stated mathematically as

$$V = (\text{a constant}) \times T \quad (2.6.2)$$

where  $T$  is the temperature in K (see Section 1.12). According to this relationship doubling the absolute temperature of a fixed quantity of gas at constant pressure doubles the volume.

### *Avogadro's Law*

The third of three fundamentally important gas laws is **Avogadro's law** relating the volume of gas to its quantity in moles. This law states that *at constant temperature and pressure the volume of a gas is directly proportional to the number of molecules of gas, commonly expressed as moles*. Mathematically, this relationship is

$$V = (\text{a constant}) \times n \quad (2.6.3)$$

where  $n$  is the number of moles of gas.

## The General Gas Law

The three gas laws just defined may be combined into a **general gas law** stating that *the volume of a quantity of ideal gas is proportional to the number of moles of gas and its absolute temperature and inversely proportional to its pressure*. Mathematically, this law is

$$V = (\text{a constant}) \times \frac{nT}{P} \quad (2.6.4)$$

Designating the proportionality constant as the **ideal gas constant, R**, yields the **ideal gas equation**:

$$V = \frac{RnT}{P} \text{ or } PV = nRT \quad (2.6.5)$$

The units of R depend upon the way in which the ideal gas equation is used. For calculations involving volume in liters and pressures in atmospheres, the value of R is 0.0821 L-atm/deg-mol.

The ideal gas equation shows that at a chosen temperature and pressure a mole of any gas should occupy the same volume. A temperature of 0°C (273.15 K) and 1 atm pressure have been chosen as **standard temperature and pressure (STP)**. *At STP the volume of 1 mole of ideal gas is 22.4 L*. This volume is called the **molar volume of a gas**. It should be remembered that at a temperature of 273 K and a pressure of 1 atm the volume of 1 mole of ideal gas is 22.4 L. Knowing these values it is always possible to calculate the value of R as follows:

$$R = \frac{PV}{nT} = \frac{1 \text{ atm} \times 22.4 \text{ L}}{1 \text{ mole} \times 273 \text{ K}} = 0.0821 \frac{\text{L} \times \text{atm}}{\text{deg} \times \text{mol}} \quad (2.6.6)$$

## Gas Law Calculations

A common calculation is that of the volume of a gas, starting with a particular volume and changing temperature and/or pressure. This kind of calculation follows logically from the ideal gas equation, but can also be reasoned out knowing that an *increase* in temperature causes an *increase* in volume, whereas an *increase* in pressure causes a *decrease* in volume. Therefore, a second volume,  $V_2$ , is calculated from an initial volume,  $V_1$ , by the relationship

$$V_2 = V_1 \times \frac{\boxed{\phantom{000}}}{\boxed{\phantom{000}}} \times \frac{\boxed{\phantom{000}}}{\boxed{\phantom{000}}} \quad (2.6.7)$$

↑                      ↑  
**Ratio of temperatures**      **Ratio of pressures**

This relationship requires only that one remember the following:

- If the temperature *increases* (if  $T_2$  is greater than  $T_1$ ), the volume *increases*. Therefore,  $T_2$  is always placed over  $T_1$  in the ratio of temperatures.

- If the pressure *increases* (if  $P_2$  is greater than  $P_1$ ), the volume *decreases*. Therefore,  $P_1$  is always placed over  $P_2$  in the ratio of pressures.
- If there is no change in temperature or pressure, the corresponding ratio remains 1.

These relationships can be illustrated by several examples.

### *Charles' Law Calculation*

When the temperature of a quantity of gas changes while the pressure stays the same, the resulting calculation is a *Charles' law calculation*. As an example of such a calculation, calculate the volume of a gas with an initial volume of 10.0 L when the temperature changes from  $-11.0^\circ\text{C}$  to  $95.0^\circ\text{C}$  at constant pressure. The first step in solving any gas law problem involving a temperature change is to convert Celsius temperatures to kelvin:

$$T_1 = 273 + (-11) = 262 \text{ K} \quad T_2 = 273 + 95 = 368 \text{ K}$$

Substitution into Equation 2.6.7 gives

$$V_2 = V_1 \times \frac{368 \text{ K}}{262 \text{ K}} \times 1 (\text{factor for pressure} = 1 \text{ because } P_2 = P_1) \quad (2.6.8)$$

$$V_2 = 10.0 \text{ L} \times \frac{368 \text{ K}}{262 \text{ K}} = 14.0 \text{ L}$$

Note that in this calculation it is seen that the temperature increases, this increases the volume, so the higher temperature is placed over the lower. Furthermore, since  $P_2 = P_1$ , the factor for the pressure ratio simply drops out.

Calculate next the volume of a gas with an initial volume of 10.0 L after the temperature decreases from  $111.0^\circ\text{C}$  to  $2.0^\circ\text{C}$  at constant pressure:

$$T_1 = 273 + (111) = 384 \text{ K} \quad T_2 = 273 + 2 = 275 \text{ K}$$

Substitution into Equation 2.6.6 gives

$$\begin{aligned} V_2 &= V_1 \times \frac{275 \text{ K}}{384 \text{ K}} = 10.0 \text{ L} \times \frac{275 \text{ K}}{384 \text{ K}} \\ &= 7.16 \text{ L} \end{aligned} \quad (2.6.9)$$

In this calculation the temperature decreases, this decreases the volume, so the lower temperature is placed over the higher (mathematically it is still  $T_2$  over  $T_1$  regardless of which way the temperature changes).

### *Boyle's Law Calculation*

When the pressure of a quantity of gas changes while the temperature stays the same, the resulting calculation is a *Boyle's law calculation*. As an example,

calculate the new volume that results when the pressure of a quantity of gas initially occupying 12.0 L is changed from 0.856 atm to 1.27 atm at constant temperature. In this case, the pressure has increased, this decreases the volume of the gas so that  $V_2$  is given by the following:

$$V_2 = V_1 \times \frac{0.856 \text{ atm}}{1.27 \text{ atm}} \times 1 \text{ (because } T_2 = T_1) = 12.0 \text{ L} \times \frac{0.856 \text{ atm}}{1.27 \text{ atm}} \quad (2.6.10)$$

$$= 8.09 \text{ L}$$

Calculate next the volume of a gas with an initial volume of 15.0 L when the pressure decreases from 1.71 atm to 1.07 atm at constant temperature. In this case the pressure decreases, this increases the volume, so the appropriate factor to multiply by is 1.71/1.07:

$$V_2 = V_1 \times \frac{1.71 \text{ atm}}{1.07 \text{ atm}} = 15.0 \text{ L} \times \frac{1.71 \text{ atm}}{1.07 \text{ atm}} = 24.0 \text{ L} \quad (2.6.11)$$

### *Calculations Using the Ideal Gas Law:*

The ideal gas law equation

$$PV = nRT \quad (2.6.5)$$

may be used to calculate any of the variables in it if the others are known. As an example, calculate the volume of 0.333 moles of gas at 300 K under a pressure of 0.950 atm:

$$V = \frac{nRT}{P} = \frac{0.333 \text{ mol} \times 0.0821 \text{ L atm/K mol} \times 300 \text{ K}}{0.950 \text{ atm}} = 8.63 \text{ L} \quad (2.6.12)$$

As another example, calculate the temperature of 2.50 mol of gas that occupies a volume of 52.6 L under a pressure of 1.15 atm:

$$T = \frac{PV}{nR} = \frac{1.15 \text{ atm} \times 52.6 \text{ L}}{2.50 \text{ mol} \times 0.0821 \text{ L atm/K mol}} = 295 \text{ K} \quad (2.6.13)$$

Other examples of gas law calculations are given in the problem section at the end of this chapter.

Where subscripts 1 and 2 denote parameters before and after any change, respectively, the following relationship can be derived from the fact that R is a constant in the ideal gas law equation:

$$\frac{P_2 V_2}{n_2 T_2} = \frac{P_1 V_1}{n_1 T_1} \quad (2.6.14)$$

derived from the fact that R is a constant, where the subscripts 1 and 2 denote values of the parameters designated before and after these are changed. By holding the appropriate parameters constant, this equation can be used to derive the mathematical relationships used in calculations involving Boyle's law, Charles' law, or Avogadro's law, as well as any combination of these laws.



Exercise: Derive an equation that will give the volume of a gas after all three of the parameters P, T, and n have been changed.

Answer: Rearrangement of Equation 2.6.13 gives the relationship needed:

$$V_2 = V_1 \frac{P_1 n_2 T_2}{P_2 n_1 T_1} \quad (2.6.15)$$

## 2.7 LIQUIDS AND SOLUTIONS

Whereas the molecules in gases are relatively far apart and have minimal attractive forces among them (none in the case of ideal gases, discussed in the preceding section), molecules of liquids are close enough together that they can be regarded as “touching” and are strongly attracted to each other. Like those of gases, however, the molecules of liquids move freely relative to each other. These characteristics give rise to several significant properties of liquids. Whereas gases are mostly empty space, most of the volume of a liquid is occupied by molecules; therefore, the densities of liquids are much higher than those of gases. Because there is little unoccupied space between molecules of liquid, liquids are only slightly *compressible*; doubling the pressure on a liquid causes a barely perceptible change in volume, whereas it halves the volume of an ideal gas. Liquids do not expand to fill their containers as do gases. Essentially, therefore, a given quantity of liquid occupies a fixed volume. Because of the free movement of molecules relative to each other in a liquid, it takes on the shape of that portion of the container that it occupies.

### Evaporation and Condensation of Liquids

Consider the surface of a body of water as illustrated in [Figure 2.8](#). The molecules of liquid water are in constant motion relative to each other. They have a distribution of energies such that there are more higher energy molecules at higher temperatures. Some molecules are energetic enough that they escape the attractive forces of the other molecules in the mass of liquid and enter the gas phase. This phenomenon is called **evaporation**.

Molecules of liquid in the gas phase, such as water vapor molecules above a body of water, may come together or strike the surface of the liquid and reenter the liquid phase. This process is called **condensation**. It is what happens, for example, when water vapor molecules in the atmosphere strike the surface of very small water droplets in clouds and enter the liquid water phase, causing the droplets to grow to sufficient size to fall as precipitation.

### Vapor Pressure

In a confined area above a liquid, equilibrium is established between the evaporation and condensation of molecules from the liquid. For a given temperature, this results in a steady-state level of the vapor, which can be described as a pressure. Such a pressure is called the **vapor pressure** of the liquid (see [Figure 2.8](#)).

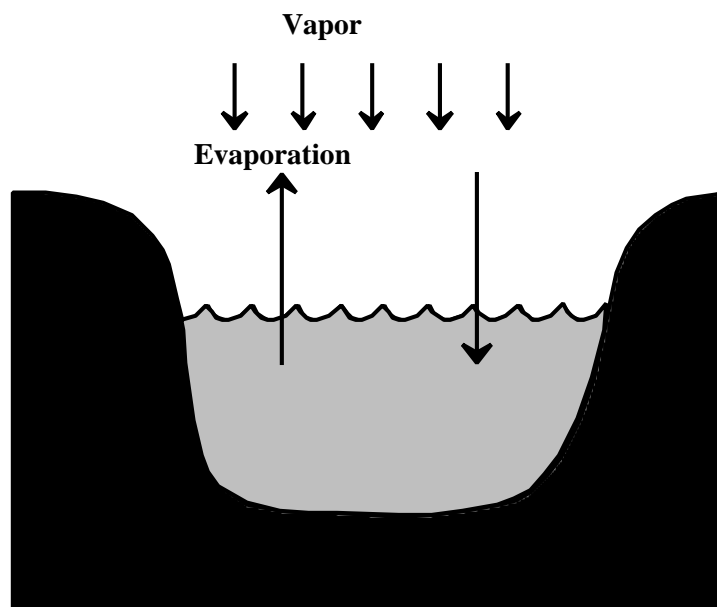


Figure 2.8 Evaporation and condensation of liquid.

Vapor pressure is very important in determining the fates and effects of substances, including hazardous substances, in the environment. Loss of a liquid by evaporation to the atmosphere increases with increasing vapor pressure and is an important mechanism by which volatile pollutants enter the atmosphere. High vapor pressure of a flammable liquid can result in the formation of explosive mixtures of vapor above a liquid, such as in a storage tank.

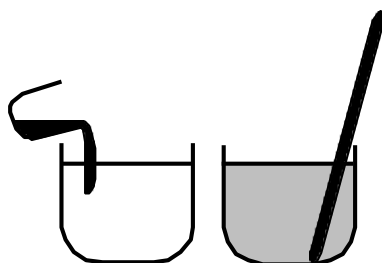
## Solutions

**Solutions** were mentioned in Section 2.2 as homogeneous mixtures. Here solutions are regarded as liquids in which quantities of gases, solids, or other liquids are dispersed as individual ions or molecules. Rainwater, for example, is a solution containing molecules of  $N_2$ ,  $O_2$  and  $CO_2$  from air dispersed among the water molecules. Rainwater from polluted air may contain harmful substances such as sulfuric acid ( $H_2SO_4$ ) in small quantities among the water molecules. The predominant liquid constituent of a solution is called the **solvent**. A substance dispersed in it is said to be **dissolved** in the solvent and is called the **solute**.

### *Solubility*

Mixing and stirring some solids, such as sugar, with water results in a rapid and obvious change in which all or part of the substance seems to disappear as it dissolves. When this happens, the substance is said to *dissolve*, and one that does so to a significant extent is said to be **soluble** in the solvent. Substances that do not dissolve to a perceptible degree are called **insoluble**. *Solubility* is a significant

physical property. It is a relative term. For example, Teflon is insoluble in water, limestone is slightly soluble, and ethanol (ethyl alcohol) is infinitely so in that it can be mixed with water in any proportion ranging from pure water to pure ethanol with only a single liquid phase consisting of an ethanol/water solution being observed.



Whether a solid dissolves significantly in water may be determined in the laboratory by stirring the solid with water in a beaker.

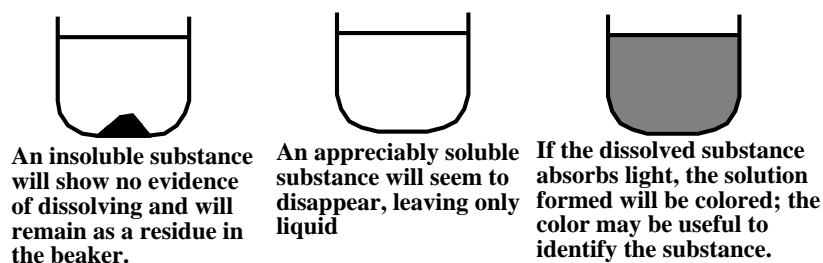


Figure 2.9 Illustration of whether a substance dissolves in a liquid.

### *Solution Concentration*

Solutions are addressed specifically and in more detail in Chapter 7. At this point, however, it is useful to have some understanding of the quantitative expression of the amount of solute dissolved in a specified amount of solution. In a relative sense, a solution that has comparatively little solute dissolved per unit volume of solution is called a **dilute solution**, whereas one with an amount of solute of the same order of magnitude as that of the solvent is a **concentrated solution**. Water saturated with air at 25°C contains only about 8 milligrams of oxygen dissolved in 1 liter of water; this composes a *dilute solution* of oxygen. A typical engine coolant solution contains about as much ethylene glycol (antifreeze) as it does water, so it is a *concentrated solution*. A solution that is at equilibrium with excess solute so that it contains the maximum amount of solute that it can dissolve is called a **saturated solution**. One that can still dissolve more solute is called an **unsaturated solution**.

For the chemist, the most useful way to express the concentration of a solution is in terms of the number of moles of solute dissolved per liter of solution. The **molar concentration,  $M$** , of a solution is *the number of moles of solute dissolved per liter of solution*. Mathematically,

$$M = \frac{\text{moles of solute}}{\text{number of liters of solution}} \quad (2.7.1)$$

Example: Exactly 34.0 g of ammonia,  $\text{NH}_3$ , were dissolved in water and the solution was made up to a volume of exactly 0.500 L. What was the molar concentration,  $M$ , of ammonia in the resulting solution?

Answer: The molar mass of  $\text{NH}_3$  is 17.0 g/mole. Therefore

$$\text{Number of moles of } \text{NH}_3 = \frac{34.0 \text{ g}}{17.0 \text{ g/mole}} = 2.00 \text{ mol} \quad (2.7.2)$$

The molar concentration of the solution is calculated from Equation 2.7.1:

$$M = \frac{2.00 \text{ mol}}{0.500 \text{ L}} = 4.00 \text{ mol/L} \quad (2.7.3)$$

Equations 2.7.1 and 2.7.2 are useful in doing calculations that relate mass of solute to solution volume and solution concentration in chemistry (see stoichiometry in Chapter 6).

## 2.8 SOLIDS

Matter in the form of *solids* is said to be in the **solid state**. The solid state is the most organized form of matter in that the atoms, molecules, and ions in it are in essentially fixed relative positions and are highly attracted to each other. Therefore, solids have a definite shape, maintain a constant volume, are virtually non-compressible under pressure, and expand and contract only slightly with changes in temperature. Like liquids, solids have generally very high densities relative to gases. (Some solids such as those made from styrofoam appear to have very low densities, but that is because such materials are composed mostly of bubbles of air in a solid matrix.) Because of the strong attraction of the atoms, molecules, and ions of solids for each other, solids do not enter the vapor phase very readily at all; the phenomenon by which this happens to a limited extent is called **sublimation**.

Some solids (quartz, sodium chloride) have very well defined geometric shapes and form characteristic crystals. These are called **crystalline solids** and occur because the molecules or ions of which they are composed assume well defined, specific positions relative to each other. Other solids have indefinite shapes because their constituents are arranged at random; glass is such a solid. These are **amorphous solids**.

## 2.9 THERMAL PROPERTIES

The behavior of a substance when heated or cooled defines several important physical properties of it, including the temperature at which it melts or vaporizes. Also included are the amounts of heat required by a given mass of the substance to raise its temperature by a unit of temperature, to melt it, or to vaporize it.

## Melting Point

The **melting point** of a pure substance is the temperature at which the substance changes from a solid to a liquid. At the melting temperature, pure solid and pure liquid composed of the substance may be present together in a state of equilibrium. An impure substance does not have a single melting temperature. Instead, as heat is applied to the substance, it begins to melt at a temperature below the melting temperature of the pure substance; melting proceeds as the temperature increases until no solid remains. Therefore, melting temperature measurements serve two purposes in characterizing a substance. The melting point of a pure substance is indicative of the identity of the substance. The melting behavior—whether melting occurs at a single temperature or over a temperature range—is a measure of substance purity.

## Boiling Point

Boiling occurs when a liquid is heated to a temperature such that bubbles of vapor of the substance are evolved. The boiling temperature depends upon pressure, and reduced pressures can be used to cause liquids to boil at lower temperatures. When the surface of a pure liquid substance is in contact with the pure vapor of the substance at 1 atm pressure, boiling occurs at a temperature called the **normal boiling point** (see [Figure 2.10](#)). Whereas a pure liquid remains at a constant temperature during boiling until it has all turned to vapor, the temperature of an impure liquid increases during boiling. As with melting point, the boiling point is useful for identifying liquids. The extent to which the boiling temperature is constant as the liquid is converted to vapor is a measure of the purity of the liquid.

## Specific Heat

As the temperature of a substance is raised, energy must be put into it to enable the molecules of the substance to move more rapidly relative to each other and to overcome the attractive forces between them. The amount of heat energy required to raise the temperature of a unit mass of a solid or liquid substance by a degree of temperature varies with the substance. Of the common liquids, the most heat is required to raise the temperature of water. This is because of the molecular structure of water, the molecules of which are strongly attracted to each other by electrical charges between the molecules (each of which has a relatively positive and a relatively more negative end). Furthermore, water molecules tend to be held in molecular networks by special kinds of chemical bonds called hydrogen bonds, in which hydrogen atoms form bridges between O atoms in different water molecules. The very high amount of energy required to increase the temperature of water has some important environmental implications because it stabilizes the temperatures of bodies of water and of geographic areas close to bodies of water. In contrast to water, hydrocarbon liquids such as those in gasoline require relatively little heat for warming because the molecules interact much less with each other than do those of water.

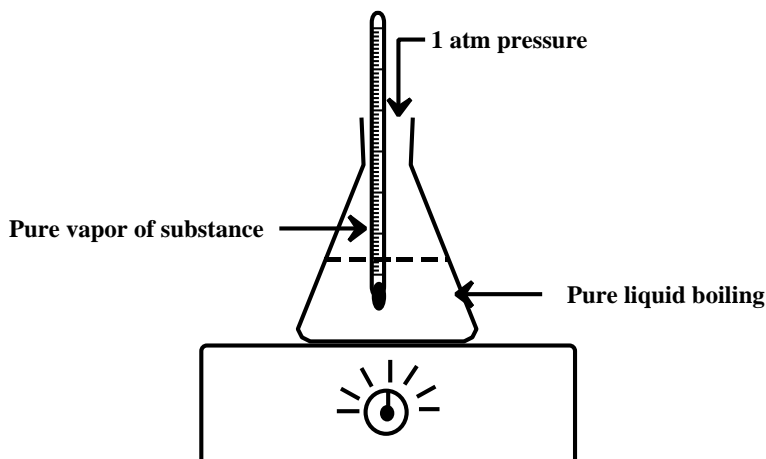


Figure 2.10 Illustration of normal boiling liquid.

The heat required to raise the temperature of a substance is called the **specific heat**, defined as the amount of heat energy required to raise the temperature of a gram of substance by 1 degree Celsius. It can be expressed by the equation

$$\text{Specific heat} = \frac{\text{heat energy absorbed, J}}{(\text{mass, g})(\text{increase in temperature, } ^\circ\text{C})} \quad (2.9.1)$$

where the heat energy is in units of joule, J. The most important value of specific heat is that of water, 4.18 J/g°C. From Equation 2.9.1, the amount of heat,  $q$ , required to raise the mass,  $m$ , of a particular substance over a temperature range of  $T$  is

$$q = (\text{specific heat}) \times m \times T \quad (2.9.2)$$

This equation can be used to calculate quantities of heat used in increasing water temperature or released when water temperature decreases. As an example, consider the amount of heat required to raise the temperature of 11.6 g of liquid water from 14.3°C to 21.8°C:

$$q = 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \times 11.6 \text{ g} \times (21.8^\circ\text{C} - 14.3^\circ\text{C}) = 364 \text{ J} \quad (2.9.2)$$

## Heat of Vaporization

Large amounts of heat energy may be required to convert liquid to vapor compared with the specific heat of the liquid. This is because changing a liquid to a gas requires breaking the molecules of liquid away from each other, not just increasing the rates at which they move relative to one another. The vaporization of a pure liquid occurs at a constant temperature (normal boiling temperature, see above). For

example, the temperature of a quantity of pure water is increased by heating to 100°C where it remains until all the water has evaporated. The **heat of vaporization** is the quantity of heat taken up in converting a unit mass of liquid entirely to vapor at a constant temperature. The heat of vaporization of water is 2,260 J/g (2.26 kJ/g) for water boiling at 100°C at 1 atm pressure. This amount of heat energy is about 540 times that required to raise the temperature of 1 gram of liquid water by 1°C. The fact that water's heat of vaporization is the largest of any common liquid has significant environmental effects. It means that enormous amounts of heat are required to produce water vapor from liquid water in bodies of water. This means that the temperature of a body of water is relatively stable when it is exposed to high temperatures because so much heat is dissipated when a fraction of the water evaporates. When water vapor condenses, similar enormous amounts of heat energy, called **heat of condensation**, are released. This occurs when water vapor forms precipitation in storm clouds and is the driving force behind the tremendous releases of heat that occur in thunderstorms and hurricanes.

The heat of vaporization of water can be used to calculate the heat required to evaporate a quantity of water. For example, the heat,  $q$ , required to evaporate 2.50 g of liquid water is

$$q = (\text{heat of vaporization})(\text{mass of water}) = 2.26 \text{ kJ/g} \times 2.50 \text{ g} = 5.65 \text{ kJ} \quad (2.9.3)$$

and that released when 5.00 g of water vapor condenses is

$$\begin{aligned} q &= -(\text{heat of vaporization})(\text{mass of water}) \\ &= -2.26 \text{ kJ/g} \times 5.00 \text{ g} = -11.3 \text{ kJ} \end{aligned} \quad (2.9.4)$$

the latter value is negative to express the fact that heat is *released*.

## Heat of Fusion

**Heat of fusion** is the quantity of heat taken up in converting a unit mass of solid entirely to liquid at a constant temperature. The heat of fusion of water is 330 J/g for ice melting at 0°C. This amount of heat energy is 80 times that required to raise the temperature of 1 gram of liquid water by 1°C. When liquid water freezes, an exactly equal amount of energy per unit mass is released, but is denoted as a negative value.

The heat of fusion of water can be used to calculate the heat required to melt a quantity of ice. For example, the heat,  $q$ , required to melt 2.50 g of ice is

$$q = (\text{heat of fusion})(\text{mass of water}) = 330 \text{ J/g} \times 2.50 \text{ g} = 825 \text{ J} \quad (2.9.5)$$

and that released when 5.00 g of liquid water freezes is

$$q = -(\text{heat of fusion})(\text{mass of water}) = -330 \text{ J/g} \times 5.00 \text{ g} = -1,650 \text{ J} \quad (2.9.6)$$

## 2.10 SEPARATION AND CHARACTERIZATION OF MATTER

A very important aspect of the understanding of matter is the separation of mixtures of matter into their constituent pure substances. Consider, the treatment of

an uncharacterized waste sludge found improperly disposed of in barrels. Typically, the sludge will appear as a black goo with an obnoxious odor from several constituents. In order to do a proper analysis of this material, it often needs to be separated into its constituent components, such as by extracting water-soluble substances into water, extracting organic materials into an organic solvent, and distilling off volatile constituents. The materials thus isolated can be analyzed to determine what is present. With this knowledge and with information about the separation itself a strategy can be devised to treat the waste in an effective, economical manner.

Many kinds of processes are used for separations, and it is beyond the scope of this chapter to go into detail about them here. However, several of the more important separation operations are discussed briefly below.

## Distillation

**Distillation** consists of evaporating a liquid by heating and cooling the vapor so that it condenses back to the liquid in a different container, as shown for water in [Figure 2.11](#). Less-volatile constituents such as solids are left behind in the distillation flask; more-volatile impurities such as volatile organic compound pollutants in water will distill off first and can be placed in different containers before the major liquid constituent is collected.

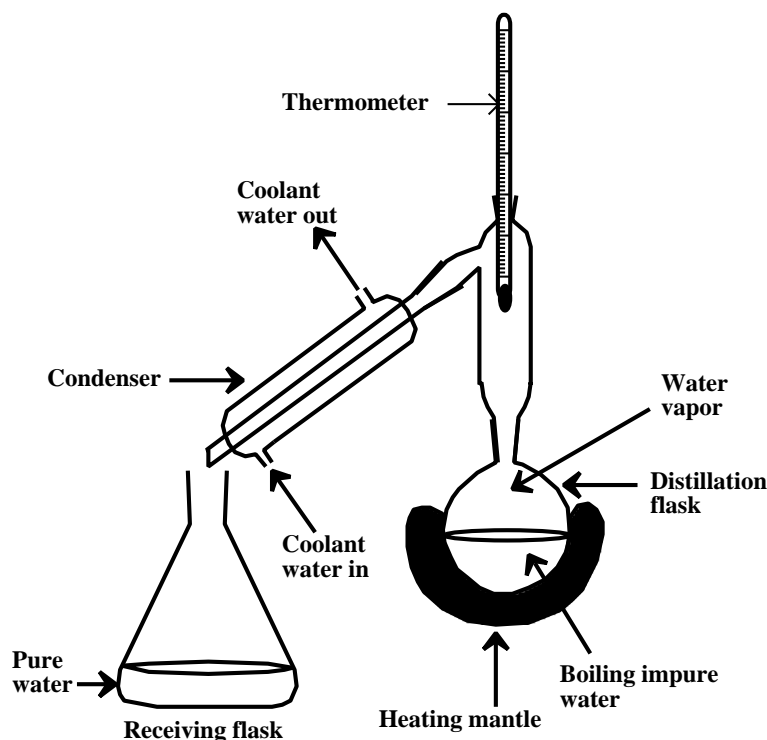


Figure 2.11 Separation by distillation.

The apparatus shown in [Figure 2.11](#) could be used, for example, to prepare fresh water from seawater, which can be considered a solution of sodium chloride in



water. The seawater to be purified is placed in a round-bottom distillation flask heated with an electrically powered heating mantle. The seawater is boiled and pure water vapor in the gaseous state flows into the condenser where it is cooled and condensed back to a purified salt-free product in the receiving flask. When most of the seawater has been evaporated, the distillation flask and its contents are cooled and part of the sodium chloride separates out as crystals that can be removed. Sophisticated versions of this distillation process are used in some arid regions of the world to produce potable (drinking) water from seawater.

Distillation is widely used in the petroleum and chemical industries. A sophisticated distillation apparatus is used to separate the numerous organic components of crude oil, including petroleum ether, gasoline, kerosene, and jet fuel fractions. This requires **fractional distillation** in which part of the vapor recondenses in a **fractionating column** that would be mounted vertically on top of the distillation flask in the apparatus shown in [Figure 2.11](#). The most volatile components enter into the condenser and are condensed back to liquid products; less volatile liquid constituents return to the distillation flask. The net effect is that of numerous distillations, which gives a very efficient separation.

In some cases the residues left from distillation, **distillation bottoms** (“still bottoms”) are waste materials. Several important categories of hazardous wastes consist of distillation bottoms, which pose disposal problems.

## Separation in Waste Treatment

In addition to distillation, several other kinds of separation procedures are used for waste treatment. These are addressed in more detail in Chapter 22 and are mentioned here as examples of separation processes.

### *Phase Transitions*

Distillation is an example of a separation in which one of the constituents being separated undergoes a **phase transition** from one phase to another, then often back again to the same phase. In distillation, for example, a liquid is converted to the vapor state, then recondensed as a liquid. Several other kinds of separations by phase transition as applied to waste treatment are the following:

**Evaporation** is usually employed to remove water from an aqueous waste to concentrate it. A special case of this technique is **thin-film evaporation**, in which volatile constituents are removed by heating a thin layer of liquid or sludge waste spread on a heated surface.

**Drying**—removal of solvent or water from a solid or semisolid (sludge) or the removal of solvent from a liquid or suspension—is a very important operation because water is often the major constituent of waste products, such as sludges. In **freeze drying**, the solvent, usually water, is sublimed from a frozen material. Hazardous-waste solids and sludges are dried to reduce the quantity of waste, to remove solvent or water that might interfere with subsequent treatment processes, and to remove hazardous volatile constituents.

**Stripping** is a means of separating volatile components from less-volatile ones in a liquid mixture by the partitioning of the more-volatile materials to a gas phase of

air or steam (steam stripping). The gas phase is introduced into the aqueous solution or suspension containing the waste in a stripping tower that is equipped with trays or packed to provide maximum turbulence and contact between the liquid and gas phases. The two major products are condensed vapor and a stripped bottoms residue. Examples of two volatile components that can be removed from water by air stripping are the organic solvents benzene and dichloromethane.

**Physical precipitation** is used here as a term to describe processes in which a dissolved solute comes out of solution as a solid as a result of a physical change in the solution. The major changes that can cause physical precipitation are cooling the solution, evaporation of solvent, or alteration of solvent composition. The most common type of physical precipitation by alteration of solvent composition occurs when a water-miscible organic solvent is added to an aqueous (water) solution of a salt, so that the solubility of a salt is lowered below its concentration in the solution.

### *Phase Transfer*

**Phase transfer** consists of the transfer of a solute in a mixture from one phase to another. An important type of phase transfer process is **solvent extraction**, shown in [Figure 2.12](#), a process in which a substance is transferred from solution in one solvent (usually water) to another (usually an organic solvent) without any chemical change taking place. When solvents are used to leach substances from solids or sludges, the process is called **leaching**. One of the more promising approaches to solvent extraction and leaching of hazardous wastes is the use of supercritical fluids, most commonly CO<sub>2</sub>, as extraction solvents. A supercritical fluid is one that has char-

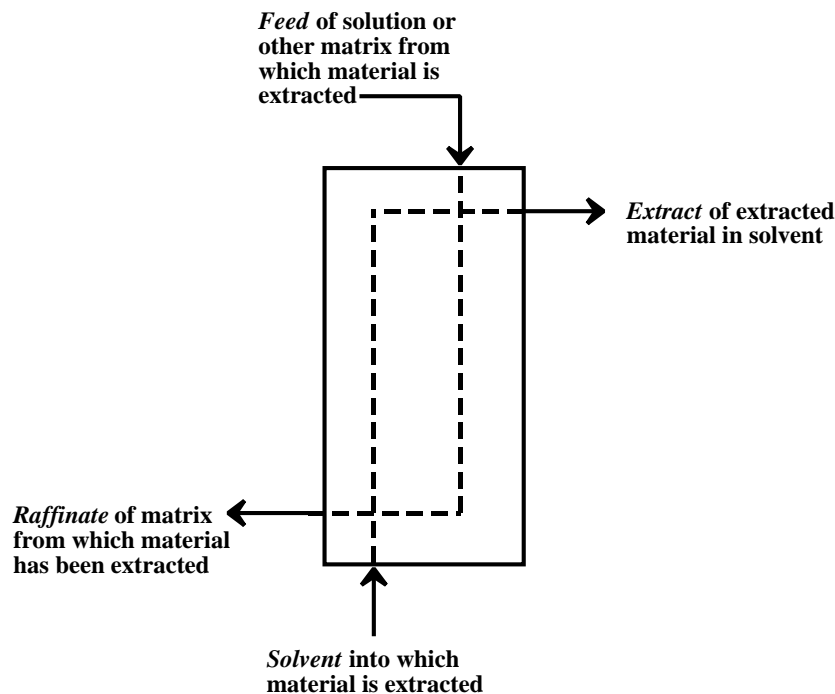


Figure 2.12 Outline of solvent extraction/leaching process with important terms in italics.

acteristics of both liquid and gas and consists of a substance above its supercritical temperature and pressure (31.1°C and 73.8 atm, respectively, for CO<sub>2</sub>). After a substance has been extracted from a waste into a supercritical fluid at high pressure, the pressure can be released, resulting in separation of the substance extracted. The fluid can then be compressed again and recirculated through the extraction system. Some possibilities for treatment of hazardous wastes by extraction with supercritical CO<sub>2</sub> include removal of organic contaminants from wastewater, extraction of organohalide pesticides from soil, extraction of oil from emulsions used in aluminum and steel processing, and regeneration of spent activated carbon.

Transfer of a substance from a solution to a solid phase is called **sorption**. The most important sorbent used in waste treatment is **activated carbon**. Activated carbon can also be applied to pretreatment of waste streams going into other processes to improve treatment efficiency and reduce fouling. Activated carbon sorption is most effective for removing from water organic materials that are poorly water soluble and that have high molecular masses.

### *Molecular Separation*

A third major class of physical separation is **molecular separation**, often based upon **membrane processes** in which dissolved contaminants or solvent pass through a size-selective membrane under pressure. **Reverse osmosis** is the most widely used of the membrane techniques. It operates by virtue of a membrane that is selectively permeable to water and excludes ionic solutes. Reverse osmosis uses high pressures to force permeate through the membrane, producing a concentrate containing high levels of dissolved salts.

## **CHAPTER SUMMARY**

*The chapter summary below is presented in a programmed format to review the main points covered in this chapter. It is used most effectively by filling in the blanks, referring back to the chapter as necessary. The correct answers are given at the end of the summary.*

Matter is defined as <sup>1</sup> \_\_\_\_\_  
\_\_\_\_\_. Chemical properties of matter  
are those that are observed <sup>2</sup> \_\_\_\_\_  
\_\_\_\_\_, whereas physical properties are those that  
are observed <sup>3</sup> \_\_\_\_\_  
\_\_\_\_\_. Matter consists of only a few <sup>4</sup> \_\_\_\_\_ and a large  
number of substances formed from them called <sup>5</sup> \_\_\_\_\_. Some general  
properties of metals are <sup>6</sup> \_\_\_\_\_  
\_\_\_\_\_. Some  
common properties of nonmetals are <sup>7</sup> \_\_\_\_\_  
\_\_\_\_\_.  
<sup>8</sup> \_\_\_\_\_ substances consist of virtually all compounds that  
contain carbon, whereas <sup>9</sup> \_\_\_\_\_ substances make up virtually all

matter that does not contain carbon. Matter consisting of <sup>10</sup> \_\_\_\_\_ is a pure substance; any other form of matter is a <sup>11</sup> \_\_\_\_\_ of two or more substances. A <sup>12</sup> \_\_\_\_\_ is not uniform throughout and possesses readily distinguishable constituents, whereas <sup>13</sup> \_\_\_\_\_ are uniform throughout. A mole is defined as the quantity of substance that <sup>14</sup> \_\_\_\_\_.

A mole of H<sub>2</sub>O contains <sup>15</sup> \_\_\_\_\_ grams of this compound. Avogadro's number is the number of specified entities <sup>16</sup> \_\_\_\_\_ and it has a value of <sup>17</sup> \_\_\_\_\_. A general term that includes atoms and molecules as well as groups of ions making up the smallest possible quantity of an ionic compound is the <sup>18</sup> \_\_\_\_\_. The average mass of a formula unit is called the <sup>19</sup> \_\_\_\_\_. Density is defined as <sup>20</sup> \_\_\_\_\_ and expressed by the formula <sup>21</sup> \_\_\_\_\_. Relative densities are expressed by means of a number called <sup>22</sup> \_\_\_\_\_ defined as <sup>23</sup> \_\_\_\_\_.

\_\_\_\_\_ . Some solutions are colored because they <sup>24</sup> \_\_\_\_\_. Solids, liquids, and gases are the three <sup>25</sup> \_\_\_\_\_ of matter. A solid has a definite <sup>26</sup> \_\_\_\_\_ and <sup>27</sup> \_\_\_\_\_ regardless of the container in which it is placed, a liquid has a definite <sup>28</sup> \_\_\_\_\_, but takes on the <sup>29</sup> \_\_\_\_\_ of its container, and a gas has the same <sup>30</sup> \_\_\_\_\_ as its container. Gas molecules move <sup>31</sup> \_\_\_\_\_, colliding with container walls to exert <sup>32</sup> \_\_\_\_\_. A quantity of a gas has a <sup>33</sup> \_\_\_\_\_ compared with the same amount of material in a liquid or solid. <sup>34</sup> \_\_\_\_\_ enables gas molecules to move large distances from their sources. The gas laws relate the four parameters of <sup>35</sup> \_\_\_\_\_ of gas. Boyle's law states that <sup>36</sup> \_\_\_\_\_.

\_\_\_\_\_ . Charles' law states that <sup>37</sup> \_\_\_\_\_.

\_\_\_\_\_ . Avogadro's law states that <sup>38</sup> \_\_\_\_\_.

\_\_\_\_\_ . The general gas law states that <sup>39</sup> \_\_\_\_\_.

\_\_\_\_\_ . The ideal gas law is given mathematically as <sup>40</sup> \_\_\_\_\_. At a temperature of 0°C and 1 atm pressure known as <sup>41</sup> \_\_\_\_\_ the volume of 1 mole of ideal gas is <sup>42</sup> \_\_\_\_\_. Molecules of liquids move <sup>43</sup> \_\_\_\_\_ relative to each other. Because there is little unoccupied space between molecules of liquid, liquids are only slightly <sup>44</sup> \_\_\_\_\_. The phenomenon by which molecules of liquid enter the gas phase is called <sup>45</sup> \_\_\_\_\_, whereas the opposite phenomenon is termed <sup>46</sup> \_\_\_\_\_. A steady-state level of the vapor of a liquid above the liquid surface is called the <sup>47</sup> \_\_\_\_\_ of the liquid. Solutions are regarded as liquids in which <sup>48</sup> \_\_\_\_\_.

The predominant liquid constituent of a solution is called the <sup>49</sup> \_\_\_\_\_, a substance dispersed in it is said to be <sup>50</sup> \_\_\_\_\_ in the solvent and is called the <sup>51</sup> \_\_\_\_\_. A substance that dissolves to a significant extent is said to be <sup>52</sup> \_\_\_\_\_ in the solvent. Substances that do not dissolve to a perceptible degree are classified as <sup>53</sup> \_\_\_\_\_. A solution that has comparatively little solute dissolved per unit volume of solution is called a <sup>54</sup> \_\_\_\_\_ solution, whereas one with an amount of solute of the same order of magnitude as that of the solvent is a <sup>55</sup> \_\_\_\_\_ solution. The molar concentration,  $M$ , of a solution is defined as <sup>56</sup> \_\_\_\_\_ and is given by formula <sup>57</sup> \_\_\_\_\_. The solid state is the most <sup>58</sup> \_\_\_\_\_ form of matter because the atoms, molecules, and ions in it are <sup>59</sup> \_\_\_\_\_. The phenomenon by which solids enter the gaseous state directly is called <sup>60</sup> \_\_\_\_\_. Solids that form well-defined crystals are called <sup>61</sup> \_\_\_\_\_, and those that have indefinite shapes because their constituents are arranged at random are called <sup>62</sup> \_\_\_\_\_. The melting point of a pure substance is the temperature at which the substance changes <sup>63</sup> \_\_\_\_\_. At this temperature, pure solid and pure liquid composed of the substance <sup>64</sup> \_\_\_\_\_. A(n) impure substance does not have <sup>65</sup> \_\_\_\_\_. The two purposes served by melting temperature measurements in characterizing a substance are <sup>66</sup> \_\_\_\_\_ and <sup>67</sup> \_\_\_\_\_.

\_\_\_\_\_ . When a pure liquid is heated such that the surface of the liquid is in contact with the pure vapor of the substance at 1 atm pressure, boiling occurs at a temperature called the <sup>68</sup> \_\_\_\_\_. The extent to which the boiling temperature is constant as the liquid is converted to vapor is a measure <sup>69</sup> \_\_\_\_\_. The amount of heat energy required to raise the temperature of a unit mass of a solid or liquid substance by a degree of temperature is called the <sup>70</sup> \_\_\_\_\_ of the substance and can be expressed by the equation <sup>71</sup> \_\_\_\_\_. The heat of vaporization is defined as <sup>72</sup> \_\_\_\_\_.

Mathematically, the heat,  $q$ , required to evaporate a specified mass of liquid water is <sup>73</sup> \_\_\_\_\_. Heat of fusion is defined as <sup>74</sup> \_\_\_\_\_.

Distillation consists of <sup>75</sup> \_\_\_\_\_.

Fractional distillation is a separation process in which part of the vapor <sup>76</sup> \_\_\_\_\_. Residues left from distillation are called <sup>77</sup> \_\_\_\_\_. A process in which a solvent is sublimed from a frozen material is called <sup>78</sup> \_\_\_\_\_. Stripping is a means of separating volatile components from less-volatile ones in a liquid mixture by <sup>79</sup> \_\_\_\_\_.

\_\_\_\_\_ . Solvent  
extraction is a process in which <sup>80</sup> \_\_\_\_\_

\_\_\_\_\_ .  
The most important sorbent used in waste treatment is <sup>81</sup> \_\_\_\_\_. A  
water purification process called <sup>82</sup> \_\_\_\_\_ uses high  
pressures to force permeate through a membrane that is selectively permeable to  
water and excludes ionic solutes.

### *Answers to Chapter Summary*

1. anything that has mass and occupies space
2. by chemical change or the potential for chemical change
3. without any chemical change
4. elements
5. compounds
6. generally solid, shiny in appearance, electrically conducting, and malleable
7. a dull appearance, not at all malleable, and frequently present as gases or liquids
8. Organic
9. inorganic substances
10. only one compound or of only one form of an element
11. mixture
12. heterogeneous mixture
13. homogeneous mixtures
14. contains the same number of specified entities as there are atoms of C in exactly 0.012 kg (12 g) of carbon-12
15. 18
16. in a mole of substance
17.  $6.02 \times 10^{23}$
18. formula unit
19. formula mass
20. mass per unit volume
21.  $d = \text{mass/volume}$
22. specific gravity
23. the ratio of the density of a substance to that of a standard substance
24. absorb light
25. states
26. shape
27. volume
28. volume
29. shape
30. shape and volume
31. independently and at random
32. pressure
33. high volume
34. Diffusion
35. moles (n), volume (V), temperature (T), and pressure (P)
36. at constant temperature, the volume of a fixed quantity of gas is inversely

- proportional to the pressure of the gas
37. the volume of a fixed quantity of gas is directly proportional to the absolute temperature at constant pressure
  38. at constant temperature and pressure the volume of a gas is directly proportional to the number of molecules of gas, commonly expressed as moles
  39. the volume of a quantity of ideal gas is proportional to the number of moles of gas and its absolute temperature and inversely proportional to its pressure
  40.  $PV = nRT$
  41. standard temperature and pressure (STP)
  42. 22.4 L
  43. freely
  44. compressible
  45. evaporation
  46. condensation
  47. vapor pressure
  48. quantities of gases, solids, or other liquids are dispersed as individual ions or molecules
  49. solvent
  50. dissolved
  51. solute
  52. soluble
  53. insoluble
  54. dilute
  55. concentrated
  56. the number of moles of solute dissolved per liter of solution
  57.  $M = \frac{\text{moles of solute}}{\text{number of liters of solution}}$
  58. organized
  59. in essentially fixed relative positions and are highly attracted to each other
  60. sublimation
  61. crystalline solids
  62. amorphous solids
  63. from a solid to a liquid
  64. may be present together in a state of equilibrium
  65. a single melting temperature
  66. the melting point of a pure substance is indicative of the identity of the substance
  67. the melting behavior—whether melting occurs at a single temperature or over a temperature range—is a measure of substance purity
  68. normal boiling point
  69. of the purity of the liquid
  70. specific heat
  71. Specific heat =  $\frac{\text{heat energy absorbed, J}}{(\text{mass, g})(\text{increase in temperature, } ^\circ\text{C})}$
  72. the quantity of heat taken up in converting a unit mass of liquid entirely to vapor at a constant temperature
  73.  $q = (\text{heat of vaporization})(\text{mass of water})$
  74. the quantity of heat taken up in converting a unit mass of solid entirely to liquid

- at a constant temperature
75. evaporating a liquid by heating and cooling the vapor so that it condenses back to the liquid in a different container
  76. recondenses in a fractionating column
  77. distillation bottoms
  78. freeze drying
  79. the partitioning of the more-volatile materials to a gas phase of air or steam
  80. a substance is transferred from solution in one solvent to another without any chemical change taking place
  81. activated carbon
  82. Reverse osmosis

## QUESTIONS AND PROBLEMS

### *Section 2.1. What is Matter*

1. What is the definition of matter?

### *Section 2.2. Classification of Matter*

2. Classify each of the following as chemical processes (c) or physical processes (p):
  - ( ) Ice melts
  - ( )  $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$
  - ( ) A small silver sphere pounded into a thin sheet
  - ( ) A copper-covered roof turns green in the atmosphere
  - ( ) Wood burns
  - ( ) Oxygen isolated by the distillation of liquid air
3. List (a) three common characteristics of metals and (b) three common characteristics of nonmetals.
4. Classify each of the following as organic substances (o) or inorganic substances (i):
  - ( ) Crude oil
  - ( ) A dime coin
  - ( ) A shirt
  - ( ) Asphalt isolated as a residue of petroleum distillation
  - ( ) A compound composed of silicon, oxygen, and aluminum
  - ( ) Concrete
5. Sandy sediment from the bottom of a river was filtered to isolate a clear liquid, which was distilled to give a liquid compound that could not be broken down further by non-chemical means. Explain how these operations illustrate heterogeneous mixtures, homogeneous mixtures, and pure substances.
6. Explain why a solution is defined as a homogeneous mixture.



Section 2.3. *Quantity of Matter: The Mole*

7. Justify the statement that “exactly 12 g of carbon-12 contain  $6.02 \times 10^{23}$  atoms of the isotope of carbon that contains 6 protons and 6 neutrons in its nucleus.” What two terms relating to quantity of matter are illustrated by this statement?
8. Where X is a number equal to the formula mass of a substance, complete the following formula: Mass of a mole of a substance = \_\_\_\_\_.
9. Fill in the blanks in the table below:

Substance	Formula unit	Formula mass	Mass of 1 mole	Numbers and kinds of individual entities in 1 mole
Neon	Ne atom	_____	_____ g	_____ Ne atoms
Nitrogen gas	N <sub>2</sub> molecule	_____	_____ g	_____ N <sub>2</sub> molecules _____ N atoms
Ethene	C <sub>2</sub> H <sub>4</sub> molecules	_____	_____ g	_____ C <sub>2</sub> H <sub>4</sub> molecules _____ C atoms _____ H atoms

Section 2.4. *Physical Properties of Matter*

10. Calculate the density of each of the following substances, for which the mass of a specified volume is given:
  - (a) 93.6 g occupying 15 mL
  - (b) 0.992 g occupying 1,005 mL
  - (c) 13.7 g occupying 11.4 mL
  - (d) 16.8 g occupying 3.19 mL
11. Calculate the volume in mL occupied by 156.3 g of water at 4°C.
12. From densities given in [Table 2.3](#), calculate each of the following:
  - (a) Volume of 105 g of iron
  - (b) Mass of 157 cm<sup>3</sup> of benzene
  - (c) Volume of 1,932 g of air
  - (d) Mass of 0.525 cm<sup>3</sup> of carbon tetrachloride
  - (e) Volume of 10.0 g of helium
  - (f) Mass of 13.2 cm<sup>3</sup> of sugar
13. From information given in [Table 2.3](#), calculate the specific gravities of (a) benzene, (b) carbon tetrachloride, and (c) sulfuric acid relative to water at 4°C.
14. From information given in this chapter, calculate the density of pure nitrogen gas, N<sub>2</sub>, at STP. To do so will require some knowledge of the mole and of the gas laws.

15. Different colors are due to different \_\_\_\_\_ of light.

*Section 2.5. States of Matter*

16. Explain how the various forms of water observed in the environment illustrate the three states of matter.

*Section 2.6. Gases*

17. Match each gas law stated on the left, below, with the phenomenon that it explains on the right:

- |                    |  |
|--------------------|--|
| A. Charles' law    | 1. The volume of a mole of ideal gas at STP is 22.4 L.   |
| B. General gas law | 2. Tripling the number of moles of gas at constant temperature and pressure triples the volume.                                    |
| C. Boyle's law     | 3. Raising the absolute temperature of a quantity of gas at constant pressure from 300 K to 450 K increases the gas volume by 50%. |
| D. Avogadro's law  | 4. Doubling the pressure on a quantity of gas at constant temperature halves the volume.   |

18. How might an individual molecule of a gas describe its surroundings? What might it see, feel, and do?

19. Explain pressure and diffusion in terms of the behavior of gas molecules.

20. Use the gas law to calculate whether the volume of a mole of water vapor at 100°C and 1 atm pressure is 30,600 mL as stated in this chapter.

21. Using the appropriate gas laws, calculate the quantities denoted by the blanks in the table below:

Conditions before				Conditions after			
$n_1$ , mol	$V_1$ , L	$P_1$ , atm	$T_1$ , K	$n_2$ , mol	$V_2$ , L	$P_2$ , atm	$T_2$ , K
1.25	13.2	1.27	298	1.25	(a) _____	0.987	298
1.25	13.2	1.27	298	1.25	(b) _____	1.27	407
1.25	13.2	1.27	298	4.00	(c) _____	1.27	298
1.25	13.2	1.27	298	1.36	(d) _____	1.06	372
1.25	13.2	1.27	298	1.25	30.5	(e) _____	298

*Chapter 2.7. Liquids and Solutions*

22. In terms of distances between molecules and movement of molecules relative to each other, distinguish between liquids and gases and liquids and solids. Compare and explain shape and compressibility of gases, liquids, and solids.

23. Define (a) evaporation, (b) condensation, (c) vapor pressure as related to liquids.

24. Define (a) solvent, (b) solute, (c) solubility, (d) saturated solution, (e)

- unsaturated solution, (f) insoluble solute.
25. From information given in Section 2.7, calculate the molar concentration of  $O_2$  in water saturated with air at  $25^\circ C$
  26. Calculate the molar concentration of solute in each of the following solutions:
    - (a) Exactly 7.59 g of  $NH_3$  dissolved in 1.500 L of solution.
    - (b) Exactly 0.291 g of  $H_2SO_4$  dissolved in 8.65 L of solution.
    - (c) Exactly 85.2 g of NaOH dissolved in 5.00 L of solution.
    - (d) Exactly 31.25 g of NaCl dissolved in 2.00 L of solution.
    - (e) Exactly 6.00 mg of atmospheric  $N_2$  dissolved in water in a total solution volume of 2.00 L.
    - (f) Exactly 1,000 g of ethylene glycol,  $C_2H_6O_2$  dissolved in water and made up to a volume of 2.00 L.

### *Section 2.8. Solids*

27. Why is the solid state called the most organized form of matter?
28. As related to solids, define (a) sublimation, (b) crystalline solids, and (c) amorphous solids.

### *Section 2.9. Thermal Properties*

29. Experimentally, how could a chemist be sure that (a) a melting point temperature was being observed, (b) the melting point observed was that of a pure substance, and (c) the melting point observed was that of a mixture.
30. Calculate the specific heat of substances for which
  - (a) 22.5 J of heat energy was required to raise the temperature of 2.60 g of the substance from  $11.5$  to  $16.2^\circ C$
  - (b) 45.6 J of heat energy was required to raise the temperature of 1.90 g of the substance from  $8.6$  to  $19.3^\circ C$
  - (c) 1.20 J of heat energy was required to raise the temperature of 3.00 g of the substance from  $25.013$  to  $25.362^\circ C$
31. Calculate the amount of heat required to raise the temperature of 7.25 g of liquid water from  $22.7^\circ C$  to  $29.2^\circ C$ .
32. Distinguish between specific heat and (a) heat of vaporization and (b) heat of fusion.
33. Calculate the heat,  $q$ , (a) required to evaporate 5.20 g of liquid water and (b) released when 6.50 g of water vapor condenses.
34. Calculate the heat,  $q$ , (a) required to melt 5.20 g of ice and (b) released when

6.50 g of liquid water freezes.

*Section 2.10. Separation and Characterization of Matter*

35. Suppose that salt were removed from seawater by distillation. In reference to the apparatus shown in [Figure 2.11](#), explain where the pure water product would be isolated and where the salt would be found.
36. How does fractional distillation differ from simple distillation?
37. What are distillation bottoms and what do they have to do in some cases with wastes?
38. Describe solvent extraction as an example of a phase transfer process.
39. For what kind of treatment process is activated carbon used?
40. Name and describe a purification process that operates by virtue of “a membrane that is selectively permeable to water and excludes ionic solutes.”