

CHAPTER GOALS

Among the questions we will answer are the following:

1. What are solutions, and what factors affect solubility?

THE GOAL: Be able to define the different kinds of mixtures and explain the influence on solubility of solvent and solute structure, temperature, and pressure.

2. How is the concentration of a solution expressed?

THE GOAL: Be able to define, use, and convert between the most common ways of expressing solution concentrations.

3. How are dilutions carried out?

THE GOAL: Be able to calculate the concentration of a solution prepared by dilution and explain how to make a desired dilution.

4. What is an electrolyte?

THE GOAL: Be able to recognize strong and weak electrolytes and nonelectrolytes, and express electrolyte concentrations.

5. How do solutions differ from pure solvents in their behavior?

THE GOAL: Be able to explain vapor pressure lowering, boiling point elevation, and freezing point depression for solutions.

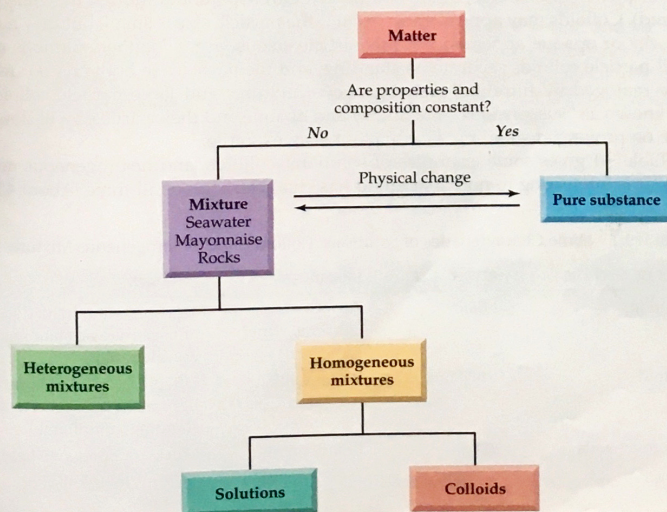
6. What is osmosis?

THE GOAL: Be able to describe osmosis and some of its applications.

Up to this point, we have been concerned primarily with pure substances, both elements and compounds. In day-to-day life, however, most of the materials we come in contact with are mixtures. Air, for example, is a gaseous mixture of primarily oxygen and nitrogen; blood is a liquid mixture of many different components; and many rocks are solid mixtures of different minerals. In this chapter, we look closely at the characteristics and properties of mixtures, with particular attention to the uniform mixtures we call *solutions*.

9.1 Mixtures and Solutions

As we saw in Section 1.3, a *mixture* is an intimate combination of two or more substances, both of which retain their chemical identities. (p. 6) Mixtures can be classified as either *heterogeneous* or *homogeneous* as indicated in Figure 9.1, depending on their appearance. **Heterogeneous mixtures** are those in which the mixing is



Heterogeneous mixture A nonuniform mixture that has regions of different composition.

FIGURE 9.1 Classification of mixtures. The components in heterogeneous mixtures are not uniformly mixed, and the composition varies with location. In homogeneous mixtures, the components are uniformly mixed at the molecular level.

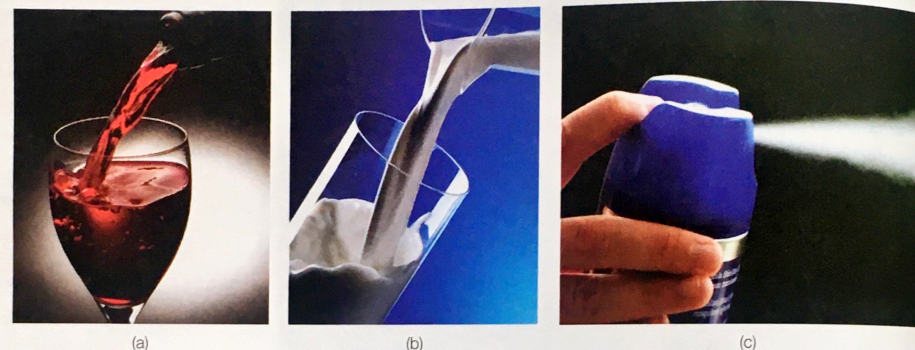
not uniform and which therefore have regions of different composition. Rocky Road ice cream, for example, is a heterogeneous mixture, with something different in every spoonful. Granite and many other rocks are also heterogeneous, having a grainy character due to the heterogeneous mixing of different minerals. **Homogeneous mixtures** are those in which the mixing is uniform and that therefore have the same composition throughout. Seawater, a homogeneous mixture of soluble ionic compounds in water, is an example.

Homogeneous mixture A uniform mixture that has the same composition throughout.

Solution A homogeneous mixture that contains particles the size of a typical ion or small molecule.

Colloid A homogeneous mixture that contains particles that range in diameter from 2 to 500 nm.

Homogeneous mixtures can be further classified as either *solutions* or *colloids* according to the size of their particles. **Solutions**, the most important class of homogeneous mixtures, contain particles the size of a typical ion or small molecule—roughly 0.1–2 nm in diameter. **Colloids**, such as milk and fog, are also homogeneous in appearance but contain larger particles than solutions—in the range 2–500 nm diameter.



▲ (a) Wine is a solution of dissolved molecules, and (b) milk is a colloid with fine particles that do not separate out on standing. (c) An aerosol spray, by contrast, is a heterogeneous mixture of small particles visible to the naked eye.

Liquid solutions, colloids, and heterogeneous mixtures can be distinguished in several ways. For example, liquid solutions are transparent (although they may be colored). Colloids may appear transparent if the particle size is small, but they have a murky or opaque appearance if the particle size is larger. Neither solutions nor small-particle colloids separate on standing, and the particles in both are too small to be removed by filtration. Heterogeneous mixtures and large-particle colloids, also known as “suspensions,” are murky and opaque and their particles will slowly settle on prolonged standing. House paint is an example.

Table 9.1 gives some examples of solutions, colloids, and heterogeneous mixtures. It is interesting to note that blood has characteristics of all three. About 45%

TABLE 9.1 Some Characteristics of Solutions, Colloids, and Heterogeneous Mixtures

TYPE OF MIXTURE	PARTICLE SIZE	EXAMPLES	CHARACTERISTICS
Solution	<2.0 nm	Air, seawater, gasoline, wine	Transparent to light; does not separate on standing; nonfilterable
Colloid	2.0–500 nm	Butter, milk, fog, pearl	Often murky or opaque to light; does not separate on standing; nonfilterable
Heterogeneous	>500 nm	Blood, paint, aerosol sprays	Murky or opaque to light; separates on standing; filterable

by volume of blood consists of suspended red and white cells, which settle slowly on standing; the remaining 55% is *plasma*, which contains ions in solution and colloidal protein molecules.

Although we usually think of solids dissolved in liquids when we talk about solutions, solutions actually occur in all three phases of matter (Table 9.2). Metal alloys like 14-karat gold (58% gold with silver and copper) and brass (10–40% zinc with copper), for instance, are solutions of one solid with another. For solutions in which a gas or solid is dissolved in a liquid, the dissolved substance is called the **solute** and the liquid is called the **solvent**. When one liquid is dissolved in another, the minor component is usually considered the solute and the major component is the solvent.

TABLE 9.2 Some Different Types of Solutions

TYPE OF SOLUTION	EXAMPLE
Gas in gas	Air (O ₂ , N ₂ , Ar, and other gases)
Gas in liquid	Seltzer water (CO ₂ in water)
Gas in solid	H ₂ in palladium metal
Liquid in liquid	Gasoline (mixture of hydrocarbons)
Liquid in solid	Dental amalgam (mercury in silver)
Solid in liquid	Seawater (NaCl and other salts in water)
Solid in solid	Metal alloys such as 14-karat gold (Au, Ag, and Cu)

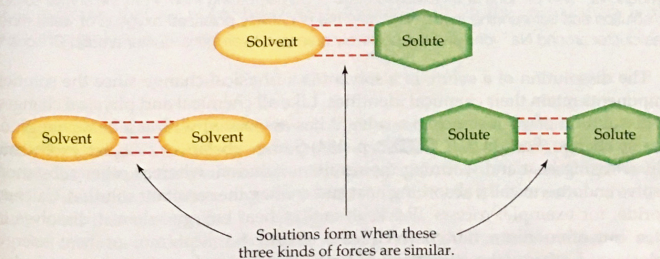
PROBLEM 9.1

Classify the following liquid mixtures as heterogeneous or homogeneous. Further classify each homogeneous mixture as a solution or colloid.

- (a) Orange juice (b) Apple juice
(c) Hand lotion (d) Tea

9.2 The Solution Process

What determines whether a substance is soluble in a given liquid? Solubility depends primarily on the strength of the attractions between solute and solvent particles relative to the strengths of the attractions within the pure substances. Ethyl alcohol is soluble in water, for example, because hydrogen bonding (Section 8.11) is nearly as strong between water and ethyl alcohol molecules as it is between water molecules alone or ethyl alcohol molecules alone. (p. 238)



A good rule of thumb for predicting solubility is that “like dissolves like,” meaning that substances with similar intermolecular forces form solutions with one another, whereas substances with different intermolecular forces do not (Section 8.11). (p. 235)

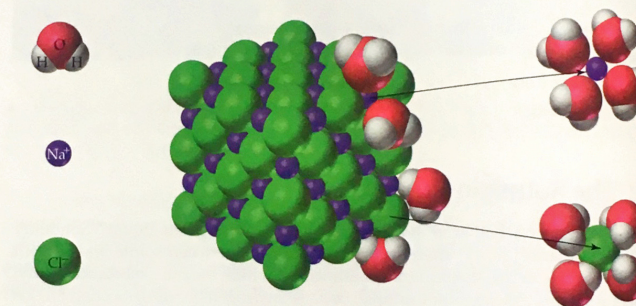
Solute A substance dissolved in a liquid.

Solvent The liquid in which another substance is dissolved.



▲ Oil and water do not mix because they have different intermolecular forces, resulting in the formation of oil slicks.

Solvation The clustering of solvent molecules around a dissolved solute molecule or ion.



▲ **FIGURE 9.2** Dissolution of an NaCl crystal in water. Polar water molecules surround the individual Na⁺ and Cl⁻ ions at an exposed edge or corner, pulling them from the crystal surface into solution and surrounding them. Note how the negatively polarized oxygens of water molecules cluster around Na⁺ ions and the positively polarized hydrogens cluster around Cl⁻ ions.



▲ Instant cold packs used to treat muscle strains and sprains often take advantage of the endothermic enthalpy of a solution of salts such as ammonium nitrate.

The dissolution of a solute in a solvent is a physical change since the solution components retain their chemical identities. Like all chemical and physical changes, the dissolution of a substance in a solvent has associated with it a heat change, or *enthalpy* change (Section 7.2). (p. 184) Some substances dissolve exothermically, releasing heat and warming the resultant solution, whereas other substances dissolve endothermically, absorbing heat and cooling the resultant solution. Calcium chloride, for example, releases 19.4 kcal/mol of heat energy when it dissolves in water, but ammonium nitrate (NH₄NO₃) absorbs 6.1 kcal/mol of heat energy. Athletes and others take advantage of both situations when they use instant hot packs or cold packs to treat injuries. Both hot and cold packs consist of a pouch of water and a dry chemical, such as CaCl₂ or MgSO₄ for hot packs and NH₄NO₃ for cold packs. Squeezing the pack breaks the pouch and the solid dissolves, either raising or lowering the temperature.

WORKED EXAMPLE 9.1 Formation of Solutions

Which of the following pairs of substances would you expect to form solutions?

- (a) Carbon tetrachloride (CCl_4) and hexane (C_6H_{14}).
 (b) Octane (C_8H_{18}) and methyl alcohol (CH_3OH).

ANALYSIS Identify the kinds of intermolecular forces in each substance (Section 8.11). Substances with similar intermolecular forces tend to form solutions.

SOLUTION

- (a) Hexane contains only C—H and C—C bonds, which are nonpolar. Carbon tetrachloride contains polar C—Cl bonds, but they are distributed symmetrically in the tetrahedral molecule so that it too is nonpolar. The major intermolecular force for both compounds is London dispersion forces, so they will form a solution.
 (b) Octane contains only C—H and C—C bonds and so is nonpolar; the major intermolecular force is dispersion. Methyl alcohol contains polar C—O and O—H bonds; it is polar and forms hydrogen bonds. The intermolecular forces for the two substances are so dissimilar that they do not form a solution.

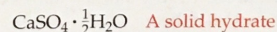
PROBLEM 9.2

Which of the following pairs of substances would you expect to form solutions?

- (a) CCl_4 and water
 (b) Benzene (C_6H_6) and MgSO_4
 (c) Hexane (C_6H_{14}) and heptane (C_7H_{16})
 (d) Ethyl alcohol ($\text{C}_2\text{H}_5\text{OH}$) and heptanol ($\text{C}_7\text{H}_{15}\text{OH}$)

9.3 Solid Hydrates

Some ionic compounds attract water strongly enough to hold onto water molecules even when crystalline, forming what are called *solid hydrates*. For example, the plaster of Paris used to make decorative objects and casts for broken limbs is calcium sulfate hemihydrate, $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$. The dot between CaSO_4 and $\frac{1}{2}\text{H}_2\text{O}$ in the formula indicates that for every two CaSO_4 formula units in the crystal there is also one water molecule present.



After being ground up and mixed with water to make plaster, $\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$ gradually changes into the crystalline dihydrate $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, known as *gypsum*. During the change, the plaster hardens and expands in volume, causing it to fill a mold or shape itself closely around a broken limb. Table 9.3 lists some other ionic compounds that are handled primarily as hydrates.

Still other ionic compounds attract water so strongly that they pull water vapor from humid air to become hydrated. Compounds that show this behavior, such as calcium chloride (CaCl_2), are called **hygroscopic** and are often used as drying agents. You might have noticed a small bag of a hygroscopic compound (probably silica gel, SiO_2) included in the packing material of a new MP3 player, camera, or other electronic device to keep humidity low during shipping.

Hygroscopic Having the ability to pull water molecules from the surrounding atmosphere.



▲ Plaster of Paris ($\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$) slowly turns into gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$) when added to water. In so doing, the plaster hardens and expands, causing it to fill a mold.

TABLE 9.3 Some Common Solid Hydrates

FORMULA	NAME	USES
$\text{AlCl}_3 \cdot 6\text{H}_2\text{O}$	Aluminum chloride hexahydrate	Antiperspirant
$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	Calcium sulfate dihydrate (gypsum)	Cements, wallboard, molds
$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$	Calcium sulfate hemihydrate (plaster of Paris)	Casts, molds
$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	Copper(II) sulfate pentahydrate (blue vitriol)	Pesticide, germicide, topical fungicide
$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	Magnesium sulfate heptahydrate (Epsom salts)	Laxative, anticonvulsant
$\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$	Sodium tetraborate decahydrate (borax)	Cleaning compounds, fireproofing agent
$\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$	Sodium thiosulfate pentahydrate (hypo)	Photographic fixer

PROBLEM 9.3

Write the formula of sodium sulfate decahydrate, known as Glauber's salt and used as a laxative.

PROBLEM 9.4

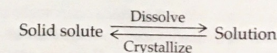
What masses of Glauber's salt must be used to provide 1.00 mol of sodium sulfate?

9.4 Solubility

We saw in Section 9.2 that ethyl alcohol is soluble in water because hydrogen bonding is nearly as strong between water and ethyl alcohol molecules as it is between water molecules alone or ethyl alcohol molecules alone. So similar are the forces in this particular case, in fact, that the two liquids are **miscible**, or mutually soluble in all proportions. Ethyl alcohol will continue to dissolve in water no matter how much is added.

Most substances, however, reach a solubility limit beyond which no more will dissolve in solution. Imagine, for instance that you are asked to prepare a saline solution (aqueous NaCl). You might measure out some water, add solid NaCl, and stir the mixture. Dissolution occurs rapidly at first but then slows down as more and more NaCl is added. Eventually the dissolution stops because an equilibrium is reached when the numbers of Na^+ and Cl^- ions leaving a crystal and going into solution are equal to the numbers of ions returning from solution to the crystal. At this point, the solution is said to be **saturated**. A maximum of 35.8 g of NaCl will dissolve in 100 mL of water at 20 °C. Any amount above this limit simply sinks to the bottom of the container and sits there.

The equilibrium reached by a saturated solution is like the equilibrium reached by a reversible reaction (Section 7.7). (Section 7.7, p. 198) Both are dynamic situations in which no *apparent* change occurs because the rates of forward and backward processes are equal. Solute particles leave the solid surface and reenter the solid from solution at the same rate.



Miscible Mutually soluble in all proportions.

Saturated solution A solution that contains the maximum amount of dissolved solute at equilibrium.

Solubility The maximum amount of a substance that will dissolve in a given amount of solvent at a specified temperature.

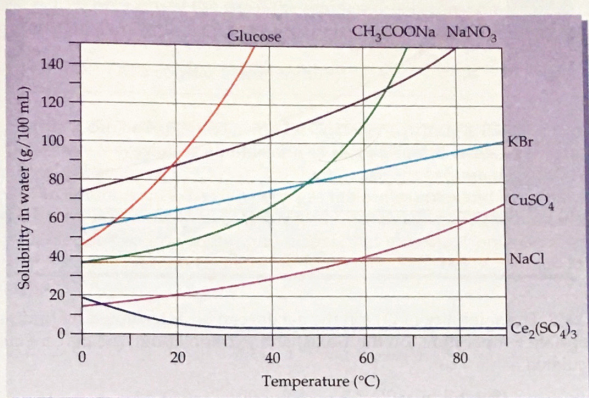
The maximum amount of a substance that will dissolve in a given amount of a solvent at a given temperature, usually expressed in grams per 100 mL (g/100 mL), is called the substance's **solubility**. Solubility is a characteristic property of a specific

solute-solvent combination, and different substances have greatly differing solubilities. Only 9.6 g of sodium hydrogen carbonate will dissolve in 100 mL of water at 20 °C, for instance, but 204 g of sucrose will dissolve under the same conditions.

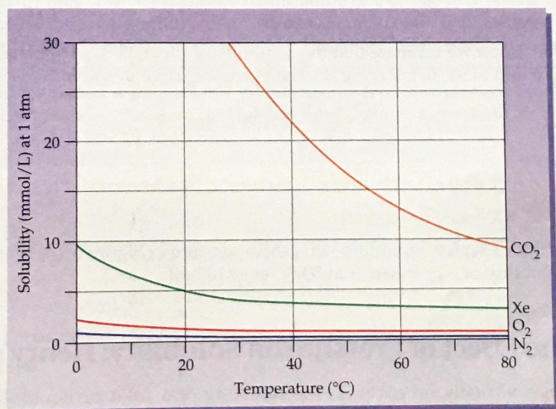
9.5 The Effect of Temperature on Solubility

As anyone who has ever made tea or coffee knows, temperature often has a dramatic effect on solubility. The compounds in tea leaves or coffee beans, for instance, dissolve easily in hot water but not in cold water. The effect of temperature is different for every substance, however, and is usually unpredictable. As shown in Figure 9.3(a), the solubilities of most molecular and ionic solids increase with increasing temperature, but the solubilities of others (NaCl) are almost unchanged, and the solubilities of still others [$\text{Ce}_2(\text{SO}_4)_3$] decrease with increasing temperature.

Solids that are more soluble at high temperature than at low temperature can sometimes form what are called **supersaturated solutions**, which contain even more solute than a saturated solution. Suppose, for instance, that a large amount of a substance is dissolved at a high temperature. As the solution cools, the solubility decreases and the excess solute should precipitate to maintain equilibrium. But if



(a)



(b)

◀ **FIGURE 9.3** Solubilities of some (a) solids and (b) gases, in water as a function of temperature. Most solid substances become more soluble as temperature rises (although the exact relationship is usually complex), while the solubility of gases decreases.

Supersaturated solution A solution that contains more than the maximum amount of dissolved solute; a nonequilibrium situation.



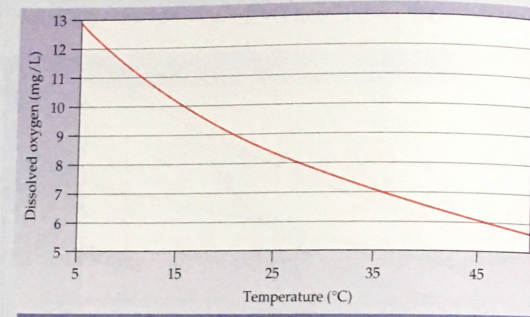
▲ **FIGURE 9.4** A supersaturated solution of sodium acetate in water. When a tiny seed crystal is added, larger crystals rapidly grow and precipitate from the solution until equilibrium is reached.

the cooling is done very slowly, and if the container stands quietly, crystallization might not occur immediately and a supersaturated solution might result. Such a solution is unstable, however, and precipitation can occur dramatically when a tiny seed crystal is added to initiate crystallization (Figure 9.4).

Unlike solids, the influence of temperature on the solubility of gases is predictable: Addition of heat decreases the solubility of most gases, as seen in Figure 9.3(b) (helium is the only common exception). One result of this temperature-dependent decrease in gas solubility can sometimes be noted in a stream or lake near the outflow of warm water from an industrial operation. As water temperature increases, the concentration of dissolved oxygen in the water decreases, killing fish that cannot tolerate the lower oxygen levels.

WORKED EXAMPLE 9.2 Solubility of Gases: Effect of Temperature

From the following graph of solubility versus temperature for O_2 , estimate the concentration of dissolved oxygen in water at 25 °C and at 35 °C. By what percentage does the concentration of O_2 change?



ANALYSIS The solubility of O_2 (on the y -axis) can be determined by finding the appropriate temperature (on the x -axis) and extrapolating. The percent change is calculated as

$$\frac{(\text{Solubility at } 25^\circ\text{C}) - (\text{Solubility at } 35^\circ\text{C})}{(\text{Solubility at } 25^\circ\text{C})} \times 100$$

SOLUTION

From the graph we estimate that the solubility of O_2 at 25 °C is approximately 8.3 mg/L and at 35 °C is 7.0 mg/L. The percent change in solubility is

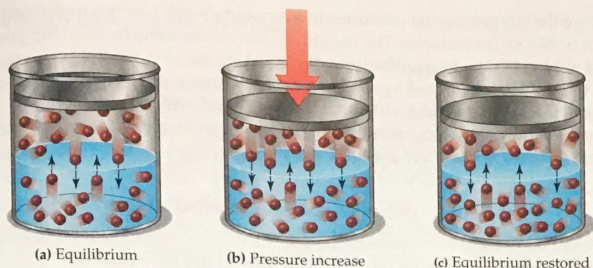
$$\frac{8.3 - 7.0}{8.3} \times 100 = 16\%$$

PROBLEM 9.5

Look at the graph of solubility versus temperature in Figure 9.3, and estimate the solubility of KBr in water at 50 °C in g/100 mL.

9.6 The Effect of Pressure on Solubility: Henry's Law

Pressure has virtually no effect on the solubility of a solid or liquid, but it has a strong effect on the solubility of a gas. According to **Henry's law**, the solubility (or concentration) of a gas in a liquid is directly proportional to the partial pressure



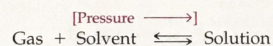
◀ **FIGURE 9.5** Henry's law. The solubility of a gas is directly proportional to its partial pressure. An increase in pressure causes more gas molecules to enter solution until equilibrium is restored between the dissolved and undissolved gas.

of the gas over the liquid. (Recall from Section 8.10 that each gas in a mixture exerts a partial pressure independent of other gases present (□□, p. 233). If the partial pressure of the gas doubles, solubility doubles; if the gas pressure is halved, solubility is halved (Figure 9.5).

Henry's law The solubility (or concentration) of a gas is directly proportional to the partial pressure of the gas if the temperature is constant. That is, concentration (C) divided by pressure (P) is constant when T is constant,

$$\text{or } \frac{C}{P_{\text{gas}}} = k \quad (\text{At a constant temperature})$$

Henry's law can be explained using Le Châtelier's principle (Section 7.9), which states that when a system at equilibrium is placed under stress, the equilibrium shifts to relieve that stress. (□□, p. 203) In the case of a saturated solution of a gas in a liquid, an equilibrium exists whereby gas molecules enter and leave the solution at the same rate. When the system is stressed by increasing the pressure of the gas, more gas molecules go into solution to relieve that increase. Conversely, when the pressure of the gas is decreased, more gas molecules come out of solution to relieve the decrease.



As an example of Henry's law in action, think about the fizzing that occurs when you open a bottle of soft drink or champagne. The bottle is sealed under greater than 1 atm of CO_2 pressure, causing some of the CO_2 to dissolve. When the bottle is opened, however, CO_2 pressure drops and gas comes fizzing out of solution.

Writing Henry's law in the form $P_{\text{gas}} = C/k$ shows that partial pressure can be used to express the concentration of a gas in a solution, a practice especially common in health-related sciences. Table 9.4 gives some typical values and illustrates the convenience of having the same unit for concentration of a gas in both air and blood.

TABLE 9.4 Partial Pressures and Normal Gas Concentrations in Body Fluids

SAMPLE	PARTIAL PRESSURE (mmHg)			
	P_{N_2}	P_{O_2}	P_{CO_2}	$P_{\text{H}_2\text{O}}$
Inspired air (dry)	597	159	0.3	3.7
Alveolar air (saturated)	573	100	40	47
Expired air (saturated)	569	116	28	47
Arterial blood	573	95	40	
Venous blood	573	40	45	
Peripheral tissues	573	40	45	



▲ The CO_2 gas dissolved under pressure comes out of solution when the bottle is opened and the pressure drops.

Compare the oxygen partial pressures in saturated alveolar air (air in the lungs) and in arterial blood, for instance. The values are almost the same because the gases dissolved in blood come to equilibrium with the same gases in the lungs.

If the partial pressure of a gas over a solution changes while the temperature is constant, the new solubility of the gas can be found easily. Because C/P is a constant value at constant temperature, Henry's law can be restated to show how one variable changes if the other changes:

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} = k \quad (\text{Where } k \text{ is constant at a fixed temperature})$$

Worked Example 9.3 gives an illustration of how to use this equation.

WORKED EXAMPLE 9.3 Solubility of Gases: Henry's Law

At a partial pressure of oxygen in the atmosphere of 159 mmHg, the solubility of oxygen in blood is 0.44 g/100 mL. What is the solubility of oxygen in blood at 11,000 ft, where the partial pressure of O_2 is 56 mmHg?

ANALYSIS According to Henry's law, the solubility of the gas divided by its pressure is constant:

$$\frac{C_1}{P_1} = \frac{C_2}{P_2}$$

Of the four variables in this equation, we know P_1 , C_1 , and P_2 , and we need to find C_2 .

BALLPARK ESTIMATE The pressure drops by a factor of about 3 (from 159 mmHg to 56 mmHg). Since the ratio of solubility to pressure is constant, the solubility must also drop by a factor of 3 (from 0.44 g/100 mL to about 0.15 g/100 mL).

SOLUTION

STEP 1: Identify known information. We have values for P_1 , C_1 , and P_2 .

$$\begin{array}{l} P_1 = 159 \text{ mmHg} \\ C_1 = 0.44 \text{ g/100 mL} \\ P_2 = 56 \text{ mmHg} \end{array}$$

$$\text{Solubility of } \text{O}_2, C_2 = ?? \text{ g/100 mL}$$

STEP 2: Identify answer and units. We are looking for the solubility of O_2 (C_2) at a partial pressure P_2 .

STEP 3: Identify conversion factors or equations. In this case, we restate Henry's law to solve for C_2 .

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} \Rightarrow C_2 = \frac{C_1 P_2}{P_1}$$

STEP 4: Solve. Substitute the known values into the equation and calculate C_2 .

$$C_2 = \frac{C_1 P_2}{P_1} = \frac{(0.44 \text{ g/100 mL})(56 \text{ mmHg})}{159 \text{ mmHg}} = 0.15 \text{ g/100 mL}$$

BALLPARK CHECK: The calculated answer matches our estimate.

PROBLEM 9.6

At 20 °C and a partial pressure of 760 mmHg, the solubility of CO_2 in water is 0.169 g/100 mL. What is the solubility of CO_2 at 2.5×10^4 mmHg?

PROBLEM 9.7

At a total atmospheric pressure of 1.00 atm, the partial pressure of CO_2 in air is approximately 4.0×10^{-4} atm. Using the data in Problem 9.6, what is the solubility of CO_2 in an open bottle of seltzer water at 20 °C?

9.7 Units of Concentration

Although we speak casually of a solution of, say, orange juice as either “dilute” or “concentrated,” laboratory work usually requires an exact knowledge of a solution’s concentration. As indicated in Table 9.5, there are several common methods for expressing concentration. The units differ, but all the methods describe how much solute is present in a given quantity of solution.

TABLE 9.5 Some Units for Expressing Concentration

CONCENTRATION MEASURE	SOLUTE MEASURE	SOLUTION MEASURE
Molarity, M	Moles	Volume (L)
Weight/volume percent, (w/v)%	Weight (g)	Volume (mL)
Volume/volume percent, (v/v)%	Volume*	Volume*
Parts per million, ppm	Parts*	10 ⁶ parts*

*Any units can be used as long as they are the same for both solute and solution.

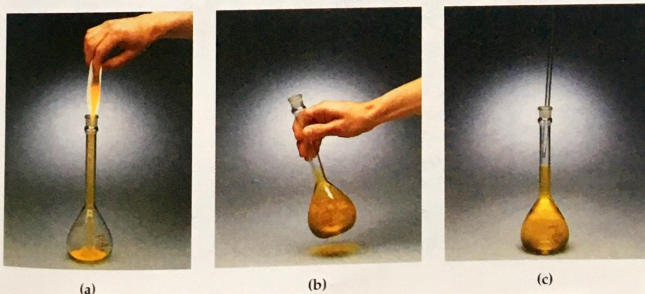
Let us look at each of the four concentration measures listed in Table 9.5 individually, beginning with *molarity*.

Mole/Volume Concentration: Molarity

We saw in Chapter 6 that the various relationships between amounts of reactants and products in chemical reactions are calculated in *moles* (Sections 6.4–6.6). Thus, the most generally useful means of expressing concentration in the laboratory is **molarity (M)**, the number of moles of solute dissolved per liter of solution. For example, a solution made by dissolving 1.00 mol (58.5 g) of NaCl in enough water to give 1.00 L of solution has a concentration of 1.00 mol/L, or 1.00 M. The molarity of any solution is found by dividing the number of moles of solute by the number of liters of solution (solute + solvent):

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Liters of solution}}$$

Note that a solution of a given molarity is prepared by dissolving the solute in enough solvent to give a *final* solution volume of 1.00 L, not by dissolving it in an *initial* volume of 1.00 L. If an initial volume of 1.00 L were used, the final solution volume might be a bit larger than 1.00 L because of the additional volume of the solute. In practice, the appropriate amount of solute is weighed and placed in a *volumetric flask*, as shown in Figure 9.6. Enough solvent is then added to dissolve the solute, and further solvent is added until an accurately calibrated final volume is reached. The solution is then shaken until it is uniformly mixed.



◀ **FIGURE 9.6** Preparing a solution of known molarity. (a) A measured number of moles of solute is placed in a volumetric flask. (b) Enough solvent is added to dissolve the solute by swirling. (c) Further solvent is carefully added until the calibration mark on the neck of the flask is reached, and the solution is shaken until uniform.

Molarity can be used as a conversion factor to relate the volume of a solution to the number of moles of solute it contains. If we know the molarity and volume of a solution, we can calculate the number of moles of solute. If we know the number of moles of solute and the molarity of the solution, we can find the solution’s volume.

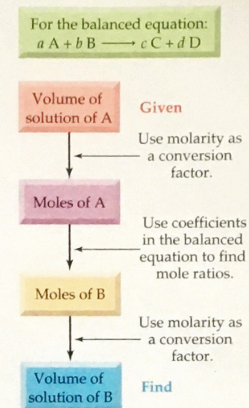
$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution (L)}}$$

$$\text{Moles of solute} = \text{Molarity} \times \text{Volume of solution}$$

$$\text{Volume of solution} = \frac{\text{Moles of solute}}{\text{Molarity}}$$

The flow diagram in Figure 9.7 shows how molarity is used in calculating the quantities of reactants or products in a chemical reaction, and Worked Examples 9.5 and 9.6 show how the calculations are done. Note that Problem 9.10 employs *millimolar* (mM) concentrations, which are useful in healthcare fields for expressing low concentrations such as are often found in body fluids (1 mM = 0.001 M).

► **FIGURE 9.7** Molarity and conversions. A flow diagram summarizing the use of molarity for conversions between solution volume and moles to find quantities of reactants and products for chemical reactions in solution.



WORKED EXAMPLE 9.4 Solution Concentration: Molarity

What is the molarity of a solution made by dissolving 2.355 g of sulfuric acid (H_2SO_4) in water and diluting to a final volume of 50.0 mL? The molar mass of H_2SO_4 is 98.1 g/mol.

ANALYSIS Molarity is defined as moles of solute per liter of solution: $M = \text{mol/L}$. Thus, we must first find the number of moles of sulfuric acid by doing a mass to mole conversion, and then divide the number of moles by the volume of the solution.

BALLPARK ESTIMATE The molar mass of sulfuric acid is about 100 g/mol, so 2.355 g is roughly 0.025 mol. The volume of the solution is 50.0 mL, or 0.05 L, so we have about 0.025 mol of acid in 0.05 L of solution, which is a concentration of about 0.5 M.

SOLUTION

STEP 1: Identify known information. We know the mass of sulfuric acid and the final volume of solution.

$$\begin{aligned} \text{Mass of H}_2\text{SO}_4 &= 2.355 \text{ g} \\ \text{Volume of solution} &= 50.0 \text{ mL} \end{aligned}$$

STEP 2: Identify answer including units. We need to find the molarity (M) in units of moles per liter.

$$\text{Molarity} = \frac{\text{Moles H}_2\text{SO}_4}{\text{Liters of solution}}$$

$$(2.355 \text{ g H}_2\text{SO}_4) \left(\frac{1 \text{ mol H}_2\text{SO}_4}{98.1 \text{ g H}_2\text{SO}_4} \right) = 0.0240 \text{ mol H}_2\text{SO}_4$$

STEP 3: Identify conversion factors and equations. We know both the amount of solute and the volume of solution, but first we must make two conversions: convert mass of H_2SO_4 to moles of H_2SO_4 , using molar mass as a conversion factor, and convert volume from milliliters to liters:

$$(50.0 \text{ mL}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.0500 \text{ L}$$

STEP 4: Solve. Substitute the moles of solute and volume of solution into the molarity expression.

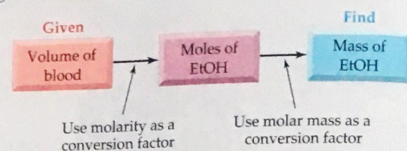
$$\text{Molarity} = \frac{0.0240 \text{ mol H}_2\text{SO}_4}{0.0500 \text{ L}} = 0.480 \text{ M}$$

BALLPARK CHECK: The calculated answer is close to our estimate, which was 0.5 M.

WORKED EXAMPLE 9.5 Molarity as Conversion Factor: Molarity to Mass

A blood concentration of 0.065 M ethyl alcohol (EtOH) is sufficient to induce a coma. At this concentration, what is the total mass of alcohol (in grams) in an adult male whose total blood volume is 5.6 L? The molar mass of ethyl alcohol is 46.0 g/mol. (Refer to the flow diagram in Figure 9.7 to identify which conversions are needed.)

ANALYSIS We are given a molarity (0.065 M) and a volume (5.6 L), which allows us to calculate the number of moles of alcohol in the blood. A mole to mass conversion then gives the mass of alcohol.



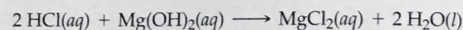
SOLUTION

$$(5.6 \text{ L blood}) \left(\frac{0.065 \text{ mol EtOH}}{1 \text{ L blood}} \right) = 0.36 \text{ mol EtOH}$$

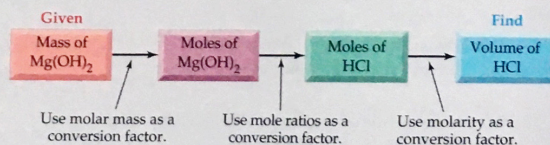
$$(0.36 \text{ mol EtOH}) \left(\frac{46.0 \text{ g EtOH}}{1 \text{ mol EtOH}} \right) = 17 \text{ g EtOH}$$

WORKED EXAMPLE 9.6 Molarity as Conversion Factor: Molarity to Volume

In our stomachs, gastric juice that is about 0.1 M in HCl aids in digestion. How many milliliters of gastric juice will react completely with an antacid tablet that contains 500 mg of magnesium hydroxide? The molar mass of $\text{Mg}(\text{OH})_2$ is 58.3 g/mol, and the balanced equation is



ANALYSIS We are given the molarity of HCl and need to find the volume. We first convert the mass of $\text{Mg}(\text{OH})_2$ to moles and then use the coefficients in the balanced equation to find the moles of HCl that will react. Once we have the moles of HCl and the molarity in moles per liter, we can find the volume. These conversions are summarized in the following flow diagram.



SOLUTION

$$[500 \text{ mg Mg}(\text{OH})_2] \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right) \left(\frac{1 \text{ mol Mg}(\text{OH})_2}{58.3 \text{ g Mg}(\text{OH})_2} \right) = 0.00858 \text{ mol Mg}(\text{OH})_2$$

$$[0.00858 \text{ mol Mg}(\text{OH})_2] \left(\frac{2 \text{ mol HCl}}{1 \text{ mol Mg}(\text{OH})_2} \right) \left(\frac{1 \text{ L HCl}}{0.1 \text{ mol HCl}} \right) = 0.2 \text{ L (200 mL)}$$

PROBLEM 9.8

What is the molarity of a solution that contains 50.0 g of vitamin B₁ hydrochloride (molar mass = 337 g/mol) in 160 mL of solution?

PROBLEM 9.9

How many moles of solute are present in the following solutions?

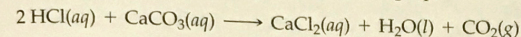
- (a) 175 mL of 0.35 M NaNO_3 (b) 480 mL of 1.4 M HNO_3

PROBLEM 9.10

The concentration of cholesterol ($\text{C}_{27}\text{H}_{46}\text{O}$) in blood is approximately 5.0 mM. How many grams of cholesterol are in 250 mL of blood?

PROBLEM 9.11

What mass (in grams) of calcium carbonate is needed to react completely with 65 mL of 0.12 M HCl according to the following equation?



Weight/Volume Percent Concentration, (w/v)%

One of the most common methods for expressing percent concentration is to give the number of grams (weight) as a percentage of the number of milliliters (volume) of the final solution—called the **weight/volume percent concentration, (w/v)%**. Mathematically, (w/v)% concentration is found by taking the number of grams of solute per milliliter of solution and multiplying by 100%:

$$(\text{w/v})\% \text{ concentration} = \frac{\text{Mass of solute (g)}}{\text{Volume of solution (mL)}} \times 100\%$$

For example, if 15 g of glucose is dissolved in enough water to give 100 mL of solution, the glucose concentration is 15 g/100 mL or 15% (w/v):

$$\frac{15 \text{ g glucose}}{100 \text{ mL solution}} \times 100\% = 15\% (\text{w/v})$$

To prepare 100 mL of a specific weight/volume solution, the weighed solute is dissolved in just enough solvent to give a final volume of 100 mL, not in an initial volume of 100 mL solvent. (If the solute is dissolved in 100 mL of solvent, the final volume of the solution will likely be a bit larger than 100 mL, since the volume of the solute is included.) In practice, solutions are prepared using a volumetric flask, as shown previously in Figure 9.5. Worked Example 9.7 illustrates how weight/volume percent concentration is found from a known mass and volume of solution.

WORKED EXAMPLE 9.7 Solution Concentration: Weight/Volume Percent

A solution of heparin sodium, an anticoagulant for blood, contains 1.8 g of heparin sodium dissolved to make a final volume of 15 mL of solution. What is the weight/volume percent concentration of this solution?

ANALYSIS Weight/volume percent concentration is defined as the mass of the solute in grams divided by the volume of solution in milliliters and multiplied by 100%.

BALLPARK ESTIMATE The mass of solute (1.8 g) is smaller than the volume of solvent (15 mL) by a little less than a factor of 10. The weight/volume percent should thus be a little greater than 10%.

SOLUTION

$$(\text{w/v})\% \text{ concentration} = \frac{1.8 \text{ g heparin sodium}}{15 \text{ mL}} \times 100\% = 12\% (\text{w/v})$$

BALLPARK CHECK: The calculated (w/v)% is reasonably close to our original estimate of 10%.

WORKED EXAMPLE 9.8 Weight/Volume Percent as Conversion Factor: Volume to Mass

How many grams of NaCl are needed to prepare 250 mL of a 1.5% (w/v) saline solution?

ANALYSIS We are given a concentration and a volume, and we need to find the mass of solute by rearranging the equation for (w/v)% concentration.

BALLPARK ESTIMATE The desired (w/v)% value, 1.5%, is between 1 and 2%. For a volume of 250 mL, we would need 2.5 g of solute for a 1% (w/v) solution and 5.0 g of solute for a 2% solution. Thus, for our 1.5% solution, we need a mass midway between 2.5 and 5.0 g, or about 3.8 g.

SOLUTION

$$\text{Since } (w/v)\% = \frac{\text{Mass of solute in g}}{\text{Volume of solution in mL}} \times 100\%$$

$$\begin{aligned} \text{then Mass of solute in g} &= \frac{(\text{Volume of solution in mL})(w/v)\%}{100\%} \\ &= \frac{(250)(1.5\%)}{100\%} = 3.75 \text{ g} = 3.8 \text{ g NaCl} \\ &\quad (2 \text{ significant figures}) \end{aligned}$$

BALLPARK CHECK: The calculated answer matches our estimate.

WORKED EXAMPLE 9.9 Weight/Volume Percent as Conversion Factor: Mass to Volume

How many milliliters of a 0.75% (w/v) solution of the food preservative sodium benzoate are needed to obtain 45 mg?

ANALYSIS We are given a concentration and a mass, and we need to find the volume of solution by rearranging the equation for (w/v)% concentration. Remember that 45 mg = 0.045 g.

BALLPARK ESTIMATE A 0.75% (w/v) solution contains 0.75 g (750 mg) for every 100 mL of solution, so 10 mL contains 75 mg. To obtain 45 mg, we need a little more than half this volume, or a little more than 5 mL.

SOLUTION

$$\text{Since } (w/v)\% = \frac{\text{Mass of solute in g}}{\text{Volume of solution in mL}} \times 100\%$$

$$\begin{aligned} \text{then Volume of solution in mL} &= \frac{(\text{Mass of solute in g})(100\%)}{(w/v)\%} \\ &= \frac{(0.045 \text{ g})(100\%)}{0.75\%} = 6.0 \text{ mL} \end{aligned}$$

BALLPARK CHECK: The calculated answer is consistent with our estimate of a little more than 5 mL.

PROBLEM 9.12

In clinical lab reports, some concentrations are given in mg/dL. Convert a Ca^{2+} concentration of 8.6 mg/dL to weight/volume percent.

PROBLEM 9.13

What is the weight/volume percent concentration of a solution that contains 23 g of potassium iodide in 350 mL of aqueous solution?

PROBLEM 9.14

How many grams of solute are needed to prepare the following solutions?

- (a) 125.0 mL of 16% (w/v) glucose ($\text{C}_6\text{H}_{12}\text{O}_6$)
 (b) 65 mL of 1.8% (w/v) KCl

Volume/Volume Percent Concentration, (v/v)%

The concentration of a solution made by dissolving one liquid in another is often given by expressing the volume of solute as a percentage of the volume of final solution—the **volume/volume percent concentration, (v/v)%**. Mathematically, the volume of the solute (usually in milliliters) per milliliter of solution is multiplied by 100%:

$$(v/v)\% \text{ concentration} = \frac{\text{Volume of solute (mL)}}{\text{Volume of solution (mL)}} \times 100\%$$

For example, if 10.0 mL of ethyl alcohol is dissolved in enough water to give 100.0 mL of solution, the ethyl alcohol concentration is $(10.0 \text{ mL}/100.0 \text{ mL}) \times 100\% = 10.0\%$ (v/v).

WORKED EXAMPLE 9.10 Volume Percent: Volume of Solution to Volume of Solute

How many milliliters of methyl alcohol are needed to prepare 75 mL of a 5.0% (v/v) solution?

ANALYSIS We are given a solution volume (75 mL) and a concentration [5.0% (v/v), meaning 5.0 mL solute/100 mL solution]. The concentration acts as a conversion factor for finding the amount of methyl alcohol needed.

BALLPARK ESTIMATE A 5% (v/v) solution contains 5 mL of solute in 100 mL of solution, so the amount of solute in 75 mL of solution must be about three-fourths of 5 mL, which means between 3 and 4 mL.

SOLUTION

$$(75 \text{ mL solution}) \left(\frac{5.0 \text{ mL methyl alcohol}}{100 \text{ mL solution}} \right) = 3.8 \text{ mL methyl alcohol}$$

BALLPARK CHECK: The calculated answer is consistent with our estimate of between 3 and 4 mL.

PROBLEM 9.15

How would you use a 500.0 mL volumetric flask to prepare a 7.5% (v/v) solution of acetic acid in water?

PROBLEM 9.16

What volume of solute (in milliliters) is needed to prepare the following solutions?

- (a) 100 mL of 22% (v/v) ethyl alcohol (b) 150 mL of 12% (v/v) acetic acid

Parts per Million (ppm)

The concentration units weight/volume percent, (w/v)%, and volume/volume percent, (v/v)%, can also be defined as *parts per hundred* (pph) since 1% means one item per 100 items. When concentrations are very small, as often occurs in dealing with trace amounts of pollutants or contaminants, it is more convenient to use **parts per million (ppm)** or **parts per billion (ppb)**. The “parts” can be in any unit of either mass or volume as long as the units of both solute and solvent are the same:

$$\text{ppm} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 10^6 \quad \text{or} \quad \frac{\text{Volume of solute (mL)}}{\text{Volume of solution (mL)}} \times 10^6$$

$$\text{ppb} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 10^9 \quad \text{or} \quad \frac{\text{Volume of solute (mL)}}{\text{Volume of solution (mL)}} \times 10^9$$

To take an example, the maximum allowable concentration in air of the organic solvent benzene (C_6H_6) is currently set by government regulation at 1 ppm. A concentration of 1 ppm means that if you take a million “parts” of air in any unit—say, mL—then 1 of those parts is benzene vapor and the other 999,999 parts are other gases:

$$1 \text{ ppm} = \frac{1 \text{ mL}}{1,000,000 \text{ mL}} \times 10^6$$

Because the density of water is approximately 1.0 g/mL at room temperature, 1.0 L (or 1000 mL) of an aqueous solution weighs 1000 g. Therefore, when dealing with very dilute concentrations of solutes dissolved in water, ppm is equivalent to mg solute/L solution, and ppb is equivalent to μg solute/L solution. To demonstrate that these units are equivalent, the conversion from ppm to mg/L is as follows:

$$1 \text{ ppm} = \left(\frac{1 \text{ g solute}}{10^6 \text{ g solution}} \right) \left(\frac{1 \text{ mg solute}}{10^{-3} \text{ g solute}} \right) \left(\frac{10^3 \text{ g solution}}{1 \text{ L solution}} \right) = \frac{1 \text{ mg solute}}{1 \text{ L solution}}$$

WORKED EXAMPLE 9.11 ppm as Conversion Factor: Mass of Solution to Mass of Solute

The maximum allowable concentration of chloroform, CHCl_3 , in drinking water is 100 ppb. What is the maximum amount (in grams) of chloroform allowed in a glass containing 400 g (400 mL) of water?

ANALYSIS We are given a solution amount (400 g) and a concentration (100 ppb). This concentration of 100 ppb means

$$100 \text{ ppb} = \frac{\text{Mass of solute (g)}}{\text{Mass of solution (g)}} \times 10^9$$

This equation can be rearranged to find the mass of solute.

BALLPARK ESTIMATE A concentration of 100 ppb means there are 100×10^{-9} g (1×10^{-7} g) of solute in 1 g of solution. In 400 g of solution, we should have 400 times this amount, or $400 \times 10^{-7} = 4 \times 10^{-5}$ g.

SOLUTION

$$\begin{aligned} \text{Mass of solute (g)} &= \frac{\text{Mass of solution (g)}}{10^9} \times 100 \text{ ppb} \\ &= \frac{400 \text{ g}}{10^9} \times 100 \text{ ppb} = 4 \times 10^{-5} \text{ g (or 0.04 mg)} \end{aligned}$$

BALLPARK CHECK: The calculated answer matches our estimate.

PROBLEM 9.17

What is the concentration in ppm of sodium fluoride in tap water that has been fluoridated by the addition of 32 mg of NaF for every 20 kg of solution?

PROBLEM 9.18

The maximum amounts of lead and copper allowed in drinking water are 0.015 mg/kg for lead and 1.3 mg/kg for copper. Express these values in parts per million, and tell the maximum amount of each (in grams) allowed in 100 g of water.

9.8 Dilution

Many solutions, from orange juice to chemical reagents, are stored in high concentrations and then prepared for use by *dilution*—that is, by adding additional solvent to lower the concentration. For example, you might make up 1/2 gal of orange juice by adding water to a canned concentrate. In the same way, you might buy a medicine or chemical reagent in concentrated solution and dilute it before use.

The key fact to remember about dilution is that the amount of *solute* remains constant; only the *volume* is changed by adding more solvent. If, for example, the initial and final concentrations are given in molarity, then we know that the number of moles of solute is the same both before and after dilution, and can be determined by multiplying molarity times volume:

$$\text{Number of moles} = \text{Molarity (mol/L)} \times \text{Volume (L)}$$

Because the number of moles remains constant, we can set up the following equation, where M_1 and V_1 refer to the solution before dilution, and M_2 and V_2 refer to the solution after dilution:

$$\text{Moles of solute} = M_1V_1 = M_2V_2$$

This equation can be rewritten to solve for M_2 , the concentration of the solution after dilution:

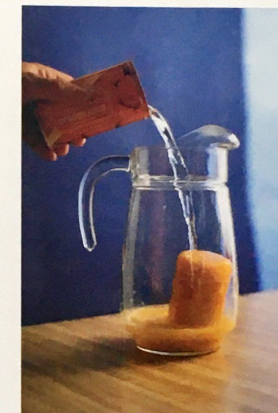
$$M_2 = M_1 \times \frac{V_1}{V_2} \quad \text{where} \quad \frac{V_1}{V_2} \quad \text{is a dilution factor}$$

The equation shows that the concentration after dilution (M_2) can be found by multiplying the initial concentration (M_1) by a **dilution factor**, which is simply the ratio of the initial and final solution volumes (V_1/V_2). If, for example, the solution volume increases by a factor of 5, from 10 mL to 50 mL, then the concentration must decrease to 1/5 its initial value because the dilution factor is 10 mL/50 mL, or 1/5. Worked Example 9.12 shows how to use this relationship for calculating dilutions.

The relationship between concentration and volume can also be used to find what volume of initial solution to start with to achieve a given dilution:

$$\begin{aligned} \text{Since} \quad M_1V_1 &= M_2V_2 \\ \text{then} \quad V_1 &= V_2 \times \frac{M_2}{M_1} \end{aligned}$$

In this case, V_1 is the initial volume that must be diluted to prepare a less concentrated solution with volume V_2 . The initial volume is found by multiplying the final volume (V_2) by the ratio of the final and initial concentrations (M_2/M_1). For example, to decrease the concentration of a solution to 1/5 its initial value, the initial volume must be 1/5 the desired final volume. Worked Example 9.13 gives a sample calculation.



▲ Orange juice concentrate is diluted with water before drinking.

Dilution factor The ratio of the initial and final solution volumes (V_1/V_2).

Although the preceding equations and following examples deal with concentration units of molarity, it is worth noting that the dilution equation can be generalized to the other concentration units presented in this section, or

$$C_1V_1 = C_2V_2$$

WORKED EXAMPLE 9.12 Dilution of Solutions: Concentration

What is the final concentration if 75 mL of a 3.5 M glucose solution is diluted to a volume of 450 mL?

ANALYSIS The number of moles of solute is constant, so

$$M_1V_1 = M_2V_2$$

Of the four variables in this equation, we know the initial concentration M_1 (3.5 M), the initial volume V_1 (75 mL), and the final volume V_2 (450 mL), and we need to find the final concentration M_2 .

BALLPARK ESTIMATE The volume increases by a factor of 6, from 75 mL to 450 mL, so the concentration must decrease by a factor of 6, from 3.5 M to about 0.6 M.

SOLUTION

Solving the above equation for M_2 and substituting in the known values gives

$$M_2 = \frac{M_1V_1}{V_2} = \frac{(3.5 \text{ M glucose})(75 \text{ mL})}{450 \text{ mL}} = 0.58 \text{ M glucose}$$

BALLPARK CHECK: The calculated answer is close to our estimate of 0.6 M.

WORKED EXAMPLE 9.13 Dilution of Solutions: Volume

Aqueous NaOH can be purchased at a concentration of 1.0 M. How would you use this concentrated solution to prepare 750 mL of 0.32 M NaOH?

ANALYSIS The number of moles of solute is constant, so

$$M_1V_1 = M_2V_2$$

Of the four variables in this equation, we know the initial concentration M_1 (1.0 M), the final volume V_2 (750 mL), and the final concentration M_2 (0.32 M), and we need to find the initial volume V_1 .

BALLPARK ESTIMATE We want the solution concentration to decrease by a factor of about 3, from 1.0 M to 0.32 M, which means we need to dilute the 1.0 M solution by a factor of 3. This means the final volume must be about 3 times greater than the initial volume. Because our final volume is to be 750 mL, we must start with an initial volume of about 250 mL.

SOLUTION

Solving the above equation for V_1 and substituting in the known values gives

$$V_1 = \frac{V_2M_2}{M_1} = \frac{(750 \text{ mL})(0.32 \text{ M})}{1.0 \text{ M}} = 240 \text{ mL}$$

To prepare the desired solution, dilute 240 mL of 1.0 M NaOH with water to make a final volume of 750 mL.

BALLPARK CHECK: The calculated answer (240 mL) is reasonably close to our estimate of 250 mL.

PROBLEM 9.19

Hydrochloric acid is normally purchased at a concentration of 12.0 M. What is the final concentration if 100.0 mL of 12.0 M HCl is diluted to 500.0 mL?

PROBLEM 9.20

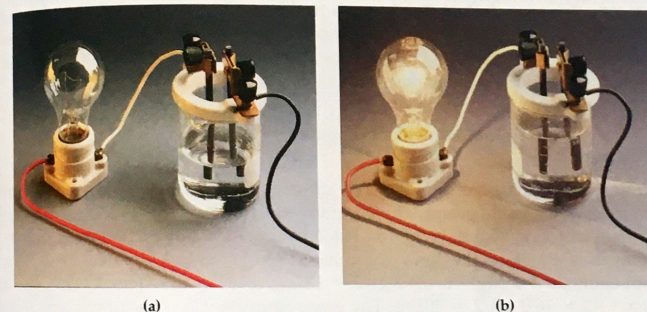
Aqueous ammonia is commercially available at a concentration of 16.0 M. How much of the concentrated solution would you use to prepare 500.0 mL of a 1.25 M solution?

PROBLEM 9.21

The Environmental Protection Agency has set the limit for arsenic in drinking water at 0.010 ppm. To what volume would you need to dilute 1.5 L of water containing 5.0 ppm arsenic to reach the acceptable limit?

9.9 Ions in Solution: Electrolytes

Look at Figure 9.8, which shows a light bulb connected to a power source through a circuit that is interrupted by two metal strips dipped into a beaker of liquid. When the strips are dipped into pure water, the bulb remains dark, but when they are dipped into an aqueous NaCl solution, the circuit is closed and the bulb lights. As mentioned previously in Section 4.1, this simple demonstration shows that ionic compounds in aqueous solution can conduct electricity. (p. 79)



▲ FIGURE 9.8 A simple demonstration shows that electricity can flow through a solution of ions. (a) With pure water in the beaker, the circuit is incomplete, no electricity flows, and the bulb does not light. (b) With a concentrated NaCl solution in the beaker, the circuit is complete, electricity flows, and the light bulb glows.

Substances like NaCl that conduct an electric current when dissolved in water are called **electrolytes**. Conduction occurs because negatively charged Cl^- anions migrate through the solution toward the metal strip connected to the positive terminal of the power source, whereas positively charged Na^+ cations migrate toward the strip connected to the negative terminal. As you might expect, the ability of a solution to conduct electricity depends on the concentration of ions in solution. Distilled water contains virtually no ions and is nonconducting; ordinary tap water contains low concentrations of dissolved ions (mostly Na^+ , K^+ , Mg^{2+} , Ca^{2+} , and Cl^-) and is weakly conducting; and a concentrated solution of NaCl is strongly conducting.

Electrolyte A substance that produces ions and therefore conducts electricity when dissolved in water.