CHAPTER 12

Solutions

Solutions are homogeneous mixtures of two or more substances in a single phase.

Solution Being Poured into a Beaker; Centrifuge Tubes in the Background

Types of Mixtures

SECTION 1

Objectives

- Distinguish between heterogeneous and homogeneous mixtures.
- List three different solute-solvent combinations.
- Compare the properties of suspensions, colloids, and solutions.
- Distinguish between electrolytes and nonelectrolytes.

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 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. A model for a homogeneous mixture such as salt water is shown in **Figure 1b.**

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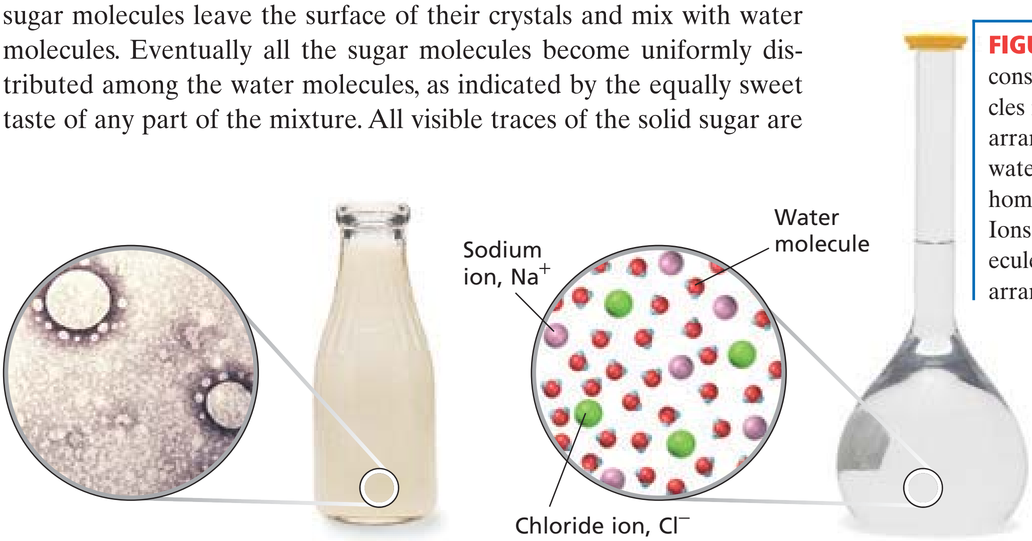
Module 1: States of Matter/Classes of Matter

Solutions

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.*

What happens as sugar dissolves? The 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



(a) Heterogeneous mixture—milk

(b) Homogeneous mixture—salt solution

FIGURE 1 (a) Milk 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.

gone. Such a mixture is called a solution. *A* **solution** *is a homogeneous mixture of two or more substances in a single phase*. In a solution, atoms, molecules, or ions are thoroughly mixed, resulting in a mixture that has the same composition and properties throughout.

Components of Solutions

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 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 in fact unnecessary, to say which is the solvent and which is the solute.

In a solution, the dissolved solute particles are so small that they cannot be seen. They remain mixed with the solvent indefinitely, so long as the existing conditions remain unchanged. If the solutions in **Figure 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 **Table 1.** In each example, one component is designated as the solvent and one as the 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 higher strength and

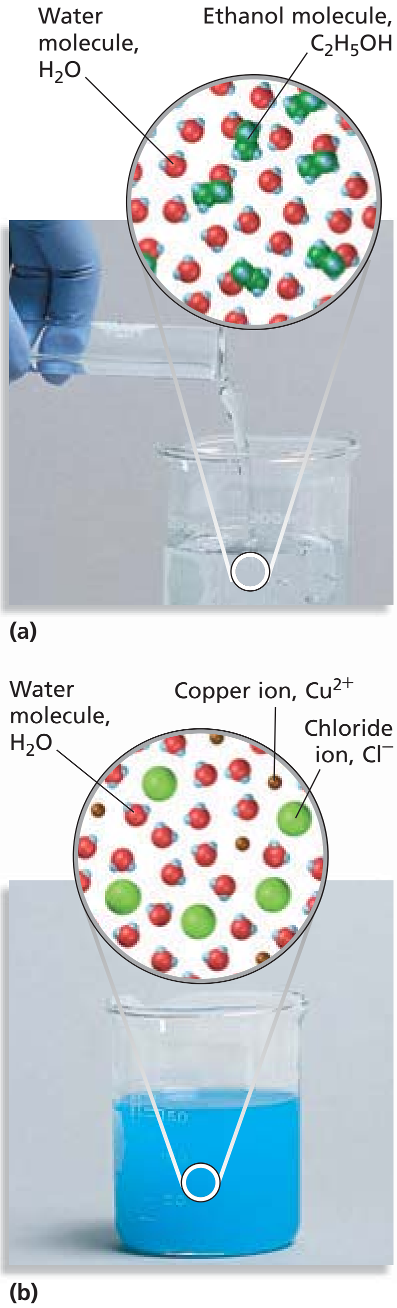


FIGURE 2 The solute in a solution can be a solid, liquid, or gas. (a) The ethanol-water solution is made from a liquid solute in a liquid solvent. (b) The copper(II) chloride–water solution is made from a solid solute in a liquid solvent. Note that the composition of each solution is uniform.

TABLE 1 *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 3 (a) 24-karat gold is pure gold. (b) 14-karat gold is an alloy of gold with silver and copper. 14-karat gold is 14/24, or 58.3%, gold.

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 3** shows a comparison between pure gold and a gold alloy. 14-karat gold is a solution because the gold, silver, and copper are evenly mixed at the atomic level.

Suspensions

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.

Colloids

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, 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*. Examples of the various types of colloids are given in **Table 2.** Note that some familiar terms, such as *emulsion* and *foam,* refer to specific types of colloids. For example, mayonnaise is an emulsion of oil

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Topic: Solutions
SciLinks code: HC61422

Topic: Colloids
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TABLE 2 ***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

droplets in water; the egg yolk in it acts as an emulsifying agent, which helps to keep the oil droplets dispersed.

FIGURE 4 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. This effect, known as the *Tyndall effect,* 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 4.**

The distinctive properties of solutions, colloids, and suspensions are summarized in **Table 3.** 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.

TABLE 3 ***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

QuickLAB

Wear safety goggles and an apron.

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

Solutes: Electrolytes Versus Nonelectrolytes

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

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Module 8: Strong and Weakly Ionized Species, pH, and Titrations

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Topic: Electrolytes/ Nonelectrolytes

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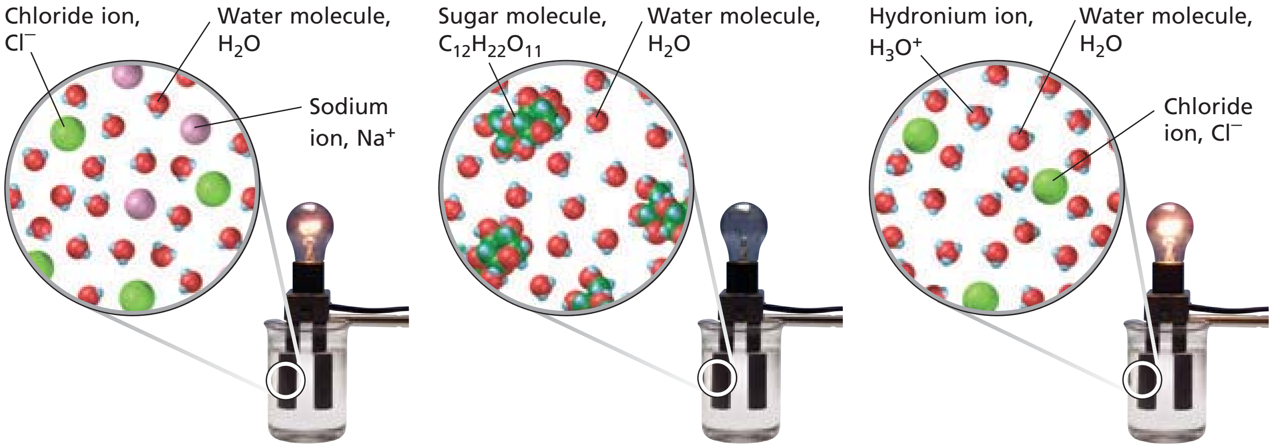


(a) Salt solution—electrolyte solute

(b) Sugar solution—nonelectrolyte solute

(c) Hydrochloric acid solution—electrolyte solute

FIGURE 5 (a) Sodium chloride dissolves in water to produce a salt solution that conducts electric current. NaCl is an electrolyte. (b) Sucrose dissolves in water to produce a sugar solution that does not conduct electricity. Sucrose is a nonelectrolyte. (c) Hydrogen chloride dissolves in water to produce a solution that conducts current. HCl is an electrolyte.

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 5** shows an apparatus for testing the conductivity of solutions. The electrodes are conductors that are attached to a power supply and that make electric contact with the test solution. For a current to pass through the light-bulb filament, the test solution must provide a conducting path between the two electrodes. 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. Such solutions contain solutes that are electrolytes. For a moderately conductive solution, however, 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. You will learn more about the strengths and behavior of electrolytes in Chapter 13.

SECTION REVIEW

1. Classify the following as either a heterogeneous or homogeneous mixture, and 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

OBJECTIVES

- List and explain three factors that affect the rate at which a solid solute dissolves in a liquid solvent.
- Explain solution equilibrium, and distinguish among saturated, unsaturated, and supersaturated solutions.
- Explain the meaning of "like dissolves like" in terms of polar and nonpolar substances.
- List the three interactions that contribute to the enthalpy of solution, and explain how they combine to cause dissolution to be exothermic or endothermic.
- Compare the effects of temperature and pressure on solubility.

Factors Affecting the Rate of Dissolution

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.

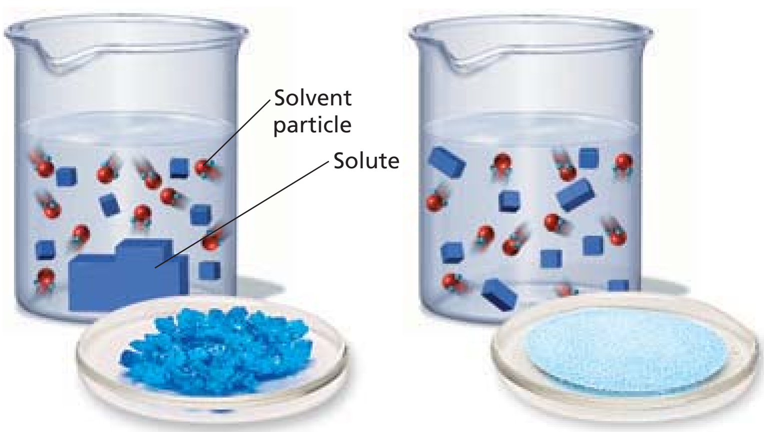
Because the dissolution process occurs at the surface of the solute, it can be speeded up if the surface area of the solute is increased. Crushing sugar that is in cubes or large crystals increases its surface area. In general, the more finely divided a substance is, the greater the surface area per unit mass and the more quickly it dissolves. **Figure 6** shows a model of solutions that are made from the same solute but have a different amount of surface area exposed to the solvent.

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

Small surface area exposed to solvent—slow rate

Large surface area exposed to solvent—faster rate



$CuSO_4$•$5H_2O$ large crystals

$CuSO_4$•$5H_2O$ powdered
Increased surface area

FIGURE 6 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.

CAREERS in Chemistry

Environmental Chemist

What happens to all of our chemical waste, such as household cleaners and shampoos that we rinse down the drain, industrial smoke, and materials that have not been removed in water treatment plants? Environmental chemists investigate the sources and effects of chemicals in all parts of the environment. Then, chemists also devise acceptable ways to dispose of chemicals. This may involve conducting tests to determine whether the air, water, or soil is contaminated; developing programs to help remove contamination; designing new production processes to reduce the amounts of waste produced; handling regulation and compliance issues; and advising on safety and emergency responses. Environmental chemists must understand and use many other disciplines, including biology, geology and ecology.

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 have probably noticed that sugar and many 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 are of higher energy than at lower temperatures. This helps to separate solute molecules from one another and to disperse them among the solvent molecules.

Solubility

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 first added to water, sugar molecules leave the solid surface and move about at random in the solvent. Some of these dissolved molecules may collide with the crystal and remain there (recrystallize). As more solid dissolves and the concentration of dissolved molecules increases, 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 7.**

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

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Topic: Solubility
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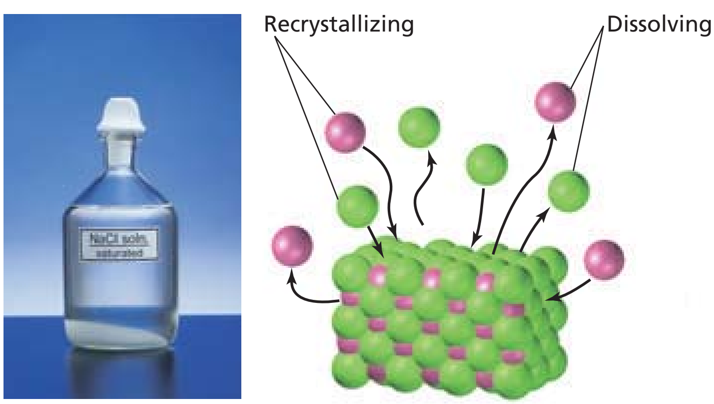

FIGURE 7 A saturated solution in a closed system is at equilibrium. The solute is recrystallizing at the same rate that it is dissolving, even though it appears that there is no activity in the system.

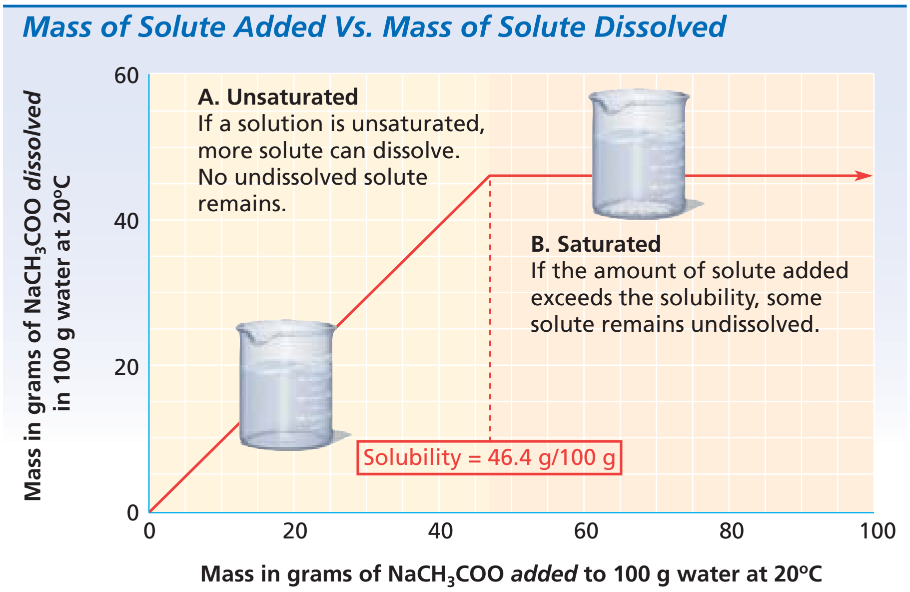

FIGURE 8 The graph shows the range of solute masses that will produce an unsaturated solution. Once the saturation point is exceeded, the system will contain undissolved solute.

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 8** 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.

extension

Chemistry in Action

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Keyword: HC6SLNX

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 they vary widely, as illustrated in **Table 4.** Solubility values can be found in chemical handbooks and are usually given as grams of solute per 100. g of solvent or per 100. mL of solvent at a given temperature.

The rate at which a solid dissolves is unrelated to its solubility at that temperature. The maximum amount of a given solute that dissolves and reaches equilibrium is always the same under the same conditions.

Solute-Solvent Interactions

Lithium chloride is highly soluble in water, but gasoline is not. On the other hand, gasoline mixes readily with benzene, C_6H_6, but lithium chloride does not. Why are there such differences in solubility?

"Like dissolves like" is a rough but useful rule for predicting whether one substance will dissolve in another. What makes substances similar depends on the type of bonding, the polarity or nonpolarity of molecules, and the intermolecular forces between the solute and solvent.

TABLE 4 ***Solubility of Solutes as a Function of Temperature (in g solute/100. g H_2O)***

	Temperature (°C)					
Substance	**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 at SP)	0.335	0.169	0.0973	0.058	—	—
O_2 (gas at SP)	0.00694	0.00537	0.00308	0.00227	0.00138	0.00

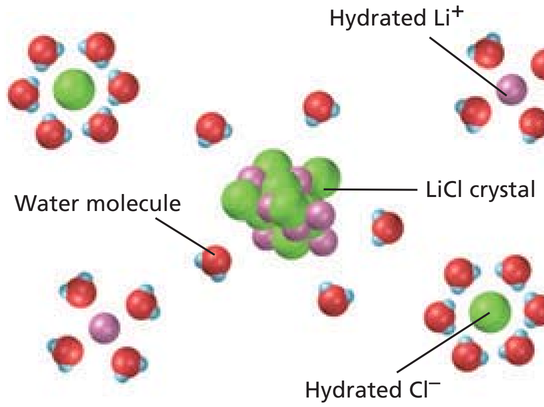


FIGURE 9 When LiCl dissolves, the ions are hydrated. The attraction between ions and water molecules is strong enough that each ion in solution is surrounded by water molecules.

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 surfaces, water molecules come into contact with Li^+ and Cl^- ions. The positive ends of the water molecules are attracted to Cl^- ions, while the negative ends are 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 9.** *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 are 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 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 form 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 and $C_6H_5CH_3$ differ widely in bonding, polarity, and intermolecular forces.

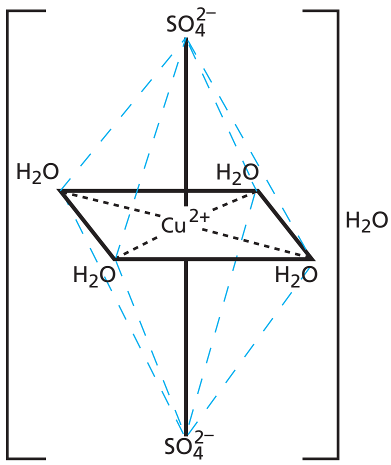


FIGURE 10 Hydrated copper(II) sulfate has water as part of its crystal structure. Heating releases the water and produces the anhydrous form of the substance, which has the formula $CuSO_4$.

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 11,** are another example of immiscible substances.

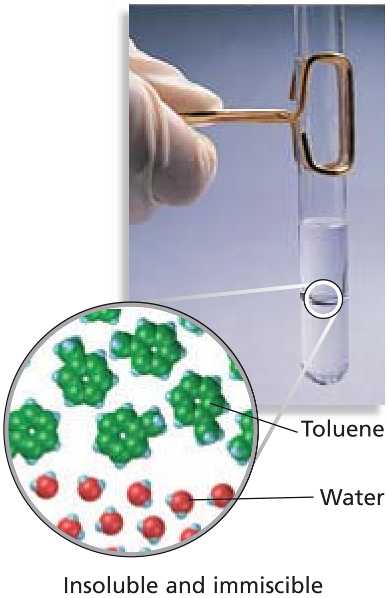
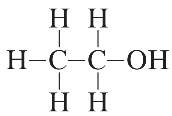


FIGURE 11 Toluene and water are immiscible. The components of this system exist in two distinct phases.

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 12,** 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.

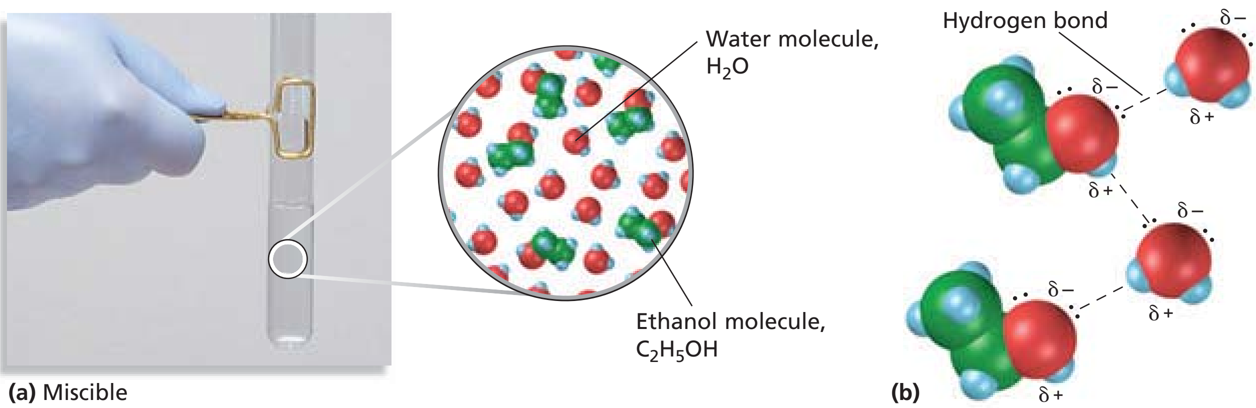


ethanol

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 a solvent for polar or ionic substances as water is. Sodium chloride is only slightly soluble in ethanol. On the other hand, ethanol is a better solvent than water is for less-polar substances because the molecule has a nonpolar region.

FIGURE 12 (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.

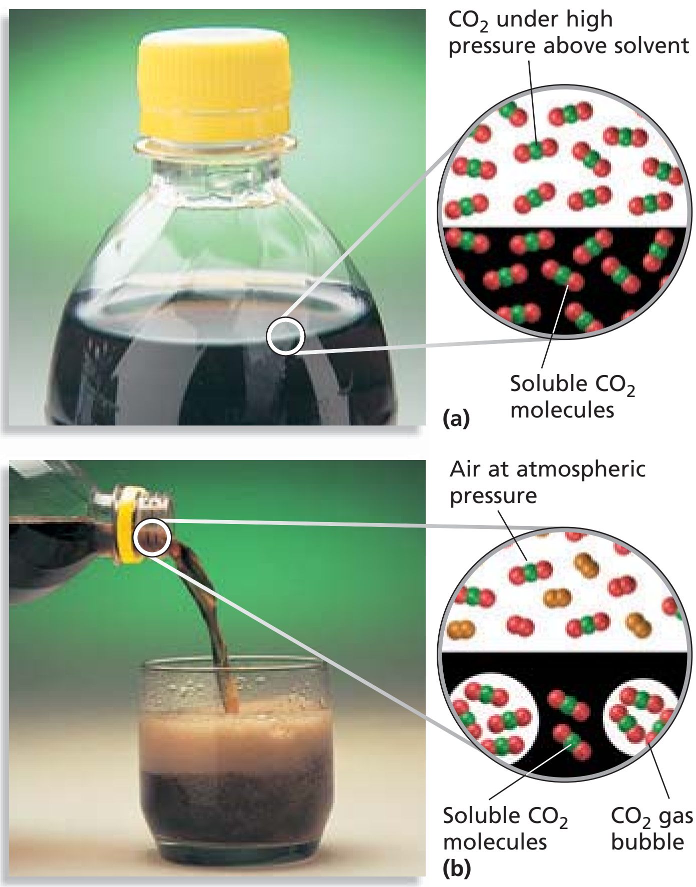
Henry's Law

The solubility of a gas in a liquid is directly proportional to the partial pressure of that gas on the surface of the liquid. This is a statement of **Henry's law,** named after the English chemist William Henry. 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** and is shown in **Figure 13.**

FIGURE 13 (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. (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.



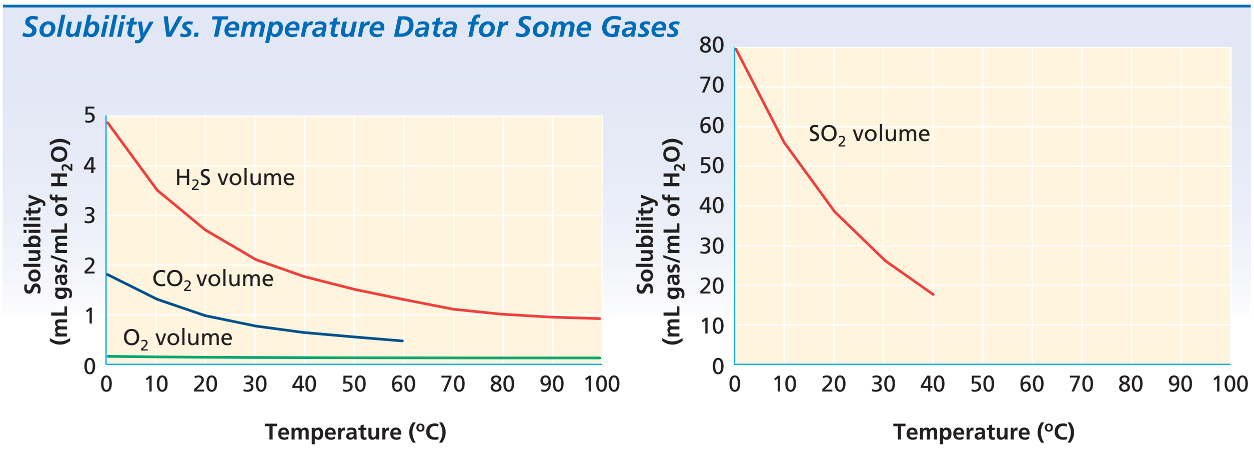


FIGURE 14 The solubility of gases in water decreases with increasing temperature. Which gas has the greater solubility at 30°C—CO_2 or SO_2?

Effects of Temperature on Solubility

First, let's consider gas solubility. Increasing the temperature usually 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 generally less soluble, as shown in **Figure 14.**

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

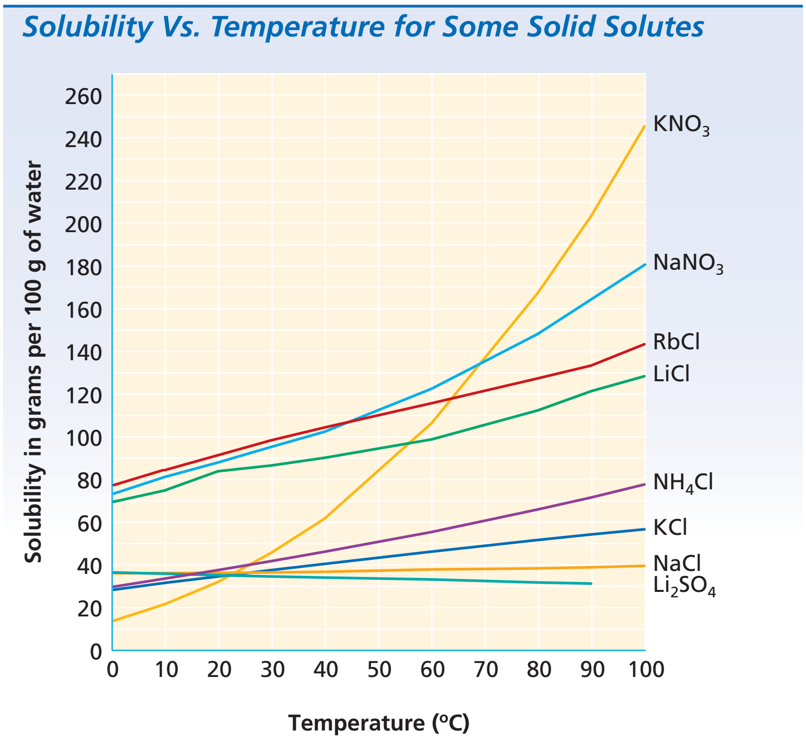


FIGURE 15 Solubility curves for various solid solutes generally show increasing solubility with increases in temperature. From the graph, you can see that the solubility of $NaNO_3$ is affected more by temperature than is NaCl.

in a large increase in solubility for some solvents and only a slight change for others.

In **Table 4** and **Figure 15,** compare the effect of temperature on the solubilities of potassium nitrate, KNO_3, and sodium chloride, NaCl. About 14 g of potassium nitrate will dissolve 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. In some cases, 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.

Enthalpies of Solution

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 the formation of a solution, solvent and solute particles experience changes in the forces attracting them to other particles. Before dissolving begins, solvent molecules are held together by intermolecular forces (solvent-solvent attraction). In the solute, molecules are held together by intermolecular forces (solute-solute attraction). Energy is required to separate solute molecules and solvent molecules from their neighbors. *A solute particle that is surrounded by solvent molecules,* as shown by the model in **Figure 9,** *is said to be* **solvated.**

Solution formation can be pictured as the result of the three interactions summarized in **Figure 16.**

FIGURE 16 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?

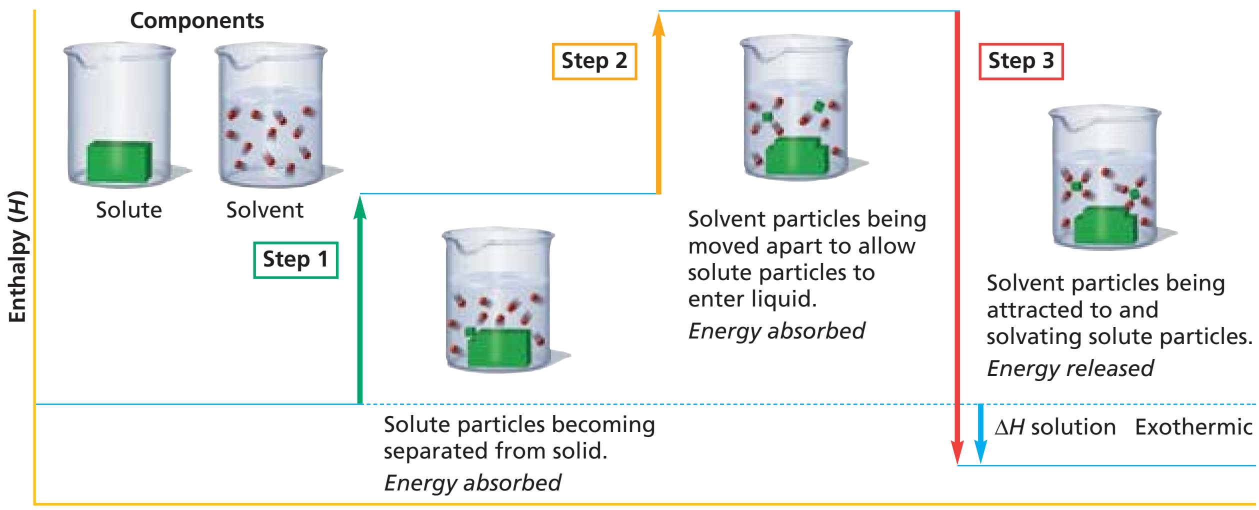


TABLE 5 Enthalpies of Solution (kJ/mol solute at 25°C)

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

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.** From the model in **Figure 16,** 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 Step 3. The enthalpy of solution is positive (energy is absorbed) when the sum of attractions from Steps 1 and 2 is greater than Step 3.

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 **Table 5** support this idea of exothermic solution processes for gaseous solutes?

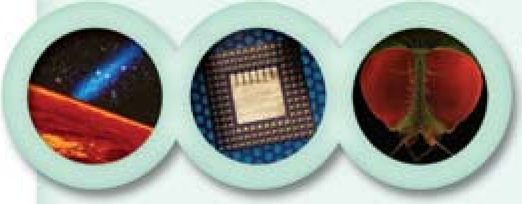
In the gaseous state, molecules are so far apart that there are virtually no intermolecular forces of attraction between them. Therefore, the solute-solute interaction has little effect on the enthalpy of a solution of a gas. Energy is released when a gas dissolves in a liquid because attraction between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

SECTION REVIEW

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 then make it a supersaturated solution?
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 your answer.
7. **INTERPRETING CONCEPTS** A commercial "fizz saver" pumps helium under pressure into a soda bottle to keep gas from escaping. Will this keep CO_2 in the soda bottle? Explain your answer.



CROSS-DISCIPLINARY CONNECTION

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!" According to Dr. Peter Keipert, Program Director of Oxygen Carriers Development at Alliance Pharmaceutical Corp., this 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. Dr. Keipert's blood substitute accomplishes 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. The blood-substitute mixture, called Oxygent™, is administered to a patient in the same way regular blood is. The perfluorocarbons are eventually exhaled through the lungs.

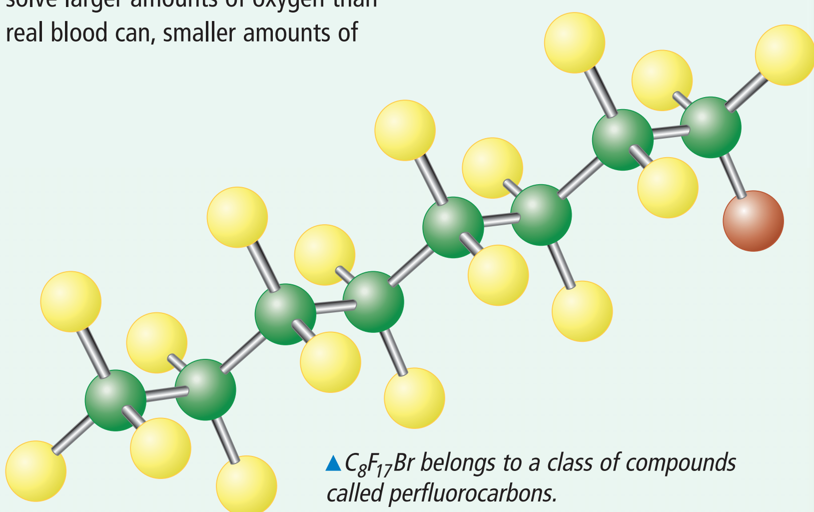
Oxygent only functions to carry gases to and from tissues; it cannot clot or perform any of the immune-system functions that blood does. Still, the substitute has several advantages over real blood. Oxygent has a shelf life of more than a year. Oxygent also eliminates many of the risks associated with blood transfusions. Because the substitute can dissolve larger amounts of oxygen than real blood can, smaller amounts of the mixture are needed.

Oxygent is currently being tested in surgical patients.

"Once this product is approved and has been demonstrated to be safe and effective in elective surgery, I think you will see its use spread into the emergency critical-care arena," says Dr. Keipert. "A patient who has lost a lot of blood and who is currently being resuscitated with normal fluids like saline solutions would be given Oxygent as an additional oxygen-delivery agent in the emergency room."

Questions

1. How would the approval of Oxygent benefit the medical community?
2. How do scientists prevent the nonpolar perfluorocarbons in Oxygent from separating from the water?



▲ $C_8F_{17}Br$ belongs to a class of compounds called perfluorocarbons.

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For a variety of links related to this chapter, go to www.scilinks.org

Topic: Perfluorocarbons
SciLinks code: HC61123

SECTION 3

Objectives

- Given the mass of solute and volume of solvent, calculate the concentration of a solution.
- Given the concentration of a solution, determine the amount of solute in a given amount of solution.
- Given the concentration of a solution, determine the amount of solution that contains a given amount of solute.

Concentration of Solutions

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.

Molarity

Molarity *is the number of moles of solute in one liter of solution*. To relate the molarity of a solution to the mass of solute present, you must know the molar mass of the solute. 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 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.

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

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

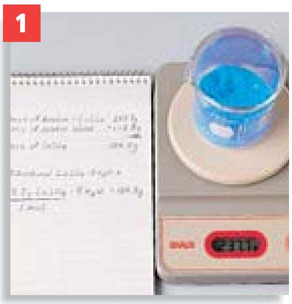
$$= 0.500 \text{ M NaOH}$$

If twice the molar mass of NaOH, 80.0 g, is dissolved in enough water to make 1 L of solution, a 2 M solution is produced. The molarity of any solution can be calculated by dividing the number of moles of solute by the number of liters of solution.

Note that a 1 M solution is *not* made by adding 1 mol of solute to 1 L of *solvent*. In such a case, the final total volume of the solution might not be 1 L. Instead, 1 mol of solute is first dissolved in less than 1 L of solvent. Then, the resulting solution is carefully diluted with more solvent to bring the *total volume* to 1 L, as shown in **Figure 17.** The following sample problem will show you how molarity is often used.

FIGURE 17 The preparation of a 0.5000 M solution of $CuSO_4{\cdot}5H_2O$ starts with calculating the mass of solute needed.

1



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 124.8 g.

2

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

3

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.

4

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

5

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

6

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

7

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

SAMPLE PROBLEM A

For more help, go to the ***Math Tutor*** *at the end of this chapter.*

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

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.

$$\text{grams of solute} \longrightarrow \text{number of moles of solute} \longrightarrow \text{molarity}$$

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

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

3 COMPUTE

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

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

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

4 EVALUATE

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.

SAMPLE PROBLEM B

For more help, go to the ***Math Tutor*** *at the end of this chapter.*

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

SOLUTION

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.

$$\text{concentration (mol of HCl/L of solution)} \times \text{volume (L of solution)} = \text{mol of HCl}$$

3 COMPUTE

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

4 EVALUATE 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.

SAMPLE PROBLEM C

For more help, go to the ***Math Tutor*** *at the end of this chapter.*

To produce 40.0 g of silver chromate, you will need at least 23.4 g of potassium chromate in solution as a reactant. 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?

SOLUTION

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
mass of product = 40.0 g Ag_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.

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

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

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

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

$$6.0 \text{ M } K_2CrO_4 = \frac{0.120 \text{ mol } K_2CrO_4}{x \text{ L } K_2CrO_4 \text{ soln}}$$

$$x = 0.020 \text{ L } K_2CrO_4 \text{ soln}$$

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

PRACTICE

Answers in Appendix E

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?

extension

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

Keyword: HC6SLNX

Molality

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.

$$\text{molality} = \frac{\text{moles solute}}{\text{mass of solvent (kg)}}$$

$$\frac{0.500 \text{ mol NaOH}}{1 \text{ kg } H_2O} = 0.500\ m \text{ NaOH}$$

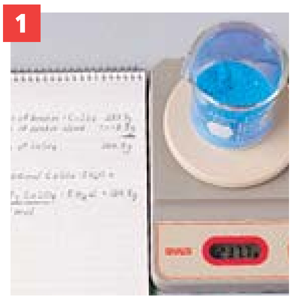
If 80.0 g of sodium hydroxide, which is 2 mol, is dissolved in 1 kg of water, a 2.00 *m* solution of NaOH is produced. The molality of any solution can be found by dividing the number of moles of solute by the mass in kilograms of the solvent in which it is dissolved. Note that if the amount of solvent is expressed in grams, the mass of solvent must be converted to kilograms by multiplying by the following conversion factor.

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

Figure 18 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 17.**

FIGURE 18 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 124.8 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 temperature changes. Molality is used because it does not change with changes in temperature. Below is a comparison of the equations for molarity and molality.

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

$$\text{molality, } m = \frac{\text{amount of A (mol)}}{\text{mass of solvent (kg)}}$$

SAMPLE PROBLEM D

A solution was prepared by dissolving 17.1 g of sucrose (table sugar, $C_{12}H_{22}O_{11}$) in 125 g of water. Find the molal concentration of this solution.

SOLUTION

1 ANALYZE

Given: solute mass = 17.1 g $C_{12}H_{22}O_{11}$
solvent mass = 125 g H_2O

Unknown: molal concentration

2 PLAN

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

$$\text{mol } C_{12}H_{22}O_{11} = \frac{\text{g } C_{12}H_{22}O_{11}}{\text{molar mass } C_{12}H_{22}O_{11}}$$

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

$$\text{molality } C_{12}H_{22}O_{11} = \frac{\text{mol } C_{12}H_{22}O_{11}}{\text{kg } H_2O}$$

3 COMPUTE

Use the periodic table to compute the molar mass of $C_{12}H_{22}O_{11}$.
$C_{12}H_{22}O_{11}$ = 342.34 g/mol

$$17.1 \text{ } \cancel{\text{g } C_{12}H_{22}O_{11}} \times \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.34 \text{ } \cancel{\text{g } C_{12}H_{22}O_{11}}} = 0.0500 \text{ mol } C_{12}H_{22}O_{11}$$

$$\frac{125 \text{ } \cancel{\text{g}} \text{ } H_2O}{1000 \text{ } \cancel{\text{g}}\text{/kg}} = 0.125 \text{ kg } H_2O$$

$$\frac{0.0500 \text{ mol } C_{12}H_{22}O_{11}}{0.125 \text{ kg } H_2O} = 0.400 \text{ } m \text{ } C_{12}H_{22}O_{11}$$

4 EVALUATE

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

SAMPLE PROBLEM E

A solution of iodine, I_2, in carbon tetrachloride, CCl_4, is used when iodine is needed for certain chemical tests. How much iodine must be added to prepare a 0.480 *m* solution of iodine in CCl_4 if 100.0 g of CCl_4 is used?

SOLUTION

1 ANALYZE

Given: molality of solution = 0.480 m I_2

mass of solvent = 100.0 g CCl_4

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 I_2.

3 COMPUTE

Use the periodic table to compute the molar mass of I_2.

I_2 = 253.8 g/mol

$$100.0\ \cancel{\text{g CCl}_4} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g CCl}_4}} = 0.100\ \text{kg CCl}_4$$

$$0.480\ m = \frac{x\ \text{mol I}_2}{0.1\ \text{kg H}_2\text{O}} \qquad x = 0.0480\ \text{mol I}_2$$

$$0.0480\ \cancel{\text{mol I}_2} \times \frac{253.8\ \text{g I}_2}{\cancel{\text{mol I}_2}} = 12.2\ \text{g I}_2$$

4 EVALUATE

The answer has three significant digits and the units for mass of I_2.

PRACTICE

Answers in Appendix E

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 quantity, in grams, of methanol, CH_3OH, is required to prepare a 0.244 m solution in 400. g of water?

extension

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

Keyword: HC6SLNX

SECTION REVIEW

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?

CHAPTER HIGHLIGHTS

Types of Mixtures

Vocabulary
soluble
solution
solvent
solute
suspension
colloid
electrolyte
nonelectrolyte

- 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.

The Solution Process

Vocabulary
solution equilibrium
saturated solution
unsaturated solution
supersaturated solution
solubility
hydration
immiscible
miscible
Henry's law
effervescence
solvated
enthalpy of solution

- 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 change when a specified amount of solute dissolved during solution formation is called the enthalpy of solution.

Concentration of Solutions

Vocabulary
concentration
molarity
molality

- 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.

CHAPTER REVIEW

For more practice, go to the Problem Bank in Appendix D.

Types of Mixtures

SECTION 1 REVIEW

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 we 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?

The Solution Process

SECTION 2 REVIEW

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 would show more effervescence and why?

11. Based on **Figure 15,** 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 15,** 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 7?**

Concentration of Solutions

SECTION 3 REVIEW

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 produce 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 pure $AgNO_3$.
a. What is step 1 in solving the problem?
b. What is the molar mass of $AgNO_3$?
c. How many milliliters of solution are needed?

24. a. Balance the 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 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* in ethanol. How many grams of C_2H_5OH are contained per 250. g of water?

MIXED REVIEW

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 6.0 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 **Table 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?

CRITICAL THINKING

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 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	cloudy	none	yes
4	colorless	clear	white solid	no

USING THE HANDBOOK

40. Review the information on alloys in the Elements Handbook.
   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* 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 & 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 $CuSO_4{\cdot}5H_2O$ 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.

extension

Graphing Calculator

Predicting Solubility from Tabular Data

Go to **go.hrw.com** for a graphing calculator exercise that asks you to predict the solubility of a substance at various temperatures.

Keyword: HC6SLNX

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. You know that 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.

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

Problem-Solving *TIPS*

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

SAMPLE

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.

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

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

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 PROBLEMS

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