CHAPTER 12
Solutions

BIG IDEA
Solutions are homogeneous mixtures of a solute in a solvent. Chemical and physical factors affect how much of one substance will dissolve in another.

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Solutions

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Types of Mixtures

Key Terms

soluble   solute   electrolyte
solution   suspension   nonelectrolyte
solvent   colloid

SECTION 1

Main Ideas

▶ Solutions are homogeneous mixtures.

▶ The particles in a suspension are large.

▶ Colloids have particles of intermediate size.

▶ Electrolytes are ionic solutions that conduct electricity.

It is easy to determine that some materials are mixtures because you can see their component parts. For example, soil is a mixture of substances, including small rocks and decomposed animal and plant matter. You can see this by picking up some soil in your hand and looking at it closely. Milk, on the other hand, does not appear to be a mixture, but in fact it is. Milk is composed principally of fats, proteins, milk sugar, and water. If you look at milk under a microscope, it will look something like **Figure 1.1a.** You can see round lipid (fat) droplets that measure from 1 to 10 μm in diameter. Irregularly shaped casein (protein) particles that are about 0.2 μm wide can also be seen. Both milk and soil are examples of *heterogeneous mixtures* because their composition is not uniform.

Salt (sodium chloride) and water form a *homogeneous mixture*. The sodium and chloride ions are interspersed among the water molecules, and the mixture appears uniform throughout, as illustrated in **Figure 1.1b.**
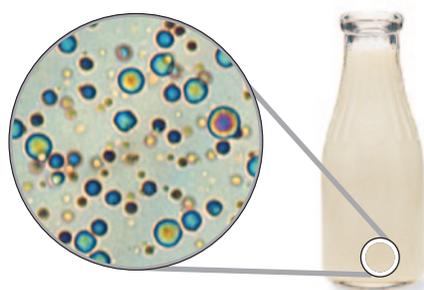
▶ MAIN IDEA

Solutions are homogeneous mixtures.

Suppose a sugar cube is dropped into a glass of water. You know from experience that the sugar will dissolve. Sugar is described as "soluble in water." **By soluble, we mean capable of being dissolved.** As it dissolves, a sugar lump gradually disappears as sugar molecules leave the surface of their crystals and mix with water molecules. Eventually, all the sugar molecules become uniformly distributed among the water molecules, as indicated by the equally sweet taste of any part of the mixture. All visible traces of the solid sugar are gone. Such a mixture is called a solution. **A solution is a homogeneous mixture of two or more substances uniformly dispersed throughout a single phase.**

FIGURE 1.1

Heterogeneous and Homogeneous Mixtures

(bl) ©Eye of Science/SPL/Photo Researchers, Inc

Sodium ion, Na$^+$

Water molecule

Chloride ion, Cl$^-$

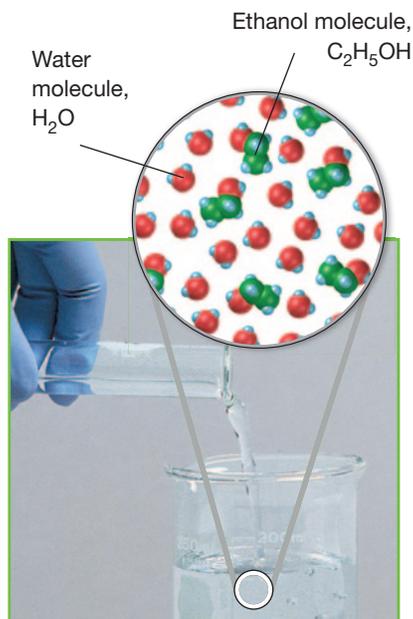**(a)** Milk is a heterogeneous mixture that consists of visible particles in a nonuniform arrangement.
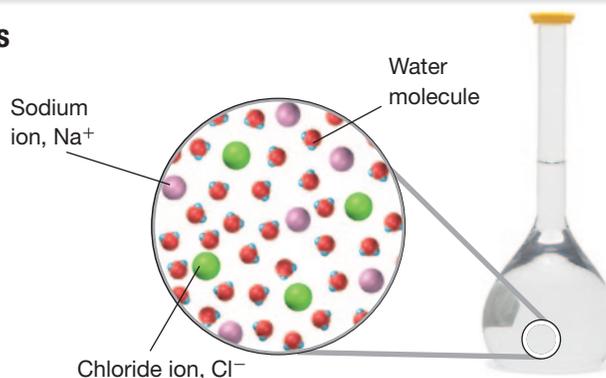
(b) Salt water is an example of a homogeneous mixture. Ions and water molecules are in a random arrangement.
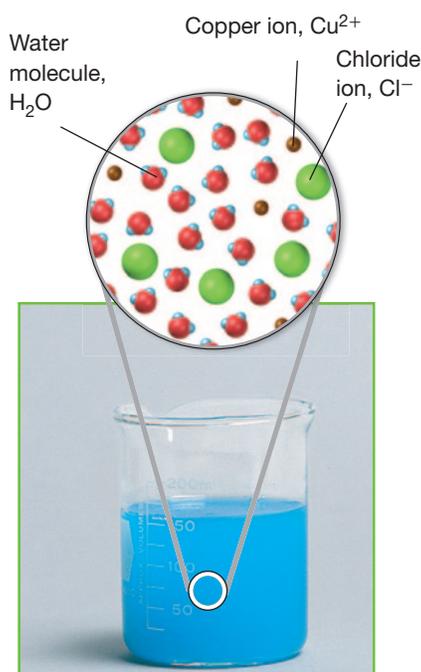
FIGURE 1.2

Solutes and Solvents The solute in a solution can be a solid, liquid, or gas.

Ethanol molecule, C_2H_5OH

Water molecule, H_2O

(a) The ethanol-water solution is made from a liquid solute in a liquid solvent.

Copper ion, Cu^{2+}

Water molecule, H_2O

Chloride ion, Cl^-

(b) The copper(II) chloride–water solution is made from a solid solute in a liquid solvent. Note that the composition of each solution is uniform.

Components of Solutions

In a solution, atoms, molecules, or ions are thoroughly mixed, resulting in a mixture that has the same composition and properties throughout. In the simplest type of solution, such as a sugar-water solution, the particles of one substance are randomly mixed with the particles of another substance. **The dissolving medium in a solution is called the solvent, and the substance dissolved in a solution is called the solute.** The solute is generally designated as that component of a solution that is of lesser quantity. In the ethanol-water solution shown in **Figure 1.2,** ethanol is the solute, and water is the solvent. Occasionally, these terms have little meaning. For example, in a 50%-50% solution of ethanol and water, it would be difficult and unnecessary to say which is the solvent and which the solute.

In a solution, the dissolved solute particles are so small that they cannot be seen. They remain mixed with the solvent indefinitely, as long as the existing conditions remain unchanged. If the solutions in **Figure 1.2** are poured through filter paper, both the solute and the solvent will pass through the paper. The solute-particle dimensions are those of atoms, molecules, and ions—which range from about 0.01 to 1 nm in diameter.

Types of Solutions

Solutions may exist as gases, liquids, or solids. Some possible solute-solvent combinations of gases, liquids, and solids in solutions are summarized in **Figure 1.3.** Note each has a defined solvent and solute.

Many alloys, such as brass (made from zinc and copper) and sterling silver (made from silver and copper), are solid solutions in which the atoms of two or more metals are uniformly mixed. By properly choosing the proportions of each metal in the alloy, many desirable properties can be obtained. For example, alloys can have more strength and greater resistance to corrosion than the pure metals. Pure gold (24K), for instance, is too soft to use in jewelry. Alloying it with silver and copper greatly increases its strength and hardness while retaining its appearance and corrosion resistance. **Figure 1.4** (on the next page) compares solutions of pure gold and a gold alloy. 14-karat gold is a solution because the gold, silver, and copper are uniformly mixed at the atomic level.

FIGURE 1.3

SOME SOLUTE-SOLVENT COMBINATIONS FOR SOLUTIONS

Solute state	Solvent state	Example
gas	gas	oxygen in nitrogen
gas	liquid	carbon dioxide in water
liquid	liquid	alcohol in water
liquid	solid	mercury in silver and tin (dental amalgam)
solid	liquid	sugar in water
solid	solid	copper in nickel (Monel™ alloy)

FIGURE 1.4

Pure Gold vs. Gold Alloy Pure gold is too soft to use in jewelry. Alloying it with other metals increases its strength and hardness while retaining its appearance and corrosion resistance.
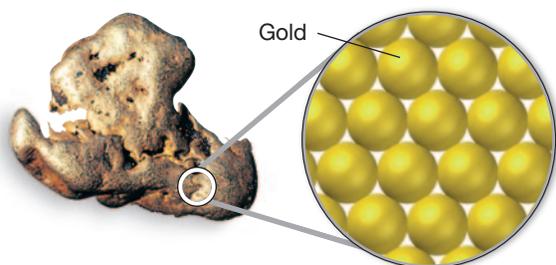
Gold

(a) 24-karat gold is pure gold.
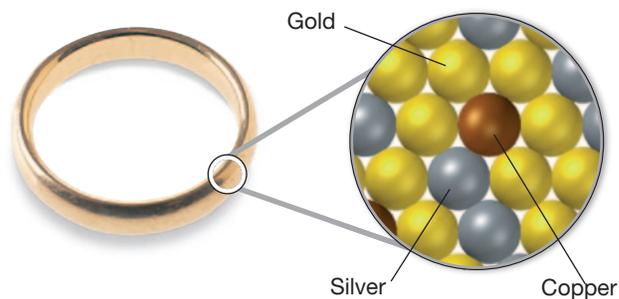
Gold

Silver   Copper

(b) 14-karat gold is an alloy of gold with silver and copper. 14-karat gold is 14/24, or 58.3%, gold.

▶ **MAIN IDEA**

The particles in a suspension are large.

If the particles in a solvent are so large that they settle out unless the mixture is constantly stirred or agitated, the mixture is called a suspension. Think of a jar of muddy water. If left undisturbed, particles of soil collect on the bottom of the jar. The soil particles are denser than the solvent, water. Gravity pulls them to the bottom of the container. Particles over 1000 nm in diameter—1000 times as large as atoms, molecules, or ions—form suspensions. The particles in suspension can be separated from heterogeneous mixtures by passing the mixture through a filter.

▶ **MAIN IDEA**

Colloids have particles of intermediate size.

Particles that are intermediate in size between those in solutions and suspensions form mixtures known as colloidal dispersions, or simply colloids. Particles between 1 nm and 1000 nm in diameter may form colloids. After large soil particles settle out of muddy water, the water is often still cloudy because colloidal particles remain dispersed in the water. If the cloudy mixture is poured through a filter, the colloidal particles will pass through the filter, and the mixture will remain cloudy. The particles in a colloid are small enough to be suspended throughout the solvent by the constant movement of the surrounding molecules. The colloidal particles make up the *dispersed phase*, and water is the *dispersing medium*. Many common things you use regularly, such as milk, hair spray, and paint, are colloids. Similar to solutions, colloids can be classified according to their dispersed phase and dispersed medium. For example, a solid might be dispersed in a liquid, as is the case with many paints, or a gas might be dispersed in a liquid, as is the case with foams such as whipped cream. The different types of colloids have common names you may recognize. For example, an *emulsion* is a liquid in a liquid, like milk. And clouds and fog, liquids dispersed in gas, are *liquid aerosols.* **Figure 1.5** on the next page lists these different types of colloids and gives some examples of each one.

FIGURE 1.5

CLASSES OF COLLOIDS

Class of colloid	Phases	Example
sol	solid dispersed in liquid	paints, mud
gel	solid network extending throughout liquid	gelatin
liquid emulsion	liquid dispersed in a liquid	milk, mayonnaise
foam	gas dispersed in liquid	shaving cream, whipped cream
solid aerosol	solid dispersed in gas	smoke, airborne particulate matter, auto exhaust
liquid aerosol	liquid dispersed in gas	fog, mist, clouds, aerosol spray
solid emulsion	liquid dispersed in solid	cheese, butter

FIGURE 1.6

Tyndall Effect A beam of light distinguishes a colloid from a solution. The particles in a colloid will scatter light, making the beam visible. The mixture of gelatin and water in the jar on the right is a colloid. The mixture of water and sodium chloride in the jar on the left is a true solution.

Tyndall Effect

Many colloids appear homogeneous because the individual particles cannot be seen. The particles are, however, large enough to scatter light. You have probably noticed that a headlight beam is visible from the side on a foggy night. Known as the *Tyndall effect*, this occurs when light is scattered by colloidal particles dispersed in a transparent medium. The Tyndall effect is a property that can be used to distinguish between a solution and a colloid, as demonstrated in **Figure 1.6.**

The distinctive properties of solutions, colloids, and suspensions are summarized in **Figure 1.7.** The individual particles of a colloid can be detected under a microscope if a bright light is cast on the specimen at a right angle. The particles, which appear as tiny specks of light, are seen to move rapidly in a random motion. This motion is due to collisions of rapidly moving molecules and is called *Brownian motion*, after its discoverer, Robert Brown. Brownian motion is not simply a casual curiosity for interesting lighting effects. In fact, it is one of the strongest macroscopic observations that science has for assuming matter is ultimately composed of particulate atoms and molecules. Only small, randomly moving particles could produce such effects.

FIGURE 1.7

PROPERTIES OF SOLUTIONS, COLLOIDS, AND SUSPENSIONS

Solutions	Colloids	Suspensions
Homogeneous	Heterogeneous	Heterogeneous
Particle size: 0.01–1 nm; can be atoms, ions, molecules	Particle size: 1–1000 nm, dispersed; can be aggregates or large molecules	Particle size: over 1000 nm, suspended; can be large particles or aggregates
Do not separate on standing	Do not separate on standing	Particles settle out
Cannot be separated by filtration	Cannot be separated by filtration	Can be separated by filtration
Do not scatter light	Scatter light (Tyndall effect)	May scatter light but are not transparent

OBSERVING SOLUTIONS, SUSPENSIONS, AND COLLOIDS

PROCEDURE

1. Prepare seven mixtures, each containing 250 mL of water and one of the following substances.

    a. 12 g of sucrose

    b. 3 g of soluble starch

    c. 5 g of clay

    d. 2 mL of food coloring

    e. 2 g of sodium borate

    f. 50 mL of cooking oil

    g. 3 g of gelatin

    Making the gelatin mixture: Soften the gelatin in 65 mL of cold water, and then add 185 mL of boiling water.

2. Observe the seven mixtures and their characteristics. Record the appearance of each mixture after stirring.

3. Transfer to individual test tubes 10 mL of each mixture that does not separate after stirring. Shine a flashlight on each mixture in a dark room. Make note of the mixtures in which the path of the light beam is visible.

DISCUSSION

1. Using your observations, classify each mixture as a solution, suspension, or colloid.

2. What characteristics did you use to classify each mixture?

MATERIALS

- balance
- 7 beakers, 400 mL
- clay
- cooking oil
- flashlight
- gelatin, plain
- hot plate (to boil H_2O)
- red food coloring
- sodium borate ($Na_2B_4O_7 \cdot 10H_2O$)
- soluble starch
- stirring rod
- sucrose
- test-tube rack
- water

SAFETY

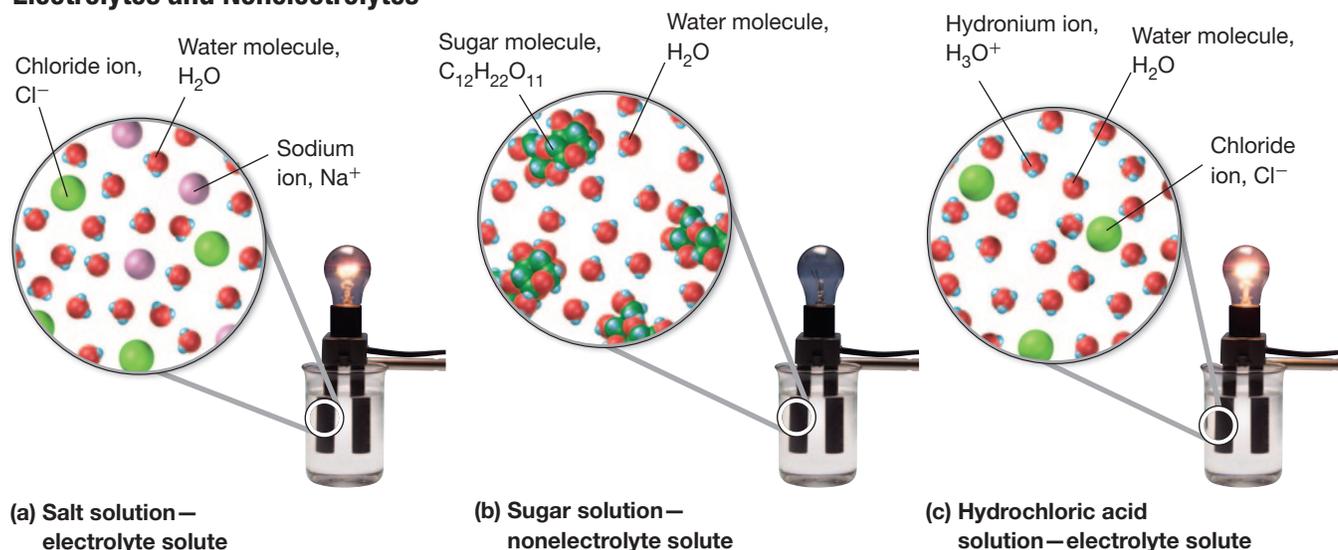Wear safety goggles and an apron.

▶ **MAIN IDEA**

Electrolytes are ionic solutions that conduct electricity.

Substances that dissolve in water are classified according to whether they yield molecules or ions in solution. When an ionic compound dissolves, the positive and negative ions separate from each other and are surrounded by water molecules. These solute ions are free to move, making it possible for an electric current to pass through the solution. **A substance that dissolves in water to give a solution that conducts electric current is called an electrolyte.** Sodium chloride, NaCl, is an electrolyte, as is any soluble ionic compound. Certain highly polar molecular compounds, such as hydrogen chloride, HCl, are also electrolytes because HCl molecules form the ions H_3O^+ and Cl^- when dissolved in water.

By contrast, a solution containing neutral solute molecules does not conduct electric current because it does not contain mobile, charged particles. **A substance that dissolves in water to give a solution that does not conduct an electric current is called a nonelectrolyte.** Sugar is a nonelectrolyte.

FIGURE 1.8

Electrolytes and Nonelectrolytes

Chloride ion, Cl^-

Water molecule, H_2O

Sodium ion, Na^+

Sugar molecule, $C_{12}H_{22}O_{11}$

Water molecule, H_2O

Hydronium ion, H_3O^+

Water molecule, H_2O

Chloride ion, Cl^-

(a) Salt solution— electrolyte solute

(b) Sugar solution— nonelectrolyte solute

(c) Hydrochloric acid solution—electrolyte solute

✔ **CRITICAL THINKING**

Analyze Observe the three solutions. Based on your analysis, explain what is happening in each part (a, b, and c).

Figure 1.8 shows an apparatus for testing the conductivity of solutions. Electrodes are attached to a power supply and make contact with the test solution. If the test solution provides a conducting path, the light bulb will glow. A nonconducting solution is like an open switch between the electrodes, and there is no current in the circuit. The light bulb glows brightly if a solution that is a good conductor is tested. For a moderately conductive solution, the light bulb is dim. If a solution is a poor conductor, the light bulb does not glow at all; such solutions contain solutes that are nonelectrolytes.

✔ SECTION 1 FORMATIVE ASSESSMENT

▶ Reviewing Main Ideas

1. Classify the following as either a heterogeneous or homogeneous mixture. Explain your answers.
   a. orange juice   b. tap water

2. a. What are substances called whose water solutions conduct electricity?
   b. Why does a salt solution conduct electricity?
   c. Why does a sugar-water solution not conduct electricity?

3. Make a drawing of the particles in an NaCl solution to show why this solution conducts electricity. Make a drawing of the particles in an NaCl crystal to show why pure salt does not conduct.

4. Describe one way to prove that a mixture of sugar and water is a solution and that a mixture of sand and water is not a solution.

5. Name the solute and solvent in the following:
   a. 14-karat gold
   b. corn syrup
   c. carbonated, or sparkling, water

✔ Critical Thinking

6. **ANALYZING INFORMATION** If you allow a container of sea water to sit in the sun, the liquid level gets lower and lower, and finally crystals appear. What is happening?

The Solution Process

SECTION 2

Main Ideas

▷ Several factors affect dissolving.

▷ Solubility is a measure of how well one substance dissolves in another.

▷ A change in energy accompanies solution formation.

Key Terms

solution equilibrium
saturated solution
unsaturated solution
supersaturated solution

solubility
hydration
immiscible
miscible

Henry's Law
effervescence
solvated
enthalpy of solution

▷ **MAIN IDEA**
Several factors affect dissolving.

If you have ever tried to dissolve sugar in iced tea, you know that temperature has something to do with how quickly a solute dissolves. What other factors affect how quickly you can dissolve sugar in iced tea?

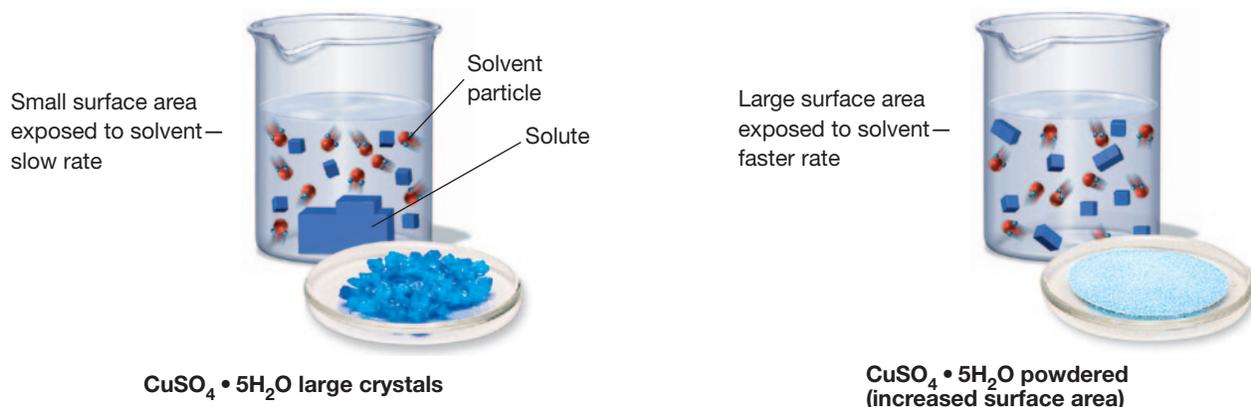### Increasing the Surface Area of the Solute

Sugar dissolves as sugar molecules leave the crystal surface and mix with water molecules. The same is true for any solid solute in a liquid solvent: molecules or ions of the solute are attracted by the solvent.

Because the dissolution process occurs at the surface of the solute, it can be sped up if the surface area of the solute is increased. Crushing sugar that is in cubes or large crystals increases its surface area. In general, the more finely divided a substance is, the greater the surface area per unit mass and the more quickly it dissolves. **Figure 2.1** shows a model of solutions that are made from the same solute but have different amounts of surface area exposed to the solvent.

FIGURE 2.1

Rates of Dissolution The rate at which a solid solute dissolves can be increased by increasing the surface area. A powdered solute has a greater surface area exposed to solvent particles and therefore dissolves faster than a solute in large crystals.

Small surface area exposed to solvent— slow rate

Solvent particle

Solute

$CuSO_4 \cdot 5H_2O$ large crystals

Large surface area exposed to solvent— faster rate

$CuSO_4 \cdot 5H_2O$ powdered (increased surface area)

Agitating a Solution

Very close to the surface of a solute, the concentration of dissolved solute is high. Stirring or shaking helps to disperse the solute particles and bring fresh solvent into contact with the solute surface. Thus, the effect of stirring is similar to that of crushing a solid—contact between the solvent and the solute surface is increased.

Heating a Solvent

You probably have noticed that sugar and other materials dissolve more quickly in warm water than in cold water. As the temperature of the solvent increases, solvent molecules move faster, and their average kinetic energy increases. Therefore, at higher temperatures, collisions between the solvent molecules and the solute are more frequent and of higher energy than at lower temperatures. This separates and disperses the solute molecules.

▶ **MAIN IDEA**

Solubility is a measure of how well one substance dissolves in another.

If you add spoonful after spoonful of sugar to tea, eventually no more sugar will dissolve. For every combination of solvent with a solid solute at a given temperature, there is a limit to the amount of solute that can be dissolved. The point at which this limit is reached for any solute-solvent combination is difficult to predict precisely and depends on the nature of the solute, the nature of the solvent, and the temperature.

The following model describes why there is a limit. When solid sugar is added to water, sugar molecules leave the solid surface and move about at random. Some of these dissolved molecules may collide with the crystal and remain there (recrystallize). As more solid dissolves, these collisions become more frequent. Eventually, molecules are returning to the crystal at the same rate at which they are going into solution, and a dynamic equilibrium is established between dissolution and crystallization. Ionic solids behave similarly, as shown in **Figure 2.2.**

Solution equilibrium is the physical state in which the opposing processes of dissolution and crystallization of a solute occur at equal rates.

FIGURE 2. 2

Solution Equilibrium A saturated solution in a closed system is at equilibrium. The solute is recrystallizing at the same rate that it is dissolving, even though it appears that there is no activity in the system.
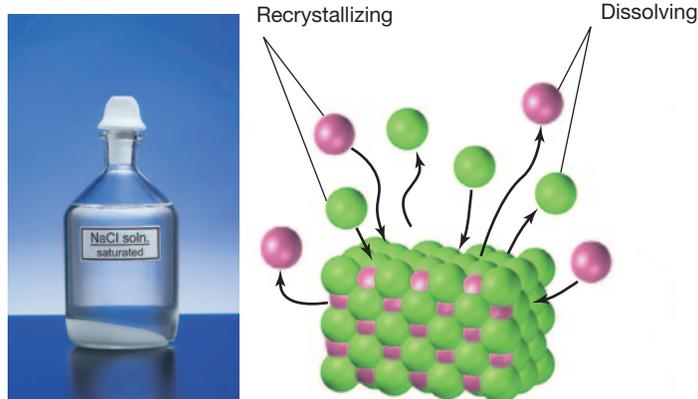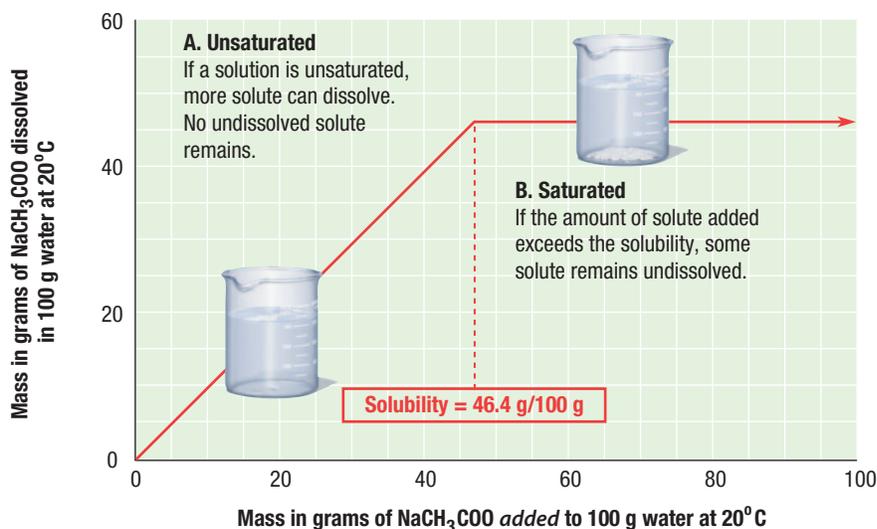

Recrystallizing   Dissolving

NaCl soln. saturated

FIGURE 2.3

Saturation Point The graph shows the range of solute masses that will produce an unsaturated solution. Once the saturation point is exceeded, the system will contain undissolved solute.

Mass of Solute Added vs. Mass of Solute Dissolved

A. Unsaturated
If a solution is unsaturated, more solute can dissolve. No undissolved solute remains.

B. Saturated
If the amount of solute added exceeds the solubility, some solute remains undissolved.

Solubility = 46.4 g/100 g

Mass in grams of $NaCH_3COO$ dissolved in 100 g water at 20°C

Mass in grams of $NaCH_3COO$ *added* to 100 g water at 20°C

Saturated Versus Unsaturated Solutions

A solution that contains the maximum amount of dissolved solute is described as a **saturated solution.** How can you tell that the $NaCH_3COO$ solution pictured in **Figure 2.3** is saturated? If more sodium acetate is added to the solution, it falls to the bottom and does not dissolve because an equilibrium has been established between ions leaving and entering the solid phase. If more water is added to the saturated solution, then more sodium acetate will dissolve in it. At 20°C, 46.4 g of $NaCH_3COO$ is the maximum amount that will dissolve in 100 g of water. A solution that contains less solute than a saturated solution under the existing conditions is an **unsaturated solution.**

Supersaturated Solutions

When a saturated solution of a solute whose solubility increases with temperature is cooled, the excess solute usually comes out of solution, leaving the solution saturated at the lower temperature. But sometimes, if the solution is left to cool undisturbed, the excess solute does not separate, and a supersaturated solution is produced. A **supersaturated solution** is a solution that contains more dissolved solute than a saturated solution contains under the same conditions. A supersaturated solution may remain unchanged for a long time if it is not disturbed, but once crystals begin to form, the process continues until equilibrium is reestablished at the lower temperature. An example of a supersaturated solution is one prepared from a saturated solution of sodium thiosulfate, $Na_2S_2O_3$, or sodium acetate, $NaCH_3COO$. Solute is added to hot water until the solution is saturated, and the hot solution is filtered. The filtrate is left to stand undisturbed as it cools. Dropping a small crystal of the solute into the supersaturated solution ("seeding") or disturbing the solution causes a rapid formation of crystals by the excess solute.

Solubility Values

The solubility of a substance is the amount of that substance required to form a saturated solution with a specific amount of solvent at a specified temperature. The solubility of sugar, for example, is 204 g per 100. g of water at 20.°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubilities must be determined experimentally and can vary widely, as shown in **Figure 2.4.** Solubility values are usually given as grams of solute per 100. g of solvent or per 100. mL of solvent at a given temperature.

"Like Dissolves Like"

Lithium chloride is soluble in water, but gasoline is not. Gasoline mixes with benzene, C_6H_6, but lithium chloride does not. Why?

"Like dissolves like" is a rough but useful rule for predicting solubility. The "like" refers to the polarity (or ionic nature) of the solute and solvent. Polar and ionic solutes tend to dissolve in polar solvents and nonpolar solutes tend to dissolve in nonpolar solvents.

FIGURE 2.4

SOLUBILITY OF SOLUTES AS A FUNCTION OF TEMPERATURE (IN g SOLUTE/100 g H₂O)

Substance	Temperature (°C)					
	0	20	40	60	80	100
$AgNO_3$	122	216	311	440	585	733
$Ba(OH)_2$	1.67	3.89	8.22	20.94	101.4	—
$C_{12}H_{22}O_{11}$	179	204	238	287	362	487
$Ca(OH)_2$	0.189	0.173	0.141	0.121	—	0.07
$Ce_2(SO_4)_3$	20.8	10.1	—	3.87	—	—
KCl	28.0	34.2	40.1	45.8	51.3	56.3
KI	128	144	162	176	192	206
KNO_3	13.9	31.6	61.3	106	167	245
$LiCl$	69.2	83.5	89.8	98.4	112	128
Li_2CO_3	1.54	1.33	1.17	1.01	0.85	0.72
$NaCl$	35.7	35.9	36.4	37.1	38.0	39.2
$NaNO_3$	73	87.6	102	122	148	180
CO_2 (gas)	0.335	0.169	0.0973	0.058	—	—
O_2 (gas)	0.00694	0.00537	0.00308	0.00227	0.00138	0.00

FIGURE 2.5

Hydration When LiCl dissolves, the ions are hydrated. The attraction between ions and water molecules is strong enough that each ion in solution is surrounded by water molecules.

Hydrated Li$^+$

Water molecule

LiCl crystal

Hydrated Cl–
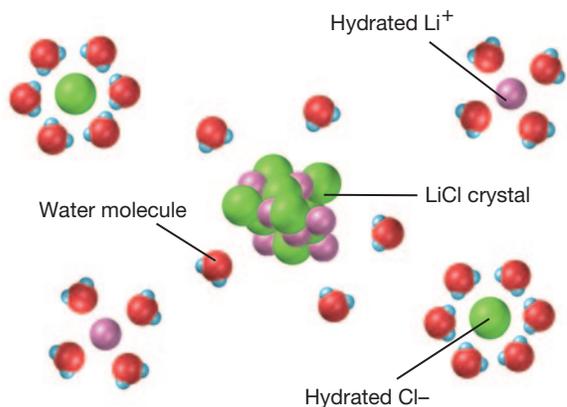
Dissolving Ionic Compounds in Aqueous Solution

The polarity of water molecules plays an important role in the formation of solutions of ionic compounds in water. The slightly charged parts of water molecules attract the ions in the ionic compounds and surround them to keep them separated from the other ions in the solution. Suppose we drop a few crystals of lithium chloride into a beaker of water. At the crystal surface, water molecules come into contact with Li$^+$ and Cl$^-$ ions. The positive region of the water molecules is attracted to Cl$^-$ ions, while the negative region is attracted to Li$^+$ ions. The attraction between water molecules and the ions is strong enough to draw the ions away from the crystal surface and into solution, as illustrated in **Figure 2.5.** This solution process with water as the solvent is referred to as hydration. The ions are said to be *hydrated.* As hydrated ions diffuse into the solution, other ions are exposed and drawn away from the crystal surface by the solvent. The entire crystal gradually dissolves, and hydrated ions become uniformly distributed in the solution.

When crystallized from aqueous solutions, some ionic substances form crystals that incorporate water molecules. These crystalline compounds, known as *hydrates,* retain specific ratios of water molecules, as shown in **Figure 2.6,** and are represented by formulas such as $CuSO_4 \bullet 5H_2O$. Heating the crystals of a hydrate can drive off the water of hydration and leave the anhydrous salt. When a crystalline hydrate dissolves in water, the water of hydration returns to the solvent. The behavior of a solution made from a hydrate is no different from the behavior of one made from the anhydrous form. Dissolving either results in a system containing hydrated ions and water.

Nonpolar Solvents

Ionic compounds are generally not soluble in nonpolar solvents such as carbon tetrachloride, CCl_4, and toluene, $C_6H_5CH_3$. The nonpolar solvent molecules do not attract the ions of the crystal strongly enough to overcome the forces holding the crystal together.

Would you expect lithium chloride to dissolve in toluene? No, LiCl is not soluble in toluene. LiCl is ionic, and $C_6H_5CH_3$ is nonpolar, so there is not much intermolecular attraction between them.

FIGURE 2.6

Hydrate Hydrated copper(II) sulfate has water as part of its crystal structure. Heating releases the water and produces the anhydrous form of the substance, which has the formula $CuSO_4$.
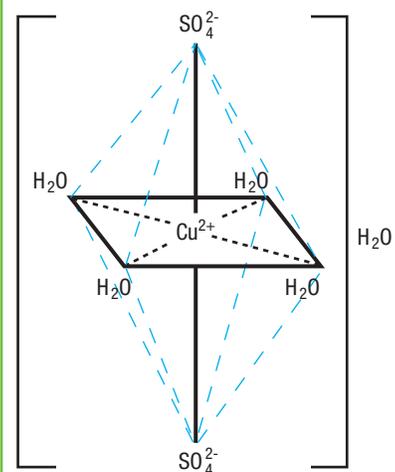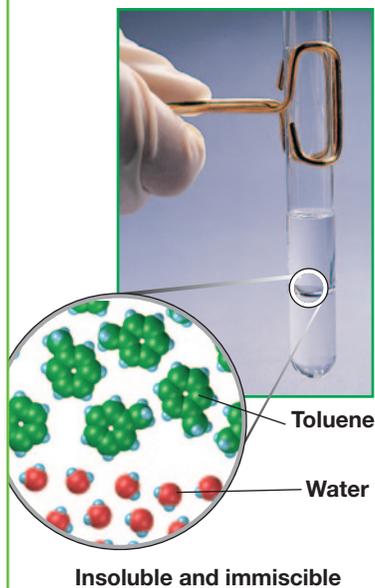
SO_4^{2-}

H_2O H_2O

Cu^{2+} H_2O

H_2O H_2O

SO_4^{2-}

FIGURE 2.7

Immiscibility

Toluene and water are immiscible. The components of this system exist in two distinct phases.

Toluene

Water

Insoluble and immiscible

☑ **CHECK FOR UNDERSTANDING**

Apply Give an example of an everyday mixture that is immiscible.

Liquid Solutes and Solvents

When you shake a bottle of salad dressing, oil droplets become dispersed in the water. As soon as you stop shaking the bottle, the strong attraction of hydrogen bonding between the water molecules squeezes out the oil droplets, forming separate layers. **Liquids that are not soluble in each other are immiscible.** Toluene and water, shown in **Figure 2.7,** are another example of immiscible substances.

Nonpolar substances, such as fats, oils, and greases, are generally quite soluble in nonpolar liquids, such as carbon tetrachloride, toluene, and gasoline. The only attractions between the nonpolar molecules are London forces, which are quite weak. The intermolecular forces existing in the solution are therefore very similar to those in pure substances. Thus, the molecules can mix freely with one another.
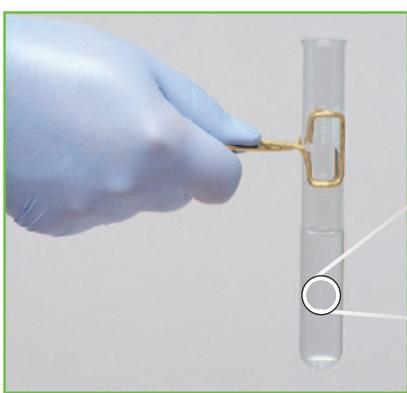
Liquids that dissolve freely in one another in any proportion are said to be miscible. Benzene and carbon tetrachloride are miscible. The nonpolar molecules of these substances exert no strong forces of attraction or repulsion, so the molecules mix freely. Ethanol and water, shown in **Figure 2.8,** also mix freely, but for a different reason. The —OH group on an ethanol molecule is somewhat polar. This group can form hydrogen bonds with water as well as with other ethanol molecules. The intermolecular forces in the mixture are so similar to those in the pure liquids that the liquids are mutually soluble in all proportions.

Gasoline is a solution composed mainly of nonpolar hydrocarbons and is also an excellent solvent for fats, oils, and greases. The major intermolecular forces acting between the nonpolar molecules are weak London forces.

Ethanol is intermediate in polarity between water and carbon tetrachloride. It is not as good as water as a solvent for polar or ionic substances. Sodium chloride is only slightly soluble in ethanol. On the other hand, ethanol is a better solvent than water for less-polar substances because the molecule has a nonpolar region.

FIGURE 2.8

Miscibility

Water molecule, H_2O

Ethanol molecule, C_2H_5OH

Hydrogen bond

$\delta -$

$\delta -$

$\delta +$

$\delta +$

$\delta -$

$\delta -$

$\delta +$

$\delta +$

(a) Water and ethanol are miscible. The components of this system exist in a single phase with a uniform arrangement.
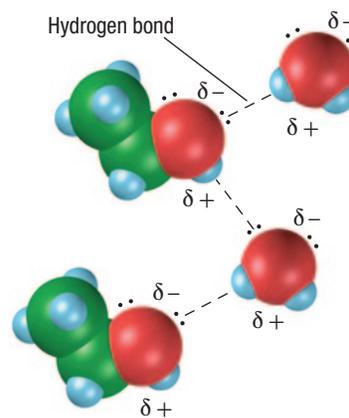
(b) Hydrogen bonding between the solute and solvent enhances the solubility of ethanol in water.

Effects of Pressure on Solubility

Changes in pressure have very little effect on the solubilities of liquids or solids in liquid solvents. However, increases in pressure increase gas solubilities in liquids.

When a gas is in contact with the surface of a liquid, gas molecules can enter the liquid. As the amount of dissolved gas increases, some molecules begin to escape and reenter the gas phase. An equilibrium is eventually established between the rates at which gas molecules enter and leave the liquid phase. As long as this equilibrium is undisturbed, the solubility of the gas in the liquid is unchanged at a given pressure.

$$\text{gas} + \text{solvent} \rightleftarrows \text{solution}$$
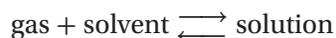
Increasing the pressure of the solute gas above the solution puts stress on the equilibrium. Molecules collide with the liquid surface more often. The increase in pressure is partially offset by an increase in the rate of gas molecules entering the solution. In turn, the increase in the amount of dissolved gas causes an increase in the rate at which molecules escape from the liquid surface and become vapor. Eventually, equilibrium is restored at a higher gas solubility. An increase in gas pressure causes the equilibrium to shift so that more molecules are in the liquid phase.

Henry's Law

Henry's law, named after the English chemist William Henry, states: The solubility of a gas in a liquid is directly proportional to the partial pressure of that gas on the surface of the liquid. Henry's law applies to gas-liquid solutions at constant temperature.

Recall that when a mixture of ideal gases is confined in a constant volume at a constant temperature, each gas exerts the same pressure it would exert if it occupied the space alone. Assuming that the gases do not react in any way, each gas dissolves to the extent it would dissolve if no other gases were present.

In carbonated beverages, the solubility of CO_2 is increased by increasing the pressure. At the bottling plant, carbon dioxide gas is forced into the solution of flavored water at a pressure of 5–10 atm. The gas-in-liquid solution is then sealed in bottles or cans. When the cap is removed, the pressure is reduced to 1 atm, and some of the carbon dioxide escapes as gas bubbles. **The rapid escape of a gas from a liquid in which it is dissolved is known as effervescence.** This is shown in **Figure 2.9.**

FIGURE 2.9

Partial Pressure

(a) There are no gas bubbles in the unopened bottle of soda because the pressure of CO_2 applied during bottling keeps the carbon dioxide gas dissolved in the liquid.

CO_2 under high pressure above solvent

Soluble CO_2 molecules

(b) When the cap on the bottle is removed, the pressure of CO_2 on the liquid is reduced, and CO_2 can escape from the liquid. The soda effervesces when the bottle is opened and the pressure is reduced.

Air at atmospheric pressure

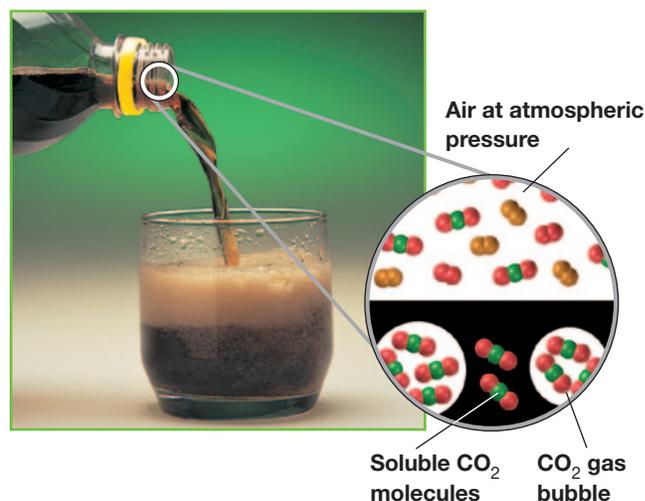Soluble CO_2 molecules

CO_2 gas bubble

FIGURE 2.10

Temperature and Solubility of Gases The solubility of gases in water decreases with increasing temperature. Which gas has the greater solubility at 30°C—CO_2 or SO_2?

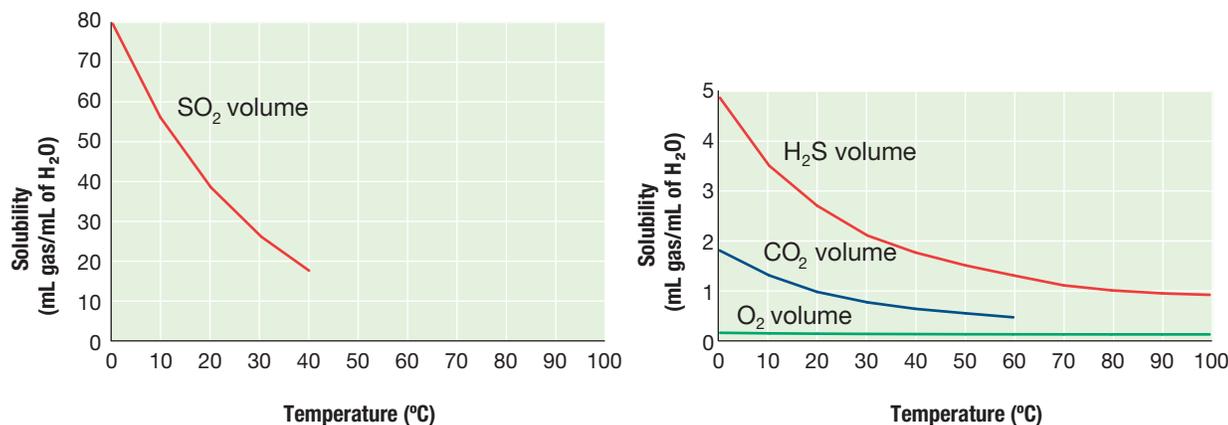**Solubility vs. Temperature Data for Some Gases**

FIGURE 2.11

Temperature and Solubility of Solids Solubility curves for various solid solutes generally show increasing solubility with increases in temperature. From the graph, you can see that the solubility of $NaNO_3$ is affected more by temperature than is NaCl.

Solubility vs. Temperature for Some Solid Solutes
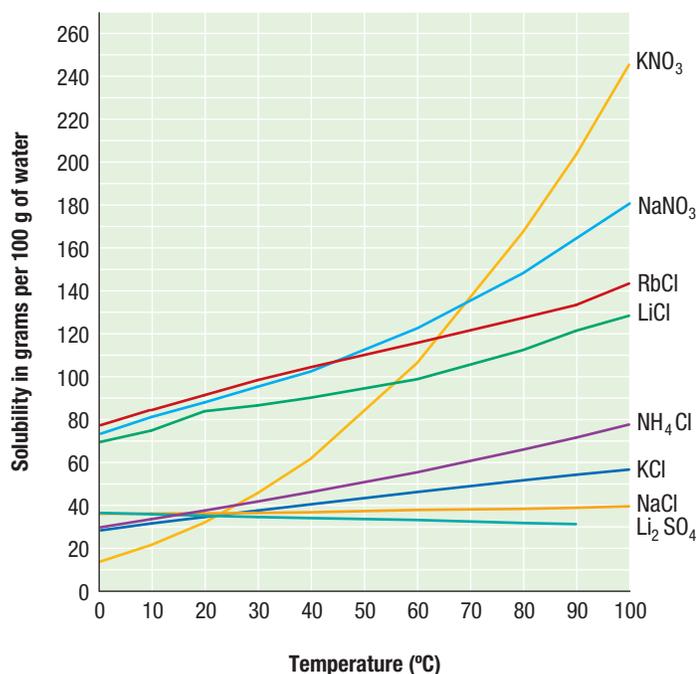
Effects of Temperature on Solubility

First, let's consider gas solubility. Increasing the temperature always decreases gas solubility. As the temperature increases, the average kinetic energy of the molecules in solution increases. A greater number of solute molecules are able to escape from the attraction of solvent molecules and return to the gas phase. At higher temperatures, therefore, equilibrium is reached with fewer gas molecules in solution, and gases are less soluble, as shown in **Figure 2.10**.

The effect of temperature on the solubility of solids in liquids is more difficult to predict. Often, increasing the temperature increases the solubility of solids. However, an equivalent temperature increase can result in a large increase in solubility for some solvents and only a slight change for others.

Compare the effect of temperature on the solubility of NaCl with that on the solubility of potassium nitrate, KNO_3 (**Figure 2.11**). About 14 g of potassium nitrate dissolves in 100 g of water at 0°C. The solubility of potassium nitrate increases by more than 150 g KNO_3 per 100 g H_2O when the temperature is raised to 80°C.

Under similar circumstances, the solubility of sodium chloride increases by only about 2 g NaCl per 100 g H_2O. Sometimes, solubility of a solid *decreases* with an increase in temperature. For example, between 0°C and 60.°C the solubility of cerium sulfate, $Ce_2(SO_4)_3$, decreases by about 17 g/100 g.

▶ **MAIN IDEA**

A change in energy accompanies solution formation.

The formation of a solution is accompanied by an energy change. If you dissolve some potassium iodide, KI, in water, you will find that the outside of the container feels cold to the touch. But if you dissolve some sodium hydroxide, NaOH, in the same way, the outside of the container feels hot. The formation of a solid-liquid solution can apparently either absorb energy (KI in water) or release energy as heat (NaOH in water).

During solution formation, changes occur in the forces between solvent and solute particles. Before dissolving begins, solvent and solute molecules are held to one another by intermolecular forces (solvent-solvent or solute-solute). Energy is required to separate each from their neighbors. **A solute particle that is surrounded by solvent molecules is said to be solvated. The net amount of energy absorbed as heat by the solution when a specific amount of solute dissolves in a solvent is the enthalpy of solution. Figure 2.12** should help you understand this process better. From the model, you can see that the enthalpy of solution is negative (energy is released) when the sum of attractions from Steps 1 and 2 is less than that of Step 3. The enthalpy of solution is positive (energy is absorbed) when the sum of attractions from Steps 1 and 2 is greater than that of Step 3.

FIGURE 2.12

Enthalpy of Solution The graph shows the changes in the enthalpy that occur during the formation of a solution. How would the graph differ for a system with an endothermic heat of solution?

Components

Solute   Solvent

Step 1
Solute particles becoming separated from solid.
Energy absorbed

Step 2
Solvent particles being moved apart to allow solute particles to enter liquid.
Energy absorbed

Step 3
Solvent particles being attracted to and solvating solute particles.
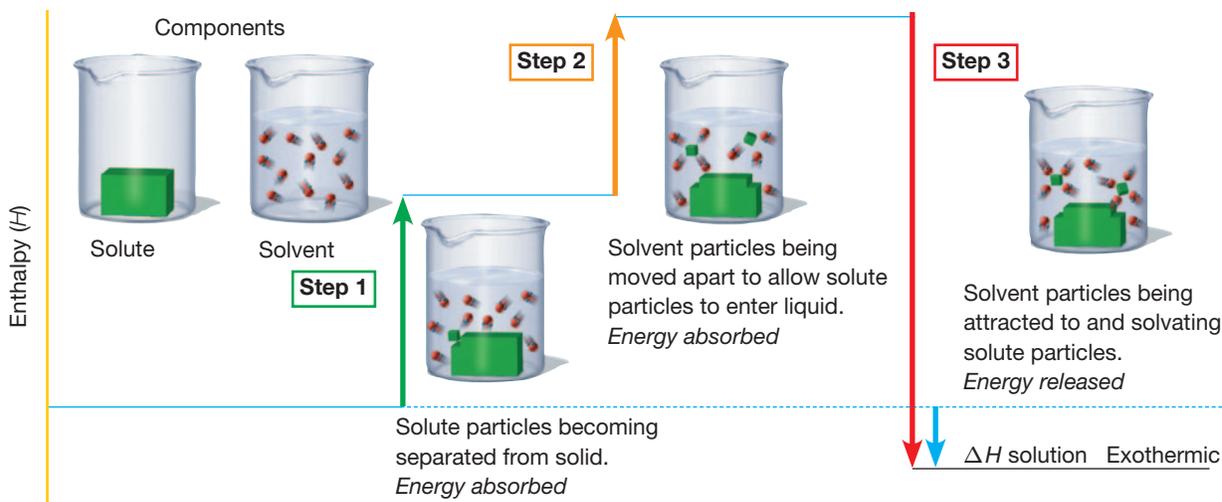*Energy released*

Enthalpy (*H*)

ΔH solution   Exothermic

FIGURE 2.13

ENTHALPIES OF SOLUTION (kJ/MOL SOLUTE AT 25°C)

Substance	Enthalpy of solution	Substance	Enthalpy of solution
$AgNO_3(s)$	+22.59	$KOH(s)$	−57.61
$CH_3COOH(l)$	−1.51	$MgSO_4(s)$	+15.9
$HCl(g)$	−74.84	$NaCl(s)$	+3.88
$HI(g)$	−81.67	$NaNO_3(s)$	+20.50
$KCl(s)$	+17.22	$NaOH(s)$	−44.51
$KClO_3(s)$	+41.38	$NH_3(g)$	−30.50
$KI(s)$	+20.33	$NH_4Cl(s)$	+14.78
$KNO_3(s)$	+34.89	$NH_4NO_3(s)$	+25.69

You know that heating decreases the solubility of a gas, so dissolution of gases is exothermic. How do the values for the enthalpies of solution in **Figure 2.13** support this idea of exothermic solution processes for gaseous solutes?

In the gaseous state, molecules are so far apart that there are virtually no intermolecular forces of attraction between them. Therefore, the solute-solute interaction has little effect on the enthalpy of a solution of a gas. Energy is released when a gas dissolves in a liquid because attraction between solute gas and solvent molecules outweighs the energy needed to separate solvent molecules.

SECTION 2 FORMATIVE ASSESSMENT

▶ Reviewing Main Ideas

1. Why would you expect a packet of sugar to dissolve faster in hot tea than in iced tea?

2. a. Explain how you would prepare a saturated solution of sugar in water.
b. How would you make it a supersaturated solution of sodium acetate in water?

3. Explain why ethanol will dissolve in water and carbon tetrachloride will not.

4. When a solute molecule is solvated, is energy released or absorbed?

5. If a warm bottle of soda and a cold bottle of soda are opened, which will effervesce more and why?

✔ Critical Thinking

6. PREDICTING OUTCOMES You get a small amount of lubricating oil on your clothing. Which would work better to remove the oil—water or toluene? Explain.

7. INTERPRETING CONCEPTS A "fizz saver" pumps air under pressure into a soda bottle to keep gas from escaping. Will this keep CO_2 in the soda bottle? Explain your answer.

Artificial Blood

A patient lies bleeding on a stretcher. The doctor leans over to check the patient's wounds and barks an order to a nearby nurse: "Get him a unit of artificial blood, stat!" Although this may sound like something from a science fiction movie, the scenario may soon be commonplace thanks to a synthetic mixture that can perform one of the main functions of human blood—transporting oxygen.

The hemoglobin inside red blood cells collects oxygen in our lungs, transports it to all the tissues of the body, and then takes carbon dioxide back to the lungs. A blood substitute could accomplish the same task, but it uses nonpolar chemicals called *perfluorocarbons* instead of hemoglobin to transport the oxygen. The perfluorocarbons are carried in a water-based saline solution, but because nonpolar substances and water do not mix well, a bonding chemical called a *surfactant* is added to hold the mixture together. The perfluorocarbons are sheared into tiny droplets and then coated with the bonding molecules. One end of these molecules attaches to the perfluorocarbon, and the other end attaches to the water, creating a milky emulsion. Blood-substitute mixtures are being researched and, if approved for treatment in the U.S., could be administered to future patients in the same way regular blood is. The perfluorocarbons are eventually exhaled through the lungs in the same way other products of respiration are.

Blood substitutes only function to carry gases to and from tissues; they cannot clot or perform any of the immune-system functions that blood does. Still, substitutes have several advantages over real blood. They could have a much longer shelf life. They could also eliminate many of the risks associated with blood transfusions. Because the substitute can dissolve larger amounts of oxygen than real blood can, smaller amounts of the mixture are needed.
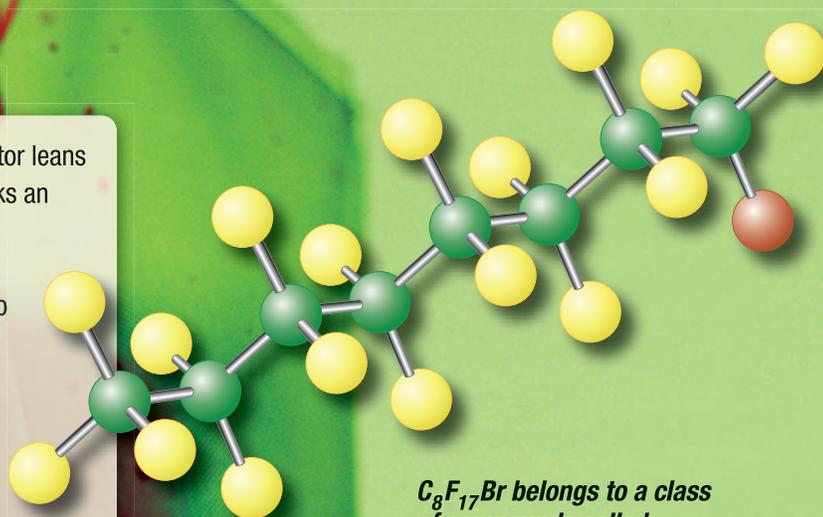
$C_8F_{17}Br$ belongs to a class of compounds called perfluorocarbons.

The collapse of the circulatory system from inadequate oxygen supply, commonly referred to as "shock," could be more effectively prevented with blood substitutes. Blood substitutes also would be useful for treating the effects of decompression sickness scuba divers experience when the higher pressures underwater allow more nitrogen to become dissolved in the blood.

Question

Do Internet research to find more information about artificial blood. Then write a letter to a friend explaining what artificial blood is, what it does in the body, and how it might benefit a family member.

Concentration of Solutions

Key Terms

concentration   molarity   molality

The concentration of a solution is a measure of the amount of solute in a given amount of solvent or solution. Some medications are solutions of drugs—a one-teaspoon dose at the correct concentration might cure the patient, while the same dose in the wrong concentration might kill the patient.

In this section, we introduce two different ways of expressing the concentrations of solutions: molarity and molality.

Sometimes, solutions are referred to as "dilute" or "concentrated," but these are not very definite terms. "Dilute" just means that there is a relatively small amount of solute in a solvent. "Concentrated," on the other hand, means that there is a relatively large amount of solute in a solvent. Note that these terms are unrelated to the degree to which a solution is saturated. A saturated solution of a substance that is not very soluble might be very dilute.

▶ **MAIN IDEA**

Molarity is moles of solute per liter of solution.

Molarity is the number of moles of solute in one liter of solution. For example, a one-molar solution of sodium hydroxide, NaOH, contains one mole of NaOH in every liter of solution. The symbol for molarity is M, and the concentration of a one-molar solution of sodium hydroxide is written as 1 M NaOH.

One mole of NaOH has a mass of 40.0 g. If this quantity of NaOH is dissolved in enough water to make exactly 1.00 L of solution, the solution is a 1.00 M solution. If 20.0 g of NaOH, which is 0.500 mol, is dissolved in enough water to make 1.00 L of solution, a 0.500 M NaOH solution is produced. This relationship between molarity, moles, and volume may be expressed in the following ways.

Molarity (*M*)

$$\text{molarity} = \frac{\text{amount of solute (mol)}}{\text{volume of solution (L)}}$$

$$= \frac{0.500 \text{ mol NaOH}}{1.00 \text{ L}} = 0.500 \text{ M NaOH}$$

If we diluted 1 L of the 0.500 M NaOH solution to 2 L, we could find the new concentration simply by using the ratio $M_1V_1 = M_2V_2$. Here, M_1 and V_1 are the molarity and volume of the original solution, and M_2 and V_2 are the molarity and volume of the new solution. We solve for M_2 as follows:

$$\frac{(0.500 \text{ M})(1 \text{ L})}{(2 \text{ L})} = 0.250 \text{ M}$$

Solutions are often diluted just before use, because those of greater molarity take up less lab space. Chemists can have solutions of greater molarity and lower volume in stock and then dilute them to ones of higher volume and lower molarity for use.

Note that a 1 M solution is *not* made by adding 1 mol of solute to 1 L of *solvent*. Instead, 1 mol of solute is first dissolved in less than 1 L of solvent. The resulting solution is carefully diluted with more solvent to bring the *total volume* to 1 L, as shown in **Figure 3.1.**

FIGURE 3.1

Preparing a 0.5000 M Solution The preparation of a 0.5000 M solution of $CuSO_4 \cdot 5H_2O$ starts with calculating the mass of solute needed.

Start by calculating the mass of $CuSO_4 \cdot 5H_2O$ needed. Making a liter of this solution requires 0.5000 mol of solute. Convert the moles to mass by multiplying by the molar mass of $CuSO_4 \cdot 5H_2O$. This mass is calculated to be 124.9 g.

Add some solvent to the solute to dissolve it, and then pour it into a 1.0-L volumetric flask.

Rinse the weighing beaker with more solvent to remove all the solute, and pour the rinse into the flask. Add water until the volume of the solution nears the neck of the flask.

Put the stopper in the flask, and swirl the solution thoroughly.

Carefully fill the flask to the 1.0-L mark with water.

Restopper the flask, and invert it at least 10 times to ensure complete mixing.

The resulting solution has 0.5000 mol of solute dissolved in 1.000 L of solution, which is a 0.5000 M concentration.

Calculating with Molarity

Sample Problem A You have 3.50 L of solution that contains 90.0 g of sodium chloride, NaCl. What is the molarity of that solution?

① ANALYZE

Given:
solute mass = 90.0 g NaCl
solution volume = 3.50 L

Unknown: molarity of NaCl solution

② PLAN

Molarity is the number of moles of solute per liter of solution. The solute is described in the problem by mass, not the amount in moles. You need one conversion (grams to moles of solute) using the inverted molar mass of NaCl to arrive at your answer.

grams of solute $\longrightarrow$ number of moles of solute $\longrightarrow$ molarity

$$\text{g NaCl} \times \frac{\text{mol NaCl}}{\text{g NaCl}} = \text{mol NaCl}$$

$$\frac{\text{amount of solute (mol)}}{V \text{ solution (L)}} = \text{molarity of solution (M)}$$

③ SOLVE

You will need the molar mass of NaCl.
NaCl = 58.44 g/mol

$$90.0 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 1.54 \text{ mol NaCl}$$

$$\frac{1.54 \text{ mol NaCl}}{3.50 \text{ L of solution}} = 0.440 \text{ M NaCl}$$

④ CHECK YOUR WORK

Because each factor involved is limited to three significant digits, the answer should have three significant digits—which it does. The units cancel correctly to give the desired moles of solute per liter of solution, which is molarity.

Calculating with Molarity

Sample Problem B You have 0.8 L of a 0.5 M HCl solution. How many moles of HCl does this solution contain?

① ANALYZE

Given:
volume of solution = 0.8 L
concentration of solution = 0.5 M HCl

Unknown: moles of HCl in a given volume

② PLAN

The molarity indicates the moles of solute that are in one liter of solution. Given the volume of the solution, the number of moles of solute can then be found.

concentration (mol of HCl/L of solution) × volume (L of solution) = mol of HCl

Calculating with Molarity (continued)

3 SOLVE

$$\frac{0.5 \text{ mol HCl}}{1.0 \cancel{\text{ L of solution}}} \times 0.8 \cancel{\text{ L of solution}} = 0.4 \text{ mol HCl}$$

4 CHECK YOUR WORK

The answer is correctly given to one significant digit. The units cancel correctly to give the desired unit, mol. There should be less than 0.5 mol HCl, because less than 1 L of solution was used.

Calculating with Molarity

Sample Problem C A 23.4 g sample of potassium chromate is needed to carry out a reaction in aqueous solution of silver chromate. All you have on hand is 5 L of a 6.0 M K_2CrO_4 solution. What volume of the solution is needed to give you the 23.4 g K_2CrO_4 needed for the reaction?

1 ANALYZE

Given: volume of solution = 5 L
concentration of solution = 6.0 M K_2CrO_4
mass of solute = 23.4 g K_2CrO_4

Unknown: volume of K_2CrO_4 solution in L

2 PLAN

The molarity indicates the moles of solute that are in 1 L of solution. Given the mass of solute needed, the amount in moles of solute can then be found. Use the molarity and the amount, in moles, of K_2CrO_4 to determine the volume of K_2CrO_4 that will provide 23.4 g.

grams of solute $\longrightarrow$ moles solute
moles solute and molarity $\longrightarrow$ liters of solution needed

3 SOLVE

To get the moles of solute, you'll need to calculate the molar mass of K_2CrO_4.

$$1 \text{ mol } K_2CrO_4 = 194.2 \text{ g } K_2CrO_4$$

$$23.4 \cancel{\text{ g } K_2CrO_4} \times \frac{1 \text{ mol } K_2CrO_4}{194.2 \cancel{\text{ g } K_2CrO_4}} = 0.120 \text{ mol } K_2CrO_4$$

Use the molarity of the solution to convert. Since you know 6.0 mol of K_2CrO_4 is in 1 L of solution, you will need:

$$0.120 \text{ mol } K_2CrO_4 \times \frac{1 \text{ L}}{6.0 \text{ mol}} =$$

$$= 0.020 \text{ L of } K_2CrO_4 \text{ solution}$$

4 CHECK YOUR WORK

The answer is correctly given to two significant digits. The units cancel correctly to give the desired unit, liters of solution.

Practice

1. What is the molarity of a solution composed of 5.85 g of potassium iodide, KI, dissolved in enough water to make 0.125 L of solution?
2. How many moles of H_2SO_4 are present in 0.500 L of a 0.150 M H_2SO_4 solution?
3. What volume of 3.00 M NaCl is needed for a reaction that requires 146.3 g of NaCl?

▶ MAIN IDEA

Molality is moles of solute per kilogram of solvent.

Molality is the concentration of a solution expressed in moles of solute per kilogram of solvent. A solution that contains 1 mol of solute, sodium hydroxide, NaOH, for example, dissolved in exactly 1 kg of solvent is a one-molal solution. The symbol for molality is m, and the concentration of this solution is written as 1 m NaOH.

One mole of NaOH has a molar mass of 40.0 g, so 40.0 g of NaOH dissolved in 1 kg of water results in a one-molal NaOH solution. If 20.0 g of NaOH, which is 0.500 mol of NaOH, is dissolved in exactly 1 kg of water, the concentration of the solution is 0.500 m NaOH.

molality (*m*)	$\text{molality} = \dfrac{\text{moles solute}}{\text{mass of solvent (kg)}}$

$$\frac{0.500 \text{ mol NaOH}}{1.00 \text{ kg H}_2\text{O}} = 0.500 \ m \text{ NaOH}$$

If 80.0 g of sodium hydroxide, which is 2.00 mol, is dissolved in 1 kg of water, a 2.00 m solution of NaOH is produced. The molality of any solution can be found by dividing the number of moles of solute by the mass in kilograms of the solvent in which it is dissolved. Note that if the amount of solvent is expressed in grams, the mass of solvent must be converted to kilograms by multiplying by the following conversion factor:

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

Figure 3.2 shows how a 0.5000 m solution of $CuSO_4 \bullet 5H_2O$ is prepared, in contrast with the 0.5000 M solution in **Figure 3.1.**

FIGURE 3.2

Preparation of a 0.5000 *m* Solution The preparation of a 0.5000 m solution of $CuSO_4 \bullet 5H_2O$ also starts with the calculation of the mass of solute needed.

① Calculate the mass of $CuSO_4 \bullet 5H_2O$ needed. Making this solution will require 0.5000 mol of $CuSO_4 \bullet 5H_2O$ per kilogram of solvent (1000 g). This mass is calculated to be 124.9 g.

② Add 1.000 kg of solvent to the solute in the beaker. Because the solvent is water, 1.000 kg will equal 1000 mL.

③ Mix thoroughly.

④ The resulting solution has 0.5000 mol of solute dissolved in 1.000 kg of solvent.

Concentrations are expressed as molalities when studying properties of solutions related to vapor pressure and freezing point and boiling point changes. Molality is used because it does not change with changes in temperature. Below is a comparison of the equations for molarity and molality.

| Molarity (*M*) | $\text{molarity} = \dfrac{\text{amount of A (mol)}}{\text{volume of solution (L)}}$ |
| Molality (*m*) | $\text{molality} = \dfrac{\text{amount of A (mol)}}{\text{mass of solvent (kg)}}$ |

Calculating with Molality

Sample Problem D A solution was prepared by dissolving 81.3 g of ethylene glycol ($C_2H_6O_2$). Find the molal concentration of this solution.

1 ANALYZE

Given: solute mass = 81.3 g $C_2H_6O_2$

Unknown: molal concentration of $C_2H_6O_2$

2 PLAN

To find molality, you need moles of solute and kilograms of solvent. The given grams of ethylene glycol must be converted to moles. The mass in grams of solvent must be converted to kilograms.

$$\text{mol } C_2H_6O_2 = \frac{\text{g } C_2H_6O_2}{\text{molar mass } C_2H_6O_2}$$

$$\text{kg } H_2O = \text{g } H_2O \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$\text{molality } C_2H_6O_2 = \frac{\text{mol } C_2H_6O_2}{\text{kg } H_2O}$$

3 SOLVE

Use the periodic table to compute the molar mass of $C_2H_6O_2$.
$C_2H_6O_2$ = 62.08 g/mol

$$81.3 \text{ g } C_2H_6O_2 \times \frac{1 \text{ mol } C_2H_6O_2}{62.08 \text{ g } C_2H_6O_2} = 1.31 \text{ mol } C_2H_6O_2$$

$$\frac{166 \text{ g } H_2O}{1000 \text{ g/kg}} = 0.166 \text{ kg } H_2O$$

$$\frac{1.31 \text{ mol } C_2H_6O_2}{0.166 \text{ kg } H_2O} = 7.89 \text{ } m \text{ } C_2H_6O_2$$

4 CHECK YOUR WORK

The answer is correctly given to three significant digits. The unit mol solute/kg solvent is correct for molality.

Calculating with Molality

Sample Problem E Solutions of urea, CH_4N_2O, in water, H_2O, are used in industries such as agriculture and automobiles. What mass of urea must be dissolved in 2250 g of water to prepare a 1.50 m solution?

1 ANALYZE

Given: molality of solution = 1.50 m CH_4N_2O
mass of solvent = 2250 g H_2O

Unknown: mass of solute

2 PLAN

Your first step should be to convert the grams of solvent to kilograms. The molality gives you the moles of solute, which can be converted to the grams of solute using the molar mass of CH_4N_2O.

3 SOLVE

Use the periodic table to compute the molar mass of CH_4N_2O.
CH_4N_2O = 60.06 g/mol

$$2250 \text{ g } \cancel{CH_4N_2O} \times \frac{1 \text{ kg}}{1000 \text{ g } \cancel{CH_4N_2O}} = 2.250 \text{ kg } CH_4N_2O$$

$$2.250 \cancel{\text{ kg}} \times \frac{1.5 \text{ mol } CH_4N_2O}{1 \cancel{\text{ kg}}} = 3.375 \text{ mol } CH_4N_2O$$

$$3.375 \text{ mol } \cancel{CH_4N_2O} \times \frac{60.06 \text{ g } CH_4N_2O}{\text{mol } \cancel{CH_4N_2O}} = 203 \text{ g } CH_4N_2O$$

4 CHECK YOUR WORK

The answer has three significant digits and the units for mass of urea.

Practice

1. What is the molality of acetone in a solution composed of 255 g of acetone, $(CH_3)_2CO$, dissolved in 200 g of water?
2. What mass of anhydrous calcium chloride $(CaCl_2)$ should be dissolved in 590.0 g of water to produce a 0.82 m solution?

✔ SECTION 3 FORMATIVE ASSESSMENT

▶ Reviewing Main Ideas

1. What quantity represents the ratio of the number of moles of solute for a given volume of solution?

2. We dissolve 5.00 grams of sugar, $C_{12}H_{22}O_{11}$, in water to make 1.000 L of solution. What is the concentration of this solution expressed as a molarity?

✔ Critical Thinking

3. **ANALYZING DATA** You evaporate all of the water from 100 mL of NaCl solution and obtain 11.3 grams of NaCl. What was the molarity of the NaCl solution?

4. **RELATING IDEAS** Suppose you know the molarity of a solution. What additional information would you need to calculate the molality of the solution?

Math Tutor | Calculating Solution Concentration

You can use the relationship below to calculate the concentration in molarity of any solution.

$$\text{molarity of solution (M)} = \frac{\text{moles of solute (mol)}}{\text{volume of solution (L)}}$$

Suppose you dissolve 20.00 g of NaOH in some water and dilute the solution to a volume of 250.0 mL (0.2500 L). You don't know the molarity of this solution until you know how many moles of NaOH were dissolved. The number of moles of a substance can be found by dividing the mass of the substance by the mass of 1 mol (molar mass) of the substance.

The molar mass of NaOH is 40.00, so the number of moles of NaOH dissolved is

$$20.00 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 0.5000 \text{ mol NaOH}$$

Now you know that the solution has 0.5000 mol NaOH dissolved in 0.2500 L of solution, so you can calculate molarity.

$$\text{molarity of NaOH} = \frac{\text{mol NaOH}}{\text{L solution}} = \frac{0.5000 \text{ mol NaOH}}{0.2500 \text{ L solution}}$$
$$= 2.000 \text{ mol/L} = 2.000 \text{ M NaOH}$$

Problem-Solving TIPS

- Remember that balances measure mass and not moles, so you often have to convert between mass and moles of solute when making or using solutions.

Sample Problem

A 0.5000 L volume of a solution contains 36.49 g of magnesium chloride, $MgCl_2$. What is the molarity of the solution?

You know the volume of the solution, but you need to find the number of moles of the solute $MgCl_2$ by the following conversion:

$$\text{mass } MgCl_2 \times \frac{1 \text{ mol } MgCl_2}{\text{molar mass } MgCl_2} = \text{mol } MgCl_2$$

$$36.49 \text{ g } MgCl_2 \times \frac{1 \text{ mol } MgCl_2}{95.20 \text{ g } MgCl_2} = 0.3833 \text{ mol } MgCl_2$$

Now you can calculate mol $MgCl_2$ per liter of solution (molarity).

$$\frac{0.3833 \text{ mol } MgCl_2}{0.5000 \text{ L solution}} = 0.7666 \text{ M } MgCl_2$$

Practice

1. What is the molarity of a 50.0 mL solution that contains 0.0350 mol of sodium sulfate, Na_2SO_4?
2. What is the molarity of a 400.0 mL solution that contains 45.00 g of cadmium nitrate, $Cd(NO_3)_2$?

BIG IDEA Solutions are homogeneous mixtures of a solute in a solvent. Chemical and physical factors affect how much of one substance will dissolve in another.

SECTION 1 **Types of Mixtures**

- Solutions are homogeneous mixtures.
- Mixtures are classified as solutions, suspensions, or colloids, depending on the size of the solute particles in the mixture.
- The dissolved substance is the solute. Solutions that have water as a solvent are aqueous solutions.
- Solutions can consist of solutes and solvents that are solids, liquids, or gases.
- Suspensions settle out upon standing. Colloids do not settle out, and they scatter light that is shined through them.
- Most ionic solutes and some molecular solutes form aqueous solutions that conduct an electric current. These solutes are called electrolytes.
- Nonelectrolytes are solutes that dissolve in water to form solutions that do not conduct electric currents.

KEY TERMS

soluble
solution
solvent
solute
suspension
colloid
electrolyte
nonelectrolyte

SECTION 2 **The Solution Process**

- A solute dissolves at a rate that depends on the surface area of the solute, how vigorously the solution is mixed, and the temperature of the solvent.
- The solubility of a substance indicates how much of that substance will dissolve in a specified amount of solvent under certain conditions.
- The solubility of a substance depends on the temperature.
- The solubility of gases in liquids increases with increases in pressure.
- The solubility of gases in liquids decreases with increases in temperature.
- The overall energy absorbed as heat by the system when a specified amount of solute dissolves during solution formation is called the enthalpy of solution.

KEY TERMS

solution equilibrium
saturated solution
unsaturated solution
supersaturated solution
solubility

hydration
immiscible
miscible
Henry's law
effervescence
solvated
enthalpy of solution

SECTION 3 **Concentration of Solutions**

- Two useful expressions of concentration are molarity and molality.
- The molar concentration of a solution represents the ratio of moles of solute to liters of solution.
- The molal concentration of a solution represents the ratio of moles of solute to kilograms of solvent.

KEY TERMS

concentration
molarity
molality

SECTION 1

Types of Mixtures

▶ **REVIEWING MAIN IDEAS**

1. **a.** What is the Tyndall effect?
   b. Identify one example of this effect.

2. Given an unknown mixture consisting of two or more substances, explain how you could determine whether that mixture is a true solution, a colloid, or a suspension.

3. Explain why a suspension is considered a heterogeneous mixture.

4. Does a solution have to involve a liquid? Explain your answer.

5. What is the difference between an electrolyte and a nonelectrolyte?

SECTION 2

The Solution Process

▶ **REVIEWING MAIN IDEAS**

6. **a.** What is solution equilibrium?
   b. What factors determine the point at which a given solute-solvent combination reaches equilibrium?

7. **a.** What is a saturated solution?
   b. What visible evidence indicates that a solution is saturated?
   c. What is an unsaturated solution?

8. **a.** What is meant by the solubility of a substance?
   b. What condition(s) must be specified when expressing the solubility of a substance?

9. **a.** What rule of thumb is useful for predicting whether one substance will dissolve in another?
   b. Describe what the rule means in terms of various combinations of polar and nonpolar solutes and solvents.

10. **a.** How does pressure affect the solubility of a gas in a liquid?
    b. What law is a statement of this relationship?

   c. If the pressure of a gas above a liquid is increased, what happens to the amount of the gas that will dissolve in the liquid, if all other conditions remain constant?
   d. Two bottles of soda are opened. One is a cold bottle, and the other is at room temperature. Which system will show more effervescence and why?

11. Based on **Figure 2.11,** determine the solubility of each of the following in grams of solute per 100 g H_2O.
    a. $NaNO_3$ at 10°C
    b. KNO_3 at 60°C
    c. NaCl at 50°C

12. Based on **Figure 2.11**, at what temperature would each of the following solubility levels be observed?
    a. 50 g KCl in 100 g H_2O
    b. 100 g $NaNO_3$ in 100 g H_2O
    c. 60 g KNO_3 in 100 g H_2O

13. The enthalpy of solution for $AgNO_3$ is +22.8 kJ/mol.
    a. Write the equation that represents the dissolution of $AgNO_3$ in water.
    b. Is the dissolution process endothermic or exothermic? Is the crystallization process endothermic or exothermic?
    c. As $AgNO_3$ dissolves, what change occurs in the temperature of the solution?
    d. When the system is at equilibrium, how do the rates of dissolution and crystallization compare?
    e. If the solution is then heated, how will the rates of dissolution and crystallization be affected? Why?
    f. How will the increased temperature affect the amount of solute that can be dissolved?
    g. If the solution is allowed to reach equilibrium and is then cooled, how will the system be affected?

14. What opposing forces are at equilibrium in the sodium chloride system shown in **Figure 2.2**?

SECTION 3

Concentration of Solutions

▶ **REVIEWING MAIN IDEAS**

15. On which property of solutions does the concept of concentration rely?

16. In what units is molarity expressed?

17. Under what circumstances might we prefer to express solution concentrations in terms of
   a. molarity?
   b. molality?

18. If you dissolve 2.00 mol KI in 1.00 L of water, will you get a 2.00 M solution? Explain.

PRACTICE PROBLEMS

19. a. Suppose you wanted to dissolve 106 g of Na_2CO_3 in enough H_2O to make 6.00 L of solution.
   (1) What is the molar mass of Na_2CO_3?
   (2) What is the molarity of this solution?
   b. What is the molarity of a solution of 14.0 g NH_4Br in enough H_2O to make 150. mL of solution?

20. a. Suppose you wanted to prepare 1.00 L of a 3.50 M aqueous solution of H_2SO_4.
   (1) What is the solute?
   (2) What is the solvent?
   (3) How many grams of solute are needed to make this solution?
   b. How many grams of solute are needed to make 2.50 L of a 1.75 M solution of $Ba(NO_3)_2$?

21. How many moles of NaOH are contained in 65.0 mL of a 2.20 M solution of NaOH in H_2O? (Hint: See Sample Problem B.)

22. A solution is made by dissolving 26.42 g of $(NH_4)_2SO_4$ in enough H_2O to make 50.00 mL of solution.
   a. What is the molar mass of $(NH_4)_2SO_4$?
   b. What is the molarity of this solution?

23. Suppose you wanted to find out how many milliliters of 1.0 M $AgNO_3$ are needed to provide 169.9 g of $AgNO_3$.
   a. What is the first step in solving the problem?
   b. What is the molar mass of $AgNO_3$?
   c. How many milliliters of solution are needed?

24. a. Balance this equation:
   $$H_3PO_4 + Ca(OH)_2 \longrightarrow Ca_3(PO_4)_2 + H_2O$$
   b. What mass of each product results if 750 mL of 6.00 M H_3PO_4 reacts completely according to the equation?

25. How many milliliters of 0.750 M H_3PO_4 are required to react with 250 mL of 0.150 M $Ba(OH)_2$ if the products are barium phosphate and water?

26. 75.0 mL of an $AgNO_3$ solution reacts with enough Cu to produce 0.250 g of Ag by single displacement. What is the molarity of the initial $AgNO_3$ solution if $Cu(NO_3)_2$ is the other product?

27. Determine the number of grams of solute needed to make each of the following molal solutions:
   a. a 4.50 m solution of H_2SO_4 in 1.00 kg H_2O
   b. a 1.00 m solution of HNO_3 in 2.00 kg H_2O

28. A solution is prepared by dissolving 17.1 g of sucrose, $C_{12}H_{22}O_{11}$, in 275 g of H_2O.
   a. What is the molar mass of sucrose?
   b. What is the molality of that solution?

29. How many kilograms of H_2O must be added to 75.5 g of $Ca(NO_3)_2$ to form a 0.500 m solution?

30. A solution made from ethanol, C_2H_5OH, and water is 1.75 m ethanol. How many grams of C_2H_5OH are contained per 250. g of water?

Mixed Review

▶ **REVIEWING MAIN IDEAS**

31. Na_2SO_4 is dissolved in water to make 450 mL of a 0.250 M solution.
   a. What is the molar mass of Na_2SO_4?
   b. How many moles of Na_2SO_4 are needed?

32. Citric acid is one component of some soft drinks. Suppose that 2.00 L of solution are made from 150. mg of citric acid, $C_6H_8O_7$.
   a. What is the molar mass of citric acid?
   b. What is the molarity of citric acid in the solution?

33. Suppose you wanted to know how many grams of KCl would be left if 350 mL of a 2.0 M KCl solution were evaporated to dryness.
   a. What is the molar mass of KCl?
   b. How would heating the solution affect the mass of KCl remaining?
   c. How many grams of KCl would remain?

34. Sodium metal reacts violently with water to form NaOH and release hydrogen gas. Suppose that 10.0 g of Na react completely with 1.00 L of water and the final solution volume is 1.00 L.
   a. What is the molar mass of NaOH?
   b. Write a balanced equation for the reaction.
   c. What is the molarity of the NaOH solution formed by the reaction?

35. In cars, ethylene glycol, $C_2H_6O_2$, is used as a coolant and antifreeze. A mechanic fills a radiator with 6.5 kg of ethylene glycol and 1.5 kg of water.
   a. What is the molar mass of ethylene glycol?
   b. What is the molality of the water in the solution?

36. Plot a solubility graph for $AgNO_3$ from the following data, with grams of solute (by increments of 50) per 100 g of H_2O on the vertical axis and with temperature in °C on the horizontal axis.

Grams solute per 100 g H_2O	Temperature (°C)
122	0
216	30
311	40
440	60
585	80
733	100

   a. How does the solubility of $AgNO_3$ vary with the temperature of the water?
   b. Estimate the solubility of $AgNO_3$ at 35°C, 55°C, and 75°C.
   c. At what temperature would the solubility of $AgNO_3$ be 275 g per 100 g of H_2O?
   d. If 100 g of $AgNO_3$ were added to 100 g of H_2O at 10°C, would the resulting solution be saturated or unsaturated? What would occur if 325 g of $AgNO_3$ were added to 100 g of H_2O at 35°C?

37. If a saturated solution of KNO_3 in 100 g of H_2O at 60°C is cooled to 20°C, approximately how many grams of the solute will precipitate out of the solution? (Use **Figure 2.4.**)

38. a. Suppose you wanted to dissolve 294.3 g of H_2SO_4 in 1.000 kg of H_2O.
   (1) What is the solute?
   (2) What is the solvent?
   (3) What is the molality of this solution?
   b. What is the molality of a solution of 63.0 g HNO_3 in 0.250 kg H_2O?

ALTERNATIVE ASSESSMENT

39. Predicting Outcomes You have been investigating the nature of suspensions, colloids, and solutions and have collected the following observational data on four unknown samples. From the data, infer whether each sample is a solution, suspension, or colloid.

		DATA TABLE 1 *Samples*		
Sample	Color	Clarity (clear or cloudy)	Settle out	Tyndall effect
1	green	clear	no	no
2	blue	cloudy	yes	no
3	colorless	clear	no	yes
4	white	cloudy	no	yes

Based on your inferences in Data Table 1, you decide to conduct one more test of the particles. You filter the samples and then reexamine the filtrate. You obtain the data found in Data Table 2. Infer the classifications of the filtrate based on the data in Data Table 2.

		DATA TABLE 2 *Filtrate of Samples*		
Sample	Color	Clarity (clear or cloudy)	On filter paper	Tyndall effect
1	green	clear	nothing	no
2	blue	cloudy	gray solid	yes
3	colorless	clear	none	yes
4	colorless	clear	white solid	no

40. Review the information on alloys in the *Elements Handbook* (Appendix A).
  a. Why is aluminum such an important component of alloys?
  b. What metals make up bronze?
  c. What metals make up brass?
  d. What is steel?
  e. What is the composition of the mixture called *cast iron*?

41. **Table 5A** of the *Elements Handbook* (Appendix A) contains carbon monoxide concentration data expressed as parts per million (ppm). The OSHA (Occupational Safety and Health Administration) limit for worker exposure to CO is 200 ppm for an eight-hour period.
  a. At what concentration do harmful effects occur in less than one hour?
  b. By what factor does the concentration in item (a) exceed the maximum limit set by OSHA?

RESEARCH AND WRITING

42. Find out about the chemistry of emulsifying agents. How do these substances affect the dissolution of immiscible substances such as oil and water? As part of your research on this topic, find out why eggs are an emulsifying agent for baking mixtures.

ALTERNATIVE ASSESSMENT

43. Make a comparison of the electrolyte concentration in various brands of sports drinks. Using the labeling information for sugar, calculate the molarity of sugar in each product or brand. Construct a poster to show the results of your analysis of the product labels.

44. Write a set of instructions on how to prepare a solution that is 1 M $CuSO_4$, using $CuSO_4 \bullet 5H_2O$ as the solute. How do the instructions differ if the solute is anhydrous $CuSO_4$? Your instructions should include a list of all materials needed.

Standards-Based Assessment

Record your answers on a separate piece of paper.

MULTIPLE CHOICE

1 Which one of the following statements is false?

A Gases are generally more soluble in water under high pressures than under low pressures.

B As temperature increases, the solubilities of solids in water increase.

C Water dissolves many ionic solutes because of its ability to hydrate ions in solution.

D Many solids dissolve more quickly in a cold solvent than in a warm solvent.

2 Two liquids are likely to be immiscible if

A both have polar molecules.

B both have nonpolar molecules.

C one is polar and the other is nonpolar.

D one is water and the other is methyl alcohol, CH_3OH.

3 The solubility of a gas in a liquid would be increased by

A the addition of an electrolyte.

B the addition of an emulsifier.

C agitation of the solution.

D an increase in its partial pressure.

4 Which of the following types of compounds is most likely to be a strong electrolyte?

A a polar compound

B a nonpolar compound

C a covalent compound

D an ionic compound

5 A saturated solution can become supersaturated under which of the following conditions?

A It contains electrolytes.

B A heated saturated solution is allowed to cool.

C More solvent is added.

D More solute is added.

6 Molarity is expressed in units of

A moles of solute per liter of solution.

B liters of solution per mole of solute.

C moles of solute per liter of solvent.

D liters of solvent per mole of solute.

7 What mass of NaOH is contained in 2.5 L of a 0.010 M solution?

A 0.010 g

B 1.0 g

C 2.5 g

D 0.40 g

GRIDDED RESPONSE

8 If 75.0 mL of an $AgNO_3$ solution reacts with enough Cu to produce 0.250 g Ag by single displacement, what is the molarity of the initial $AgNO_3$ solution if $Cu(NO_3)_2$ is the other product? (Give your answer to three significant figures.)

Test Tip

Before choosing an answer to a question, try to answer the question without looking at the answer choices on the test.