

Solutions



The ocean is a salt solution covering the majority of the Earth's surface.

Chapter Outline

- **14.1** General Properties of Solutions
- 14.2 Solubility
- **14.3** Factors Related to Solubility
- **14.4** Rate of Dissolving Solids
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14.1 GENERAL PROPERTIES OF SOLUTIONS

M ost substances we encounter in our daily lives are mixtures. Often they are homogeneous mixtures, which are called *solutions*. Some solutions we commonly encounter are shampoo, soft drinks, or wine. Blood plasma is a complex mixture composed of compounds and ions dissolved in water and proteins suspended in the solution. These solutions all have water as a main component, but many common items, such as air, gasoline, and steel, are also solutions that do not contain water. What are the necessary components of a solution? Why do some substances dissolve, while others do not? What effect does a dissolved substance have on the properties of the solution? Answering these questions is the first step in understanding the solutions we encounter in our daily lives.

14.1 General Properties of Solutions

The term **solution** is used in chemistry to describe a system in which one or more substances are homogeneously mixed or dissolved in another substance. A simple solution has two components: a solute and a solvent. The **solute** is the component that is dissolved or is the least abundant component in the solution. The **solvent** is the dissolving agent or the most abundant component in the solution. For example, when salt is dissolved in water to form a solution, salt is the solute and water is the solvent. Complex solutions containing more than one solute and/or more than one solvent are common.

The three states of matter—solid, liquid, and gas—give us nine different types of solutions: solid dissolved in solid, solid dissolved in liquid, solid dissolved in gas, liquid dissolved in liquid, and so on. Of these, the most common solutions are solid dissolved in liquid, liquid dissolved in liquid, gas dissolved in liquid, and gas dissolved in gas. Some common types of solutions are listed in Table 14.1.

A true solution is one in which the particles of dissolved solute are molecular or ionic in size, generally in the range of 0.1 to 1 nm (10^{-8} to 10^{-7} cm). The properties of a true solution are as follows:

- 1. A mixture of two or more components—solute and solvent—is homogeneous and has a variable composition; that is, the ratio of solute to solvent can be varied.
- 2. The dissolved solute is molecular or ionic in size.
- 3. It is either colored or colorless and is usually transparent.
- **4.** The solute remains uniformly distributed throughout the solution and will not settle out with time.
- **5.** The solute can generally be separated from the solvent by purely physical means (for example, by evaporation).

Phase of solution	Solute	Solvent	Example
Gas	gas	gas	air
Liquid	gas	liquid	soft drinks
Liquid	liquid	liquid	antifreeze
Liquid	solid	liquid	salt water
Solid	gas	solid	H ₂ in Pt
Solid	solid	solid	brass

Table 14.1 Common Types of Solutions

solution

solute

solvent



CHEMISTRY IN ACTION • Exploding Silicon

We usually consider silicon, a main ingredient in sand and computer chips, to be reasonably inert. So imagine researcher Frederic Mikulec's surprise

when he cut a silicon wafer with a diamond scribe and it blew up in his face. Michael Sailor, professor of chemistry and biochemistry at the University of California, San Diego, recalls, "It was just a small explosion, like a cap going off in a cap gun." The event inspired the research team to investigate the process further.

The silicon wafers being used were specially processed to have many tiny holes. Gadolinium nitrate was dissolved in ethanol and added to the porous silicon wafers. After the ethanol evaporated, the gadolinium nitrate remained inside the pores in the silicon chip. Sailor's team discovered that they could explode these chips using a 9-volt battery or by scratching them.

Next the team began to look for ways to use these exploding silicon chips. The scientists added metal ions



to the wafers and exploded them. Their results indicated that the elements could be identified by their emission spectra in the flame from the explosion. Currently, groundwater samples must be taken to the lab and analyzed in an emission spectrometer by being mixed with other chemicals and burned. Sailor thinks that these wafers could be used as a sensor for analyzing water contaminants in the field. Imagine being able to test water for heavy metals using an instrument you could hold in the palm of your hand.

Since the silicon wafers explode like gunpowder, these tiny wafers could also be used to operate tiny micromotors or microvehicles. Sailor also thinks that these silicon wafers could replace blasting caps in explosives.



Note the beautiful purple trails of $KMnO_4$ as the crystals dissolve.

solubility

Let's illustrate these properties using water solutions of sugar and of potassium permanganate. We prepare two sugar solutions, the first containing 10 g of sugar added to 100 mL of water and the second containing 20 g of sugar added to 100 mL of water. Each solution is stirred until all the solute dissolves, demonstrating that we can vary the composition of a solution. Every portion of the solution has the same sweet taste because the sugar molecules are uniformly distributed throughout. If confined so that no solvent is lost, the solution will taste and appear the same a week or a month later. A solution cannot be separated into its components by filtering it. But by carefully evaporating the water, we can recover the sugar from the solution.

To observe the dissolving of potassium permanganate $(KMnO_4)$, we drop a few crystals of it in a beaker of water. Almost at once, the beautiful purple color of dissolved permanganate ions (MnO_4) appears and streams to the bottom of the beaker as the crystals dissolve. After a while, the purple color disperses until it's evenly distributed throughout the solution. This dispersal demonstrates that molecules and ions move about freely and spontaneously (diffuse) in a liquid or solution.

Solution permanency is explained in terms of the kinetic-molecular theory (see Section 12.2). According to the KMT, both the solute and solvent particles (molecules or ions) are in constant random motion. This motion is energetic enough to prevent the solute particles from settling out under the influence of gravity.

14.2 Solubility

The term **solubility** describes the amount of one substance (solute) that will dissolve in a specified amount of another substance (solvent) under stated conditions. For example, 36.0 g of sodium chloride will dissolve in 100 g of water at 20°C. We say then that the solubility of NaCl in water is 36.0 g/100 g H₂O at 20°C. Solubility is often used in a relative way. For instance, we say that a

14.2 SOLUBILITY

substance is very soluble, moderately soluble, slightly soluble, or insoluble. Although these terms do not accurately indicate how much solute will dissolve, they are frequently used to describe the solubility of a substance qualitatively.

Two other terms often used to describe solubility are *miscible* and *immiscible*. Liquids that are capable of mixing and forming a homogeneous solution are miscible; those that do not form solutions or are generally insoluble in each other are **immiscible**. Methyl alcohol and water are miscible in each other in all proportions. Oil and water are immiscible, forming two separate layers when they are mixed, as shown in Figure 14.1.

The general guidelines for the solubility of common ionic compounds (salts) are given in Figure 14.2. These guidelines have some exceptions, but they provide a solid foundation for the compounds considered in this course. The solubilities of over 200 compounds are given in the Solubility Table in Appendix V. Solubility data for thousands of compounds can be found by consulting standard reference sources.*

The quantitative expression of the amount of dissolved solute in a particular quantity of solvent is known as the **concentration of a solution**. Several methods of expressing concentration are described in Section 14.6.

miscible immiscible

The term salt is used interchangeably with ionic compound by many chemists.

concentration of a solution



Figure 14.1 An immiscible mixture of oil and water.

Figure 14.2

The solubility of various common ions. Substances containing the ions on the left are generally soluble in cold water, while those substances containing the ions on the right are insoluble in cold water. The arrows point to the exceptions.

*Two commonly used handbooks are Lange's Handbook of Chemistry, 15th ed. (New York: McGrawHill, 1998), and the Handbook of Chemistry and Physics, 85th ed. (Cleveland: Chemical Rubber Co., 2004).

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14.3 Factors Related to Solubility

Predicting solubilities is complex and difficult. Many variables, such as size of ions, charge on ions, interaction between ions, interaction between solute and solvent, and temperature, complicate the problem. Because of the factors involved, the general rules of solubility given in Figure 14.2 have many exceptions. However, these rules are useful because they do apply to many of the more common compounds that we encounter in the study of chemistry. Keep in mind that these are rules, not laws, and are therefore subject to exceptions. Fortunately, the solubility of a solute is relatively easy to determine experimentally. Now let's examine the factors related to solubility.

The Nature of the Solute and Solvent

The old adage "like dissolves like" has merit, in a general way. Polar or ionic substances tend to be more compatible with other polar substances. Nonpolar substances tend to be compatible with other nonpolar substances and less miscible with polar substances. Thus ionic compounds, which are polar, tend to be much more soluble in water, which is polar, than in solvents such as ether, hexane, or benzene, which are essentially nonpolar. Sodium chloride, an ionic substance, is soluble in water, slightly soluble in ethyl alcohol (less polar than water), and insoluble in ether and benzene. Pentane, C_5H_{12} , a nonpolar substance, is only slightly soluble in water but is very soluble in benzene and ether.

At the molecular level the formation of a solution from two nonpolar substances, such as hexane and benzene, can be visualized as a process of simple mixing. These nonpolar molecules, having little tendency to either attract or repel one another, easily intermingle to form a homogeneous solution.

Solution formation between polar substances is much more complex. See, for example, the process by which sodium chloride dissolves in water (Figure 14.3). Water molecules are very polar and are attracted to other polar molecules or ions. When salt crystals are put into water, polar water molecules become attracted to the sodium and chloride ions on the crystal surfaces and weaken the attraction between Na⁺ and Cl⁻ ions. The positive end of the water dipole is attracted to the Cl⁻ ions, and the negative end of the water dipole to the Na⁺ ions. The weakened attraction permits the ions to move apart, making room for more water dipoles. Thus the surface ions are surrounded by water molecules, becoming hydrated ions, Na⁺(aq) and Cl⁻(aq), and slowly diffuse away from the crystals and dissolve in solution:

$$\operatorname{NaCl}(\operatorname{crystal}) \xrightarrow{H_2O} \operatorname{Na}^+(aq) + \operatorname{Cl}^-(aq)$$

Examination of the data in Table 14.2 reveals some of the complex questions relating to solubility.

The Effect of Temperature on Solubility

Temperature affects the solubility of most substances, as shown by the data in Table 14.2. Most solutes have a limited solubility in a specific solvent at a fixed temperature. For most solids dissolved in a liquid, an increase in temperature



Figure 14.3

Dissolution of sodium chloride in water. Polar water molecules are attracted to Na^+ and Cl^- ions in the salt crystal, weakening the attraction between the ions. As the attraction between the ions weakens, the ions move apart and become surrounded by water dipoles. The hydrated ions slowly diffuse away from the crystal to become dissolved in solution.

14.3 FACTORS RELATED TO SOLUBILITY

		Solubility (g salt/100 g H ₂ O)		
Salt	0°C	100°C		
LiF	0.12	0.14 (at 35°C)		
LiCl	67	127.5		
LiBr	143	266		
LiI	151	481		
NaF	4	5		
NaCl	35.7	39.8		
NaBr	79.5	121		
NaI	158.7	302		
KF	92.3 (at 18°C)	Very soluble		
KCl	27.6	57.6		
KBr	53.5	104		
KI	127.5	208		

Table 14.2 Solubility of Alkali Metal Halides in Water

results in increased solubility (see Figure 14.4). However, no single rule governs the solubility of solids in liquids with change in temperature. Some solids increase in solubility only slightly with increasing temperature (see NaCl in Figure 14.4); other solids decrease in solubility with increasing temperature (see Li_2SO_4 in Figure 14.4).

On the other hand, the solubility of a gas in water usually decreases with increasing temperature (see HCl and SO_2 in Figure 14.4). The tiny bubbles that form when water is heated are due to the decreased solubility of air at higher temperatures. The decreased solubility of gases at higher temperatures is explained in terms of the KMT by assuming that, in order to dissolve, the gas





Pouring root beer into a glass illustrates the effect of pressure on solubility. The escaping CO₂ produces the foam.

saturated solution

unsaturated solution

molecules must form bonds of some sort with the molecules of the liquid. An increase in temperature decreases the solubility of the gas because it increases the kinetic energy (speed) of the gas molecules and thereby decreases their ability to form "bonds" with the liquid molecules.

The Effect of Pressure on Solubility

Small changes in pressure have little effect on the solubility of solids in liquids or liquids in liquids but have a marked effect on the solubility of gases in liquids. The solubility of a gas in a liquid is directly proportional to the pressure of that gas above the solution. Thus the amount of a gas dissolved in solution will double if the pressure of that gas over the solution is doubled. For example, carbonated beverages contain dissolved carbon dioxide under pressures greater than atmospheric pressure. When a can of carbonated soda is opened, the pressure is immediately reduced to the atmospheric pressure, and the excess dissolved carbon dioxide bubbles out of the solution.

Saturated, Unsaturated, and Supersaturated Solutions

At a specific temperature there is a limit to the amount of solute that will dissolve in a given amount of solvent. When this limit is reached, the resulting solution is said to be saturated. For example, when we put 40.0 g of KCl into 100 g of H_2O at 20°C, we find that 34.0 g of KCl dissolve and 6.0 g of KCl remain undissolved. The solution formed is a saturated solution of KCl.

Two processes are occurring simultaneously in a saturated solution. The solid is dissolving into solution and, at the same time, the dissolved solute is crystallizing out of solution. This may be expressed as

solute (undissolved) \implies solute (dissolved)

When these two opposing processes are occurring at the same rate, the amount of solute in solution is constant, and a condition of equilibrium is established between dissolved and undissolved solute. Therefore, a saturated solution contains dissolved solute in equilibrium with undissolved solute. Thus, any point on any solubility curve (Figure 14.4) represents a saturated solution of that solute. For example, a solution containing 60 g NH₄Cl per 100 g H₂O is saturated at 70°C.

It's important to state the temperature of a saturated solution, because a solution that is saturated at one temperature may not be saturated at another. If the temperature of a saturated solution is changed, the equilibrium is disturbed, and the amount of dissolved solute will change to reestablish equilibrium.

A saturated solution may be either dilute or concentrated, depending on the solubility of the solute. A saturated solution can be conveniently prepared by dissolving a little more than the saturated amount of solute at a temperature somewhat higher than room temperature. Then the amount of solute in solution will be in excess of its solubility at room temperature, and, when the solution cools, the excess solute will crystallize, leaving the solution saturated. (In this case, the solute must be more soluble at higher temperatures and must not form a supersaturated solution.) Examples expressing the solubility of saturated solutions at two different temperatures are given in Table 14.3.

An **unsaturated solution** contains less solute per unit of volume than does its corresponding saturated solution. In other words, additional solute can be dissolved in an unsaturated solution without altering any other conditions.



Solute

NaCl

KCl

NaNO₃

KClO₃

AgNO₃

C12H22O11

14.3 FACTORS RELATED TO SOLUBILITY

Table 14.3 Saturated Solutions at 20°C and 50°C

Consider a solution made by adding 40 g of KCI to 100 g of H_2O at 20°C (see Table 14.3). The solution formed will be saturated and will contain about 6 g of undissolved salt, because the maximum amount of KCI that can dissolve in 100 g of H_2O at 20°C is 34 g. If the solution is now heated and maintained at 50°C, all the salt will dissolve and even more can be dissolved. The solution at 50°C is unsaturated.

20°C

36.0

34.0

88.0

222.0

203.9

7.4

Solubility (g solute/100 g H_2O)

50°C

37.0

42.6

114.0

19.3

455.0

260.4

In some circumstances, solutions can be prepared that contain more solute than needed for a saturated solution at a particular temperature. These solutions are said to be **supersaturated**. However, we must qualify this definition by noting that a supersaturated solution is unstable. Disturbances, such as jarring, stirring, scratching the walls of the container, or dropping in a "seed" crystal, cause the supersaturation to return to saturation, releasing heat. When a supersaturated solution is disturbed, the excess solute crystallizes out rapidly, returning the solution to a saturated state.

Supersaturated solutions are not easy to prepare but may be made from certain substances by dissolving, in warm solvent, an amount of solute greater than that needed for a saturated solution at room temperature. The warm solution is then allowed to cool very slowly. With the proper solute and careful work, a supersaturated solution will result.

Will a solution made by adding 2.5 g of $CuSO_4$ to 10 g of H_2O be saturated or **Example 14.1** unsaturated at 20°C?

We first need to know the solubility of $CuSO_4$ at 20°C. From Figure 14.4, we see that the solubility of $CuSO_4$ at 20°C is about 21 g per 100 g of H₂O. This amount is equivalent to 2.1 g of $CuSO_4$ per 10 g of H₂O.

Since 2.5 g per 10 g of H_2O is greater than 2.1 g per 10 g of H_2O , the solution will be saturated and 0.4 g of CuSO₄ will be undissolved.

Practice 14.1

Will a solution made by adding 9.0 g NH_4Cl to 20 g of H_2O be saturated or unsaturated at 50°C?

The heat released in this hot pack results from the crystallization of a supersaturated solution of sodium acetate.

supersaturated solution

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14.4 Rate of Dissolving Solids

The rate at which a solid dissolves is governed by (1) the size of the solute particles, (2) the temperature, (3) the concentration of the solution, and (4) agitation or stirring. Let's look at each of these conditions:

- 1. *Particle size*. A solid can dissolve only at the surface that is in contact with the solvent. Because the surface-to-volume ratio increases as size decreases, smaller crystals dissolve faster than large ones. For example, if a salt crystal 1 cm on a side (a surface area of 6 cm²) is divided into 1000 cubes, each 0.1 cm on a side, the total surface of the smaller cubes is 60 cm²—a 10-fold increase in surface area (see Figure 14.5).
- **2.** *Temperature.* In most cases, the rate of dissolving of a solid increases with temperature. This increase is due to kinetic effects. The solvent molecules move more rapidly at higher temperatures and strike the solid surfaces more often, causing the rate of dissolving to increase.
- **3.** *Concentration of the solution.* When the solute and solvent are first mixed, the rate of dissolving is at its maximum. As the concentration of the solution increases and the solution becomes more nearly saturated with the solute, the rate of dissolving decreases greatly. The rate of dissolving is graphed in Figure 14.6. Note that about 17 g dissolve in the first five-minute interval, but only about 1 g dissolves in the fourth five-minute interval. Although different solutes show different rates, the rate of dissolving always becomes very slow as the concentration approaches the saturation point.
- **4.** Agitation or stirring. The effect of agitation or stirring is kinetic. When a solid is first put into water, it comes in contact only with solvent in its immediate vicinity. As the solid dissolves, the amount of dissolved solute around the solid becomes more and more concentrated, and the rate of dissolving slows down. If the mixture is not stirred, the dissolved solute





Surface area of crystals. A crystal 1 cm on a side has a surface area of 6 cm². Subdivided into 1000 smaller crystals, each 0.1 cm on a side, the total surface area is increased to 60 cm^2 .

14.5 SOLUTIONS: A REACTION MEDIUM



Figure 14.6 Rate of dissolution of a solid solute in a solvent. The rate is maximum at the beginning and decreases as the concentration approaches saturation.

diffuses very slowly through the solution; weeks may pass before the solid is entirely dissolved. Stirring distributes the dissolved solute rapidly through the solution, and more solvent is brought into contact with the solid, causing it to dissolve more rapidly.

14.5 Solutions: A Reaction Medium

Many solids must be put into solution to undergo appreciable chemical reaction. We can write the equation for the double-displacement reaction between sodium chloride and silver nitrate:

 $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

But suppose we mix solid NaCl and solid AgNO₃ and look for a chemical change. If any reaction occurs, it is slow and virtually undetectable. In fact, the crystalline structures of NaCl and AgNO₃ are so different that we could separate them by tediously picking out each kind of crystal from the mixture. But if we dissolve the NaCl and AgNO₃ separately in water and mix the two solutions, we observe the immediate formation of a white, curdlike precipitate of silver chloride.

Molecules or ions must collide with one another in order to react. In the foregoing example, the two solids did not react because the ions were securely locked within their crystal structures. But when the NaCl and AgNO₃ are dissolved, their crystal lattices are broken down and the ions become mobile. When the two solutions are mixed, the mobile Ag^+ and Cl^- ions come into contact and react to form insoluble AgCl, which precipitates out of solution. The soluble Na⁺ and NO₃⁻ ions remain mobile in solution but form the crystalline salt NaNO₃ when the water is evaporated:

 $\begin{aligned} &\operatorname{NaCl}(aq) + \operatorname{AgNO}_3(aq) \longrightarrow \operatorname{AgCl}(s) + \operatorname{NaNO}_3(aq) \\ &\operatorname{Na}^+(aq) + \operatorname{Cl}^-(aq) + \operatorname{Ag}^+(aq) + \operatorname{NO}_3^-(aq) \longrightarrow \operatorname{AgCl}(s) + \operatorname{Na}^+(aq) + \operatorname{NO}_3^-(aq) \\ &\operatorname{solution} & \operatorname{silver nitrate} & \operatorname{silver} & \operatorname{solution nitrate} \\ &\operatorname{solution} & \operatorname{solution} & \operatorname{chloride} & \operatorname{in solution} \end{aligned}$

The mixture of the two solutions provides a medium or space in which the Ag^+ and CI^- ions can react. (See Chapter 15 for further discussion of ionic reactions.)

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Solutions also function as diluting agents in reactions in which the undiluted reactants would combine with each other too violently. Moreover, a solution of known concentration provides a convenient method for delivering a specific amount of reactant.

14.6 Concentration of Solutions

The concentration of a solution expresses the amount of solute dissolved in a given quantity of solvent or solution. Because reactions are often conducted in solution, it's important to understand the methods of expressing concentration and to know how to prepare solutions of particular concentrations. The concentration of a solution may be expressed qualitatively or quantitatively. Let's begin with a look at the qualitative methods of expressing concentration.

Dilute and Concentrated Solutions

When we say that a solution is *dilute* or *concentrated*, we are expressing, in a relative way, the amount of solute present. One gram of a compound and 2 g of a compound in solution are both dilute solutions when compared with the same volume of a solution containing 20 g of a compound. Ordinary concentrated hydrochloric acid contains 12 mol of HCl per liter of solution. In some laboratories, the dilute acid is made by mixing equal volumes of water and the concentrated acid. In other laboratories the concentrated acid is diluted with two or three volumes of water, depending on its use. The term **dilute solution**, then, describes a solution that contains a relatively small amount of dissolved solute.

Mass Percent Solution

The mass percent method expresses the concentration of the solution as the percent of solute in a given mass of solution. It says that for a given mass of solution, a certain percent of that mass is solute. Suppose we take a bottle from the reagent shelf that reads "sodium hydroxide, NaOH, 10%." This statement means that for every 100 g of this solution, 10 g will be NaOH and 90 g will be water. (Note that this amount of solution is 100 g, not 100 mL.) We could also make this same concentration of solution by dissolving 2.0 g of NaOH in 18 g of water. Mass percent concentrations are most generally used for solids dissolved in liquids:

mass percent =
$$\frac{\text{g solute}}{\text{g solute} + \text{g solvent}} \times 100 = \frac{\text{g solute}}{\text{g solution}} \times 100$$

As instrumentation advances are made in chemistry, our ability to measure the concentration of dilute solutions is increasing as well. In addition to mass percent, chemists now commonly use **parts per million (ppm)**:

parts per million = $\frac{\text{g solute}}{\text{g solute} + \text{g solvent}} \times 1,000,000$

Currently, air and water contaminants, drugs in the human body, and pesticide residues are measured in parts per million.

dilute solution

concentrated solution

Note that mass percent is independent of the formula for the solute.

parts per million (ppm)

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The percent expresses the mass of the solute:

 $\begin{array}{l} 250.\,g\,=\,total\,mass\,of\,solution\\ 5.00\%\,of\,250.\,g\,=\,(0.0500)(250.\,g)\,=\,12.5\,g\,KCl\,(solute)\\ 250.\,g\,-\,12.5\,g\,=\,238\,g\,H_2O \end{array}$

Dissolving 12.5 g KCl in 238 g H_2O gives a 5.00% KCl solution.

A 34.0% sulfuric acid solution has a density of 1.25 g/mL. How many grams of **Example 14.4** H_2SO_4 are contained in 1.00 L of this solution?

Since H_2SO_4 is the solute, we first solve the mass percent equation for grams **SOLUTION** of solute:

mass percent =
$$\left(\frac{\text{g solute}}{\text{g solution}}\right)100$$

$$g \text{ solute } = \frac{(\text{mass percent})(g \text{ solution})}{100}$$

The mass percent is given, so we need to determine the grams of solution. The mass of the solution can be calculated from the density data. Convert density (g/mL) to grams:

$$\frac{1.00 \text{ L} = 1.00 \times 10^3 \text{ mL}}{\left(\frac{1.25 \text{ g}}{\text{mL}}\right)(1.00 \times 10^3 \text{ mL})} = 1250 \text{ g} \text{ (mass of solution)}$$

Now we have all the figures to calculate the grams of solute:

g solute =
$$\frac{(34.0 \text{ g H}_2\text{SO}_4)(1250 \text{ g})}{100 \text{ g}} = 425 \text{ g H}_2\text{SO}_4$$

Thus, 1.00 L of 34.0% H_2SO_4 solution contains 425 g H_2SO_4 .

Practice 14.2

What is the mass percent of Na_2SO_4 in a solution made by dissolving 25.0 g Na_2SO_4 in 225.0 g H_2O ?

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Mass/Volume Percent (m/v)

This method expresses concentration as grams of solute per 100 mL of solution. With this system, a 10.0%-m/v-glucose solution is made by dissolving 10.0 g of glucose in water, diluting to 100 mL, and mixing. The 10.0%-m/v solution could also be made by diluting 20.0 g to 200 mL, 50.0 g to 500 mL, and so on. Of course, any other appropriate dilution ratio may be used:

mass/volume percent = $\frac{\text{g solute}}{\text{mL solution}} \times 100$



Example 14.5 A 3.0% H₂O₂ solution is commonly used as a topical antiseptic to prevent infection. What volume of this solution will contain 10. g of H₂O₂?

SOLUTION First solve the mass/volume percent equation for mL of solution:

mL solution = $\frac{(g \text{ solute})}{(m/v \text{ percent})}(100)$ mL solution = $\frac{(10. g \text{ solute})}{(3.0 m/v \text{ percent})}(100) = 330 \text{ mL}$

Volume Percent

Solutions that are formulated from two liquids are often expressed as *volume percent* with respect to the solute. The volume percent is the volume of a liquid in 100 mL of solution. The label on a bottle of ordinary rubbing alcohol reads "isopropyl alcohol, 70% by volume." Such a solution could be made by mixing 70 mL of alcohol with water to make a total volume of 100 mL, but we cannot use 30 mL of water, because the two volumes are not necessarily additive:

volume percent = $\frac{\text{volume of liquid in question}}{\text{total volume of solution}} \times 100$

Volume percent is used to express the concentration of alcohol in beverages. Wines generally contain 12% alcohol by volume. This translates into 12 mL of alcohol in each 100 mL of wine. The beverage industry also uses the concentration unit of *proof* (twice the volume percent). Pure ethyl alcohol is 100% and therefore 200 proof. Scotch whiskey is 86 proof, or 43% alcohol.

Molarity

Mass percent solutions do not equate or express the molar masses of the solute in solution. For example, 1000. g of 10.0% NaOH solution contains 100. g NaOH; 1000. g of 10.0% KOH solution contains 100. g KOH. In terms of moles of NaOH and KOH, these solutions contain

$$mol NaOH = (100. g NaOH) \left(\frac{1 mol NaOH}{40.00 g NaOH}\right) = 2.50 mol NaOH$$
$$mol KOH = (100. g KOH) \left(\frac{1 mol KOH}{56.11 g KOH}\right) = 1.78 mol KOH$$

From these figures, we see that the two 10.0% solutions do not contain the same number of moles of NaOH and KOH. As a result, we find that a 10.0% NaOH solution contains more reactive base than a 10.0% KOH solution.



14.6 CONCENTRATION OF SOLUTIONS

We need a method of expressing concentration that will easily indicate how many moles of solute are present per unit volume of solution. For this purpose, the concentration known as molarity is used in calculations involving chemical reactions.

A 1-molar solution contains 1 mol of solute per liter of solution. For example, to make a 1-molar solution of sodium hydroxide (NaOH) we dissolve 40.00 g NaOH (1 mol) in water and dilute the solution with more water to a volume of 1 L. The solution contains 1 mol of the solute in 1 L of solution and is said to be 1 molar in concentration. Figure 14.7 illustrates the preparation of a 1-molar solution. Note that the volume of the solute and the solvent together is 1 L.

The concentration of a solution can, of course, be varied by using more or less solute or solvent; but in any case the **molarity** of a solution is the number of moles of solute per liter of solution. The abbreviation for molarity is *M*. The units of molarity are moles per liter. The expression "2.0 *M* NaOH" means a 2.0-molar solution of NaOH (2.0 mol, or 80. g, of NaOH dissolved in water to make 1.0 L of solution):

molarity – M –	number of moles of solute _	moles
monanty - m -	liter of solution	liter

Flasks that are calibrated to contain specific volumes at a particular temperature are used to prepare solutions of a desired concentration. These *volumetric flasks* have a calibration mark on the neck that accurately indicates the measured volume. Molarity is based on a specific volume of solution and therefore will vary slightly with temperature because volume varies with temperature (1000 mL H₂O at 20°C = 1001 mL at 25°C).



Figure 14.7 Preparation of a 1 *M* solution.

molarity (M)



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Suppose we want to make 500 mL of 1 M solution. This solution can be prepared by determining the mass of 0.5 mol of the solute and diluting with water in a 500-mL (0.5-L) volumetric flask. The molarity will be

$$M = \frac{0.5 \text{ mol solute}}{0.5 \text{ L solution}} = 1 \text{ molar}$$

You can see that it isn't necessary to have a liter of solution to express molarity. All we need to know is the number of moles of dissolved solute and the volume of solution. Thus, 0.001 mol NaOH in 10 mL of solution is 0.1 M:

$$\left(\frac{0.001 \text{ mol}}{10 \text{ mL}}\right) \left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 0.1 M$$

When we stop to think that a balance is not calibrated in moles, but in grams, we can incorporate grams into the molarity formula. We do so by using the relationship

malas -	_ grams of solute
moles -	molar mass

Molarities of concentrated acids commonly used in the laboratory: HCl 12 *M* HC₂H₃O₂ 17 *M* HNO₃ 16 *M* H₂SO₄ 18 *M*

$$M = \frac{\text{mol}}{\text{L}} = \frac{\text{g solute}}{\text{molar mass solute} \times \text{L solution}}$$

We can now determine the mass of any amount of a solute that has a known formula, dilute it to any volume, and calculate the molarity of the solution using this formula.

Example 14.6	What is the molarity of a solution containing 1.4 mol of acetic acid $(HC_2H_3O_2)$ in 250. mL of solution?
SOLUTION	By the unit-conversion method, we note that the concentration given in the problem statement is 1.4 mol per 250. mL (mol/mL). Since molarity = mol/L, the needed conversion is
	$\frac{\text{mol}}{\text{mL}} \longrightarrow \frac{\text{mol}}{\text{L}} = M$
	$\left(\frac{1.4 \text{ mol}}{250. \text{ m}L}\right) \left(\frac{1000 \text{ m}L}{L}\right) = \frac{5.6 \text{ mol}}{L} = 5.6 M$

Example 14.7 What is the molarity of a solution made by dissolving 2.00 g of potassium chlorate in enough water to make 150. mL of solution?

SOLUTION We use the unit-conversion method. The steps in the conversions must lead to units of moles/liter:

$$\frac{g \operatorname{KClO}_3}{\mathrm{mL}} \longrightarrow \frac{g \operatorname{KClO}_3}{\mathrm{L}} \longrightarrow \frac{\mathrm{mol} \operatorname{KClO}_3}{\mathrm{L}} = M$$

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14.6 CONCENTRATION OF SOLUTIONS

The data are

mass $KCIO_3 = 2.00 \text{ g}$ molar mass $KCIO_3 = 122.6 \text{ g/mol}$ volume = 150. mL

$$\left(\frac{2.00 \text{ gKClO}_{3}}{150. \text{ mL}}\right) \left(\frac{1000 \text{ mL}}{L}\right) \left(\frac{1 \text{ mol KClO}_{3}}{122.6 \text{ gKClO}_{3}}\right) = \frac{0.109 \text{ mol}}{L} = 0.109 M \text{ KClO}_{3}$$

How many grams of potassium hydroxide are required to prepare 600. mL of **Example 14.8** 0.450 *M* KOH solution?

The conversion is

milliliters \longrightarrow liters \longrightarrow moles \longrightarrow grams

The data are

volume = 600. mL $M = \frac{0.450 \text{ mol}}{\text{L}}$ molar mass KOH $= \frac{56.11 \text{ g}}{\text{mol}}$

The calculation is

 $(600. \text{ mL})\left(\frac{1 \text{L}}{1000 \text{ mL}}\right)\left(\frac{0.450 \text{ mot}}{\text{L}}\right)\left(\frac{56.11 \text{ g KOH}}{\text{mot}}\right) = 15.1 \text{ g KOH}$

Practice 14.3

What is the molarity of a solution made by dissolving 7.50 g of magnesium nitrate $[Mg(NO_3)_2]$ in enough water to make 25.0 mL of solution?

Practice 14.4

How many grams of sodium chloride are needed to prepare 125 mL of a 0.037 *M* NaCl solution?

Calculate the number of moles of nitric acid in 325 mL of 16 *M* HNO₃ solution.

Use the equation

moles = liters $\times M$

Substitute the data given in the problem and solve:

moles =
$$(0.325 \mathcal{L}) \left(\frac{16 \text{ mol HNO}_3}{1 \mathcal{L}} \right) = 5.2 \text{ mol HNO}_3$$

What volume of 0.250 *M* solution can be prepared from 16.0 g of potassium **Example 14.10** carbonate?

We start with 16.0 g K_2CO_3 ; we need to find the volume of 0.250 *M* solution **SOLUTION** that can be prepared from this amount of K_2CO_3 . The conversion therefore is

 $g K_2 CO_3 \longrightarrow mol K_2 CO_3 \longrightarrow L$ solution



Example 14.9 SOLUTION

SOLUTION



The data are

$$\max K_{2}CO_{3} = 16.0 \text{ g} \quad M = \frac{0.250 \text{ mol}}{1 \text{ L}} \quad \text{molar mass } K_{2}CO_{3} = \frac{138.2 \text{ g}}{1 \text{ mol}}$$
$$(16.0 \text{ g} \text{K}_{2}CO_{3}) \left(\frac{1 \text{ mot } \text{K}_{2}CO_{3}}{138.2 \text{ g} \text{K}_{2}CO_{3}}\right) \left(\frac{1 \text{ L}}{0.250 \text{ mot } \text{K}_{2}CO_{3}}\right) = 0.463 \text{ L} (463 \text{ mL})$$

Thus, 463 mL of 0.250 M solution can be made from 16.0 g K_2CO_3 .



Practice 14.5

What volume of 0.035 M AgNO₃ can be made from 5.0 g of AgNO₃?

Practice 14.6

How many milliliters of 0.50 *M* NaOH are required to react completely with 25.00 mL of 1.5 *M* HCl?

We've now examined several ways to measure concentration of solutions quantitatively. A summary of these concentration units is found in Table 14.4.

14.6 CONCENTRATION OF SOLUTIONS

Table 14.4 Concentration Units for Solutions

Units	Symbol	Definition
Mass percent	% m/m	$\frac{\text{mass solute}}{\text{mass solution}} \times 100$
Parts per million	ppm	$\frac{\text{mass solute}}{\text{mass solution}} \times 1,000,000$
Mass/volume percent	% m/v	$\frac{\text{mass solute}}{\text{mL solution}} \times 100$
Volume percent	% v/v	$\frac{\text{mL solute}}{\text{mL solution}} \times 100$
Molarity	М	moles solute L solution
Molality	т	moles solute kg solvent

Molality is covered in Section 14.7

Dilution Problems

Chemists often find it necessary to dilute solutions from one concentration to another by adding more solvent to the solution. If a solution is diluted by adding pure solvent, the volume of the solution increases, but the number of moles of solute in the solution remains the same. Thus the moles/liter (molarity) of the solution decreases. Always read a problem carefully to distinguish between (1) how much solvent must be added to dilute a solution to a particular concentration and (2) to what volume a solution must be diluted to prepare a solution of a particular concentration.

Calculate the molarity of a sodium hydroxide solution that is prepared by mixing 100. mL of 0.20 *M* NaOH with 150. mL of water. Assume that the volumes are additive.

This problem is a dilution problem. If we double the volume of a solution by adding water, we cut the concentration in half. Therefore, the concentration of the above solution should be less than 0.10 M. In the dilution, the moles of NaOH remain constant; the molarity and volume change. The final volume is (100. mL + 150. mL), or 250. mL.

To solve this problem, (1) calculate the moles of NaOH in the original solution and (2) divide the moles of NaOH by the final volume of the solution to obtain the new molarity:

Step 1 Calculate the moles of NaOH in the original solution:

$$M = \frac{\text{mol}}{\text{L}} \quad \text{mol} = \text{L} \times M$$
$$(0.100 \, \text{L}) \left(\frac{0.20 \, \text{mol NaOH}}{1 \, \text{L}} \right) = 0.020 \, \text{mol NaOH}$$

Example 14.12



SOLUTION

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Step 2	Solve for the new molarity, taking into account that the total volume
	of the solution after dilution is $250. \text{ mL} (0.250 \text{ L})$:

$$M = \frac{0.020 \text{ mol NaOH}}{0.250 \text{ L}} = 0.080 M \text{ NaOH}$$

ALTERNATIVE SOLUTION When the moles of solute in a solution before and after dilution are the same, the moles before and after dilution may be set equal to each other. That is, $mol_1 = mol_2$

where $mol_1 = moles$ before dilution and $mol_2 = moles$ after dilution. Then

 $\operatorname{mol}_1 = \operatorname{L}_1 \times M_1$ $\operatorname{mol}_2 = \operatorname{L}_2 \times M_2$ $\mathbf{L}_1 \times M_1 = \mathbf{L}_2 \times M_2$

When both volumes are in the same units, a more general statement can be made:

 $V_1 \times M_1 = V_2 \times M_2$

For this problem,

 $V_1 = 100. \,\mathrm{mL}$ $M_1 = 0.20 M$ $V_2 = 150. \,\mathrm{mL} + 100. \,\mathrm{mL}$ $M_2 = (\mathrm{unknown})$

Then

 $(100. \text{ mL})(0.20 M) = (250. \text{ mL})M_2$

Solving for M_2 , we get

$$M_2 = \frac{(100. \text{ mL})(0.20 M)}{250. \text{ mL}} = 0.080 M \text{ NaOH}$$

Practice 14.7

Calculate the molarity of a solution prepared by diluting 125 mL of 0.400 M K₂Cr₂O₇ with 875 mL of water.

Example 14.13	How many grams of silver chloride will be precipitated by adding sufficient silver nitrate to react with 1500. mL of $0.400 M$ barium chloride solution?			
	$2 \operatorname{AgNO}_3(aq) + \operatorname{BaCl}_2(aq) \longrightarrow 2 \operatorname{AgCl}(s) + \operatorname{Ba}(\operatorname{NO}_3)_2(aq)$			
SOLUTION	This problem is a stoichiometry problem. The fact that $BaCl_2$ is in solution means that we need to consider the volume and concentration of the solution in order to determine the number of moles of $BaCl_2$ reacting.			
	Step 1	Step 1 Determine the number of moles of BaCl ₂ in 1500. mL of 0.400 M solution:		
		$M = \frac{\text{mol}}{\text{L}} \qquad \text{mol} = \text{L} \times M \qquad 1500. \text{ mL} = 1.500 \text{ L}$		
		$(1.500 \mathcal{L}) \left(\frac{0.400 \operatorname{mol} \operatorname{BaCl}_2}{\mathcal{L}} \right) = 0.600 \operatorname{mol} \operatorname{BaCl}_2$		

 \mathcal{K}

14.7 COLLIGATIVE PROPERTIES OF SOLUTIONS

Step 2 Calculate the moles and grams of AgCl:

 $mol BaCl_2 \longrightarrow mol AgCl \longrightarrow g AgCl$

 $(0.600 \text{ mot } \text{BaCt}_2)\left(\frac{2 \text{ mot } \text{AgCt}}{1 \text{ mot } \text{BaCt}_2}\right)\left(\frac{143.4 \text{ g AgCl}}{\text{ mot } \text{AgCt}}\right) = 172 \text{ g AgCl}$

Practice 14.8

How many grams of lead(II) iodide will be precipitated by adding sufficient $Pb(NO_3)_2$ to react with 750 mL of 0.250 *M* KI solution?

 $2 \operatorname{KI}(aq) + \operatorname{Pb}(\operatorname{NO}_3)_2(aq) \longrightarrow \operatorname{PbI}_2(s) + 2 \operatorname{KNO}_3(aq)$

14.7 Colligative Properties of Solutions

Two solutions—one containing 1 mol (60.06 g) of urea (NH₂CONH₂) and the other containing 1 mol (342.3 g) of sucrose (C₁₂H₂₂O₁₁) each in 1 kg of water—both have a freezing point of -1.86° C, not 0°C as for pure water. Urea and sucrose are distinct substances, yet they lower the freezing point of the water by the same amount. The only thing apparently common to these two solutions is that each contains 1 mol (6.022×10^{23} molecules) of solute and 1 kg of solvent. In fact, when we dissolve 1 mol of any nonionizable solute in 1 kg of water, the freezing point of the resulting solution is -1.86° C.

These results lead us to conclude that the freezing point depression for a solution containing 6.022×10^{23} solute molecules (particles) and 1 kg of water is a constant, namely, 1.86°C. Freezing point depression is a general property of solutions. Furthermore, the amount by which the freezing point is depressed is the same for all solutions made with a given solvent; that is, each solvent shows a characteristic *freezing point depression constant*. Freezing point depression constants for several solvents are given in Table 14.5.

The solution formed by the addition of a nonvolatile solute to a solvent has a lower freezing point, a higher boiling point, and a lower vapor pressure than that of the pure solvent. These effects are related and are known as colligative

F o Solvent	Freezing point f pure solvent (°C)	Freezing po depression cons $\left(\frac{^{\circ}C \text{ kg solve}}{\text{mol solut}}\right)$	$\left(\begin{array}{c} \operatorname{bint} & & \\ \operatorname{tant}, K_f & & \\ \operatorname{Boiling point} & \\ \operatorname{of pure solvent} & \\ & (^{\circ}\mathrm{C}) \end{array} \right)$	Boiling point elevation constant, K_b $\left(\frac{^{\circ}C \text{ kg solvent}}{\text{mol solute}}\right)$
Water	0.00	1.86	100.0	0.512
Acetic acid	16.6	3.90	118.5	3.07
Benzene	5.5	5.1	80.1	2.53
Camphor	178	40	208.2	5.95

Table 14.5Freezing Point Depression and Boiling Point Elevation
Constants of Selected Solvents



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(a) boiling point elevation and(b) freezing point depression

Figure 14.8

effects (concentration: 1 mol solute/1 kg water).

Vapor pressure curves of pure water and water solutions, showing

colligative properties

properties. The **colligative properties** are properties that depend only on the number of solute particles in a solution, not on the nature of those particles. Freezing point depression, boiling point elevation, and vapor pressure lowering are colligative properties of solutions.

The colligative properties of a solution can be considered in terms of vapor pressure. The vapor pressure of a pure liquid depends on the tendency of molecules to escape from its surface. If 10% of the molecules in a solution are nonvolatile solute molecules, the vapor pressure of the solution is 10% lower than that of the pure solvent. The vapor pressure is lower because the surface of the solution contains 10% nonvolatile molecules and 90% of the volatile solvent molecules. A liquid boils when its vapor pressure equals the pressure of the atmosphere. We can thus see that the solution just described as having a lower vapor pressure will have a higher boiling point than the pure solvent. The solution with a lowered vapor pressure doesn't boil until it has been heated above the boiling point of the solvent (see Figure 14.8a). Each solvent has its own characteristic boiling point elevation constant (Table 14.5). The boiling point elevation constant is based on a solution that contains 1 mol of solute particles per kilogram of solvent. For example, the boiling point elevation constant for a solution containing 1 mol of solute particles per kilogram of water is 0.512°C, which means that this water solution will boil at 100.512°C.

The freezing behavior of a solution can also be considered in terms of lowered vapor pressure. Figure 14.8b shows the vapor pressure relationships of ice, water, and a solution containing 1 mol of solute per kilogram of water. The freezing point of water is at the intersection of the liquid and solid vapor pressure curves (i.e., at the point where water and ice have the same vapor pressure). Because the vapor pressure of the liquid is lowered by the solute, the vapor pressure curve of the solution does not intersect the vapor pressure curve of the solid until the solution has been cooled below the freezing point of pure water. So the solution must be cooled below 0°C in order for it to freeze.

The foregoing discussion dealing with freezing point depressions is restricted to *un-ionized* substances. The discussion of boiling point elevations is restricted to *nonvolatile* and un-ionized substances. The colligative properties of ionized substances are not under consideration at this point; we will discuss them in Chapter 15.



Engine coolant is one application of colligative properties. The addition of coolant to the water in a radiator raises its boiling point and lowers its freezing point.

14.7 COLLIGATIVE PROPERTIES OF SOLUTIONS



Sodium chloride or calcium chloride is used to melt ice on snowy streets and highways.

Some practical applications involving colligative properties are (1) use of salt-ice mixtures to provide low freezing temperatures for homemade ice cream, (2) use of sodium chloride or calcium chloride to melt ice from streets, and (3) use of ethylene glycol-water mixtures as antifreeze in automobile radiators (ethylene glycol also raises the boiling point of radiator fluid, thus allowing the engine to operate at a higher temperature).

Both the freezing point depression and the boiling point elevation are directly proportional to the number of moles of solute per kilogram of solvent. When we deal with the colligative properties of solutions, another concentration expression, *molality*, is used. The **molality** (*m*) of a solute is the number of moles of solute per kilogram of solvent:

molality (m)

$m = \frac{\text{mol solute}}{\text{kg solvent}}$

Note that a lowercase *m* is used for molality concentrations and a capital *M* for molarity. The difference between molality and molarity is that molality refers to moles of solute *per kilogram of solvent*, whereas molarity refers to moles of solute *per liter of solution*. For un-ionized substances, the colligative properties of a solution are directly proportional to its molality.

Molality is independent of volume. It is a mass-to-mass relationship of solute to solvent and allows for experiments, such as freezing point depression and boiling point elevation, to be conducted at variable temperatures.

The following equations are used in calculations involving colligative properties and molality:

$$\Delta t_f = mK_f$$
 $\Delta t_b = mK_b$ $m = \frac{\text{mol solute}}{\text{kg solvent}}$

m =molality; mol solute/kg solvent

 Δt_f = freezing point depression; °C

 Δt_b = boiling point elevation; °C

 K_f = freezing point depression constant; °C kg solvent/mol solute

 K_b = boiling point elevation constant; °C kg solvent/mol solute

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CHAPTER 14 SOLUTIONS

Example 14.14	What is the molality (<i>m</i>) of a solution prepared by dissolving 2.70 g CH ₃ OH in 25.0 g H ₂ O?
SOLUTION	Since $m = \frac{\text{mol solute}}{\text{kg solvent}}$, the conversion is
	$\frac{2.70 \text{ g CH}_3\text{OH}}{25.0 \text{ g H}_2\text{O}} \longrightarrow \frac{\text{mol CH}_3\text{OH}}{25.0 \text{ g H}_2\text{O}} \longrightarrow \frac{\text{mol CH}_3\text{OH}}{1 \text{ kg H}_2\text{O}}$
	The molar mass of CH ₃ OH is $(12.01 + 4.032 + 16.00)$, or 32.04 g/mol:
	$\left(\frac{2.70 \text{ gCH}_{3}\text{OH}}{25.0 \text{ gH}_{2}\text{O}}\right) \left(\frac{1 \text{ mol CH}_{3}\text{OH}}{32.04 \text{ gCH}_{3}\text{OH}}\right) \left(\frac{1000 \text{ gH}_{2}\text{O}}{1 \text{ kg H}_{2}\text{O}}\right) = \frac{3.37 \text{ mol CH}_{3}\text{OH}}{1 \text{ kg H}_{2}\text{O}}$
	The molality is $3.37 m$.
	Practice 14.9
	What is the molality of a solution prepared by dissolving 150.0 g $C_6H_{12}O_6$ in 600.0 g H_2O ?
Example 14.15	A solution is made by dissolving 100. g of ethylene glycol ($C_2H_6O_2$) in 200. g of water. What is the freezing point of this solution?
SOLUTION	To calculate the freezing point of the solution, we first need to calculate Δt_f , the change in freezing point. Use the equation
	$\Delta t_f = mK_f = \frac{\text{mol solute}}{\text{kg solvent}} \times K_f$
	K_f (for water): $\frac{1.86^{\circ}\text{C kg solvent}}{\text{mol solute}}$ (from Table 14.5)
	mol solute: $(100. \text{gC}_2\text{H}_6\text{O}_2) \left(\frac{1 \text{ mol } \text{C}_2\text{H}_6\text{O}_2}{62.07 \text{gC}_2\text{H}_6\text{O}_2} \right) = 1.61 \text{ mol } \text{C}_2\text{H}_6\text{O}_2$
	kg solvent: $(200. \text{g H}_2\text{O}) \left(\frac{1 \text{ kg}}{1000 \text{ g}}\right) = 0.200 \text{ kg H}_2\text{O}$
	$\Delta t_f = \left(\frac{1.61 \text{ mol} \text{C}_2 \text{H}_6 \text{O}_2}{0.200 \text{kg} \text{H}_2 \text{O}}\right) \left(\frac{1.86^{\circ} \text{C} \text{kg} \text{H}_2 \text{O}}{1 \text{mol} \text{C}_2 \text{H}_6 \text{O}_2}\right) = 15.0^{\circ} \text{C}$
	The freezing point depression, 15.0°C, must be subtracted from 0°C, the freez- ing point of the pure solvent (water):
	freezing point of solution = freezing point of solvent $-\Delta t_f$
	$= 0.0^{\circ}C - 15.0^{\circ}C = -15.0^{\circ}C$
	Therefore, the freezing point of the solution is -15.0 °C. This solution will protect an automobile radiator down to -15.0 °C (5 °F).



CHEMISTRY IN ACTION • The Scoop on Ice Cream

ce cream is mainly composed of water (from milk and cream), milk solids, milk fats, and, frequently, various sweet-

eners (corn syrup or sugar), flavorings, emulsifiers, and stabilizers. But that smooth, creamy, rich flavor and texture are the result of the chemistry of the mixing and freezing process. The rich, smooth texture of great ice cream results from the milk fat. By law, if the carton is labeled "ice cream," it must contain a minimum of 10% milk fat. That carton of ice cream also contains 20%–50% air whipped into the ingredients during the initial mixing process.

H. Douglas Goff, a professor of food science and ice cream expert from Ontario, Canada, says, "There are no real chemical reactions that take place when you make ice cream, but that doesn't mean that there isn't plenty of chemistry." The structure of ice cream contributes greatly to its taste. Tiny air bubbles are formed in the initial whipping process. These bubbles are distributed through a network of fat globules and liquid water. The milk fat has surface proteins on the globules in the milk to keep the fat dissolved in solution. Ice cream manufacturers destabilize these globules using emulsifiers (such as egg yolks, mono- or



diglycerides) and let them come together in larger networks sort of like grape clusters.

Once the mixture is fully whipped, it is cooled to begin the freezing process. But ice cream does not freeze at 0°C even though it is 55–64% water. The freezing point is depressed as a colligative property of the ice cream solution. Once the freezing of the water begins, the concentration of the solution increases, which continues to depress the freezing point. Goff tells us that even at a typical serving temperature (-16° C for most ice cream), only about 72% of the water in the ice cream is frozen. The unfrozen solution keeps the ice cream "scoopable."

One last factor that affects the quality of ice cream is the size of the ice crystals. For very smooth ice cream, tiny crystals are needed. To produce these, the ice cream must freeze very slowly. Large crystals give a coarse, grainy texture. Now, as you savor that premium ice cream cone, you'll know just how colligative properties and the chemistry of freezing helped make it so delicious!

A solution made by dissolving 4.71 g of a compound of unknown molar mass in 100.0 g of water has a freezing point of -1.46 °C. What is the molar mass of the compound?

First substitute the data in $\Delta t_f = mK_f$ and solve for *m*:

$$\Delta t_f = +1.46$$
 (since the solvent, water, freezes at 0°C)

$$K_{f} = \frac{1.86^{\circ} \text{C kg H}_{2}\text{O}}{\text{mol solute}}$$

$$1.46^{\circ} \text{C} = mK_{f} = m \times \frac{1.86^{\circ} \text{C kg H}_{2}\text{O}}{\text{mol solute}}$$

$$m = \frac{1.46^{\circ} \text{C} \times \text{mol solute}}{1.86^{\circ} \text{C} \times \text{kg H}_{2}\text{O}} = \frac{0.785 \text{ mol solute}}{\text{kg H}_{2}\text{O}}$$

Now convert the data, 4.71 g solute/100.0 g H_2O , to g/mol:

$$\left(\frac{4.71 \text{ g solute}}{100.0 \text{ gH}_2 \text{O}}\right) \left(\frac{1000 \text{ gH}_2 \text{O}}{1 \text{ kgH}_2 \text{O}}\right) \left(\frac{1 \text{ kgH}_2 \text{O}}{0.785 \text{ mol solute}}\right) = 60.0 \text{ g/mol}$$

The molar mass of the compound is 60.0 g/mol.

Example 14.16



SOLUTION

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Practice 14.10

What is the freezing point of the solution in Practice Exercise 14.9? What is the boiling point?

14.8 Osmosis and Osmotic Pressure

When red blood cells are put into distilled water, they gradually swell and in time may burst. If red blood cells are put in a 5%-urea (or a 5%-salt) solution, they gradually shrink and take on a wrinkled appearance. The cells behave in this fashion because they are enclosed in semipermeable membranes. A **semipermeable membrane** allows the passage of water (solvent) molecules through it in either direction but prevents the passage of larger solute molecules or ions. When two solutions of different concentrations (or water and a water solution) are separated by a semipermeable membrane, water diffuses through the membrane from the solution of lower concentration into the solution of higher concentration. The diffusion of water, either from a dilute solution or from pure water, through a semipermeable membrane into a solution of higher concentration is called **osmosis**.

A 0.90% (0.15 *M*) sodium chloride solution is known as a *physiological saline solution* because it is *isotonic* with blood plasma; that is, it has the same concentration of NaCl as blood plasma. Because each mole of NaCl yields about 2 mol of ions when in solution, the solute particle concentration in physiological saline solution is nearly 0.30 *M*. Five-percent-glucose solution (0.28 *M*) is also approximately isotonic with blood plasma. Blood cells neither swell nor shrink in an isotonic solution. The cells described in the preceding paragraph swell in water because water is *bypotonic* to cell plasma. The cells shrink in 5%-urea solution because the urea solution is *bypertonic* to the cell plasma. To prevent possible injury to blood cells by osmosis, fluids for intravenous use are usually made up at approximately isotonic concentration.

semipermeable membrane

osmosis

Human red blood cells. Left: In an isotonic solution the concentration is the same inside and outside the cell (0.9% saline). Center: In a hypertonic solution (1.6% saline) water leaves the cells, causing them to crenate (shrink). Right: In a hypotonic solution (0.2% saline) the cells swell as water moves into the cell center. Magnification is $260,000 \times$.







14.8 OSMOSIS AND OSMOTIC PRESSURE



Figure 14.9

Laboratory demonstration of osmosis: As a result of osmosis, water passes through the membrane, causing the solution to rise in the thistle tube.

All solutions exhibit *osmotic pressure*, which is another colligative property. Osmotic pressure is a pressure difference between the system and atmospheric pressure. The osmotic pressure of a system can be measured by applying enough pressure to stop the flow of water due to osmosis in the system. The difference between the applied pressure and atmospheric pressure is the **osmotic pressure**. When pressure greater than the osmotic pressure is applied to a system, the flow of water can be reversed from that of osmosis. This process can be used to obtain useful drinking water from seawater and is known as *reverse osmosis*. Osmotic pressure is dependent only on the concentration of the solute particles and is independent of their nature. The osmotic pressure needed to prevent osmosis; this pressure can be very large. The osmotic pressure of a solution containing 1 mol of solute particles in 1 kg of water is about 22.4 atm, which is about the same as the pressure exerted by 1 mol of a gas confined in a volume of 1 L at 0°C.

Osmosis has a role in many biological processes, and semipermeable membranes occur commonly in living organisms. An example is the roots of plants, which are covered with tiny structures called root hairs; soil water enters the plant by osmosis, passing through the semipermeable membranes covering the root hairs. Artificial or synthetic membranes can also be made.

Osmosis can be demonstrated with the simple laboratory setup shown in Figure 14.9. As a result of osmotic pressure, water passes through the cellophane membrane into the thistle tube, causing the solution level to rise. In osmosis, the net transfer of water is always from a less concentrated to a more concentrated solution; that is, the effect is toward equalization of the concentration on both sides of the membrane. Note that the effective movement of water in osmosis is always from the region of *higher water concentration* to the region of *lower water concentration*.

Osmosis can be explained by assuming that a semipermeable membrane has passages that permit water molecules and other small molecules to pass in either direction. Both sides of the membrane are constantly being struck by water molecules in random motion. The number of water molecules crossing the

osmotic pressure

membrane is proportional to the number of water molecule-to-membrane impacts per unit of time. Because the solute molecules or ions reduce the concentration of water, there are more water molecules and thus more water molecule impacts on the side with the lower solute concentration (more dilute solution). The greater number of water molecule-to-membrane impacts on the dilute side thus causes a net transfer of water to the more concentrated solution. Again, note that the overall process involves the net transfer, by diffusion through the membrane, of water molecules from a region of higher water concentrated solution).

This is a simplified picture of osmosis. No one has ever seen the hypothetical passages that allow water molecules and other small molecules or ions to pass through them. Alternative explanations have been proposed, but our discussion has been confined to water solutions. Osmotic pressure is a general colligative property, however, and is known to occur in nonaqueous systems.

Chapter 14 Review

14.1 General Properties of Solutions

KEY TERMS

Solution Solute Solvent

- A solution is a homogeneous mixture of two or more substances:
 - Consists of solvent—the dissolving agent—and solute—the component(s) dissolved in the solvent
 - Is a homogeneous mixture
 - Contains molecular or ionic particles
 - Can be colored or colorless
 - Can be separated into solute and solvent by a physical separation process

14.2 Solubility

KET	I EKIN 2	
Solu	bility	

Miscible	
mmiscible	
Concentration of a solution	

- Solubility describes the amount of solute that will dissolve in a specified amount of solvent.
- Solubility can also be qualitative.
- General guideline for ionic solubility are:



REVIEW QUESTIONS

- Liquids can also be classified as miscible (soluble in each other) or immiscible (not soluble in each other).
- The concentration of a solution is the quantitative measurement of the amount of solute that is dissolved in a solution.

14.3 Factors Related to Solubility

KEY TERMS

Saturated solution Unsaturated solution

- Supersaturated solution
- Like tends to dissolve like is a general rule for solvents and solutes.
- As temperature increases:
 - Solubility of a solid in a liquid tends to increase.
 - Solubility of a gas in a liquid tends to decrease.
- As pressure increases:
- Solubility of a solid in a liquid remains constant.Solubility of a gas in a liquid tends to increase.
- At a specific temperature, the amount of solute that can dissolve in a solvent has a limit:
- Unsaturated solutions contain less solute than the limit.
- Saturated solutions contain dissolved solute at the limit.
 Supersaturated solutions contain more solute than
- the limit and are therefore unstable:If disturbed, the excess solute will precipitate out of solution.

14.4 Rate of Dissolving Solids

- The rate at which a solute dissolves is determined by these factors:
 - Particle size
 - Temperature
 - Concentration of solution
 - Agitation

14.5 Solutions: A Reaction Medium

- Molecules or ions must collide in order to react.
- Solutions provide a medium for the molecules or ions to collide.

14.6 Concentration of Solutions

KEY TERMS

- Dilute solution Concentrated solution Parts per million (ppm)
- **Review Questions**

Molarity (M)

All questions with blue numbers have answers in the appendix of the text.

- 1. Sketch the orientation of water molecules (a) about a single sodium ion and (b) about a single chloride ion in solution.
- **2.** Estimate the number of grams of sodium fluoride that would dissolve in 100 g of water at 50°C. (Table 14.2)

• Concentrations can be measured in many ways:

Mass percent	% m/m	$\frac{\text{mass solute}}{100}$
mass percent	/0 111/111	mass solution
Parts per million	ppm	$\frac{\text{mass solute}}{\text{mass solution}} \times 1,000,000$
Mass/volume	% m/v	$\frac{\text{mass solute}}{\text{mL solution}} \times 100$
Volume percent	% v/v	$\frac{mL \text{ solute}}{mL \text{ solution}} \times 100$
Molarity	М	moles solute
Molality	т	moles solute kg solvent

- Dilution of solutions requires the addition of more solvent to an existing solution:
 - The number of moles in the diluted solution is the same as that in the original solution.
 - $M_1V_1 = M_2V_2$.

14.7 Colligative Properties of Solutions

KEY TERMS

Colligative properties

Molality (m)

- Properties of a solution that depend only on the number of solute particles in solution are called colligative properties:
 - Freezing point depression
 - $\Delta t_f = mK_f$
 - Boiling point elevation
 - $\Delta t_b = mK_b$
 - Osmotic pressure
- Molality is used in working with colligative properties.

14.8 Osmosis and Osmotic Pressure

KEY TERMS

Semipermeable membrane Osmosis

- Osmotic pressure
- Osmosis is the diffusion of water through a semipermeable membrane:
- Occurs from dilute solution to a solution of higher concentration
- Osmosis results in osmotic pressure, which is a colligative property of a solution.
- 3. What is the solubility at 25°C of these substances (Figure 14.4)?
 - (a) potassium chloride
 - (b) potassium chlorate
 - (c) potassium nitrate

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- **4.** What is different in the solubility trend of the potassium halides compared with that of the lithium halides and the sodium halides? (Table 14.2)
- 5. What is the solubility, in grams of solute per 100 g of H₂O, of (a) KClO₃ at 60°C; (b) HCl at 20°C; (c) Li₂SO₄ at 80°C; and (d) KNO₃ at 0°C? (Figure 14.4)
- 6. Which substance, KNO₃ or NH₄Cl, shows the greater increase in solubility with increased temperature? (Figure 14.4)
- Does a 2-molal solution in benzene or a 1-molal solution in camphor show the greater freezing point depression? (Table 14.5)
- 8. What would be the total surface area if the 1-cm cube in Figure 14.5 were cut into cubes 0.01 cm on a side?
- At which temperatures—10°C, 20°C, 30°C, 40°C, or 50°C—would you expect a solution made from 63 g of ammonium chloride and 150 g of water to be unsaturated? (Figure 14.4)
- **10.** Explain why the rate of dissolving decreases. (Figure 14.6)
- **11.** Explain how a supersaturated solution of NaC₂H₃O₂ can be prepared and proven to be supersaturated.
- **12.** Assume that the thistle tube in Figure 14.9 contains 1.0 *M* sugar solution and that the water in the beaker has just been replaced by a 2.0 *M* solution of urea. Would the solution level in the thistle tube continue to rise, remain constant, or fall? Explain.
- **13.** What is a true solution?
- **14.** Name and distinguish between the two components of a solution.
- **15.** Is it always apparent in a solution which component is the solute, for example, in a solution of a liquid in a liquid?
- **16.** Explain why the solute does not settle out of a solution.
- 17. Is it possible to have one solid dissolved in another? Explain.
- **18.** An aqueous solution of KCl is colorless, KMnO₄ is purple, and K₂Cr₂O₇ is orange. What color would you expect of an aqueous solution of Na₂Cr₂O₇? Explain.
- **19.** Explain why hexane will dissolve benzene but will not dissolve sodium chloride.
- **20.** Some drinks like tea are consumed either hot or cold, whereas others like Coca-Cola are drunk only cold. Why?
- **21.** Why is air considered to be a solution?
- **22.** In which will a teaspoonful of sugar dissolve more rapidly, 200 mL of iced tea or 200 mL of hot coffee? Explain in terms of the KMT.
- **23.** What is the effect of pressure on the solubility of gases in liquids? of solids in liquids?
- 24. Why do smaller particles dissolve faster than large ones?
- **25.** In a saturated solution containing undissolved solute, solute is continuously dissolving, but the concentration of the solution remains unchanged. Explain.
- **26.** Explain why there is no apparent reaction when crystals of AgNO₃ and NaCl are mixed, but a reaction is apparent immediately when solutions of AgNO₃ and NaCl are mixed.

- **27.** Why do salt trucks distribute salt over icy roads in the winter?
- **28.** What do we mean when we say that concentrated nitric acid (HNO₃) is 16 molar?
- **29.** Will 1 L of 1 *M* NaCl contain more chloride ions than 0.5 L of 1 *M* MgCl₂? Explain.
- **30.** Champagne is usually cooled in a refrigerator prior to opening. It's also opened very carefully. What would happen if a warm bottle of champagne were shaken and opened quickly and forcefully?
- **31.** Describe how you would prepare 750 mL of 5.0 *M* NaCl solution.
- **32.** Explain in terms of the KMT how a semipermeable membrane functions when placed between pure water and a 10% sugar solution.
- **33.** Which has the higher osmotic pressure, a solution containing 100 g of urea (NH₂CONH₂) in 1 kg H₂O or a solution containing 150 g of glucose, C₆H₁₂O₆, in 1 kg H₂O?
- **34.** Explain why a lettuce leaf in contact with salad dressing containing salt and vinegar soon becomes wilted and limp whereas another lettuce leaf in contact with plain water remains crisp.
- **35.** A group of shipwreck survivors floated for several days on a life raft before being rescued. Those who had drunk some seawater were found to be suffering the most from dehydration. Explain.
- **36.** Arrange the following bases (in descending order) according to the volume of each that will react with 1 L of 1 *M* HCl:
 - (a) 1 *M* NaOH
 - (b) $1.5 M Ca(OH)_2$
 - (c) 2 *M* KOH
 - (d) $0.6 M Ba(OH)_2$
- **37.** Explain in terms of vapor pressure why the boiling point of a solution containing a nonvolatile solute is higher than that of the pure solvent.
- **38.** Explain why the freezing point of a solution is lower than the freezing point of the pure solvent.
- **39.** When water and ice are mixed, the temperature of the mixture is 0° C. But if methyl alcohol and ice are mixed, a temperature of -10° C is readily attained. Explain why the two mixtures show such different temperature behavior.
- **40.** Which would be more effective in lowering the freezing point of 500. g of water?
 - (a) 100. g of sucrose, $C_{12}H_{22}O_{11}$, or 100. g of ethyl alcohol, C_2H_5OH
 - (b) 100. g of sucrose or 20.0 g of ethyl alcohol
 - (c) 20.0 g of ethyl alcohol or 20.0 g of methyl alcohol, CH₃OH.
- **41.** What is the difference between molarity and molality?
- **42.** Is the molarity of a 5 *m* aqueous solution of NaCl greater or less than 5 *M*? Explain.

PAIRED EXERCISES

Paired Exercises

All exercises with blue numbers have answers in the appendix of the text.

 Which of the substances listed below are reasonably soluble and which are insoluble in water? (See Figure 14.2 or Appendix V.)
 (a) KOH
 (b) NiCl₂
 (c) ZnS

Percent Solution

- Calculate the mass percent of the following solutions:
 (a) 15.0 g KCl + 100.0 g H₂O
 (b) 2.50 g Na₃PO₄ + 10.0 g H₂O
 (c) 0.20 mol NH₄C₂H₃O₂ + 125 g H₂O
 (d) 1.50 mol NaOH in 33.0 mol H₂O
- **5.** A chemistry lab experiment requires 25.2 g of silver nitrate. How many grams of a 15.5% by mass solution of silver nitrate should be used?
- 7. In 25 g of a 7.5% by mass solution of CaSO₄
 (a) how many grams of solute are present?
 (b) how many grams of solvent are present?
- **9.** Calculate the mass/volume percent of a solution made by dissolving 22.0 g of CH₃OH (methanol) in C₂H₅OH (ethanol) to make 100. mL of solution.
- **11.** What is the volume percent of 10.0 mL of CH₃OH (methanol) dissolved in water to a volume of 40.0 mL?

Molarity

- 13. Calculate the molarity of the following solutions:(a) 0.25 mol of solute in 75.0 mL of solution(b) 1.75 mol of KBr in 0.75 L of solution
 - (c) $35.0 \text{ g of } \text{NaC}_2\text{H}_3\text{O}_2$ in 1.25 L of solution
 - (d) 75 g of $CuSO_4 \cdot 5 H_2O$ in 1.0 L of solution
- 15. Calculate the number of moles of solute in each of the following solutions:
 (a) 1.5 L of 1.20 M H₂SO₄
 (b) 25.0 mL of 0.0015 M BaCl₂
 - (c) 125 mL of $0.35 M K_3 PO_4$
- 17. Calculate the grams of solute in each of the following solutions:(a) 2.5 L of 0.75 M K₂CrO₄
 - (b) 75.2 mL of 0.050 M HC₂H₃O₂
 - (c) 250 mL of 16 *M* HNO₃
- How many milliliters of 0.750 *M* H₃PO₄ will contain the following?
 (a) 0.15 mol H₃PO₄
 (b) 35.5 g H₃PO₄

- Which of the substances listed below are reasonably soluble and which are insoluble in water?
 (See Figure 14.2 or Appendix V.)
 (a) PbI₂
 (b) MgCO₃
 (c) CaCl₂
- 4. Calculate the mass percent of the following solutions:
 (a) 25.0 g NaNO₃ in 125.0 g H₂O
 (b) 1.25 g CaCl₂ in 35.0 g H₂O
 (c) 0.75 mol K₂CrO₄ in 225 g H₂O
 (d) 1.20 mol H₂SO₄ in 72.5 mol H₂O
- 6. A reaction requires 25.0 g of sodium chloride. How many grams of a 10.0% by mass solution would provide this amount of solute?
- 8. In 75 g of a 12.0% by mass solution of BaCl₂
 (a) how many grams of solute are present?
 (b) how many grams of solvent are present?
- 10. Calculate the mass/volume percent of a solution made by dissolving 4.20 g of NaCl in H_2O to make 12.5 mL of solution.
- **12.** What is the volume percent of 2.0 mL of hexane, C_6H_{14} , dissolved in benzene, C_6H_6 , to a volume of 9.0 mL?
- 14. Calculate the molarity of the following solutions:
 (a) 0.50 mol of solute in 125 mL of solution
 (b) 2.25 mol of CaCl₂ in 1.50 L of solution
 (c) 275 g C₆H₁₂O₆ in 775 mL of solution
 (d) 125 g MgSO₄ · 7 H₂O in 2.50 L of solution
- 16. Calculate the number of moles of solute in each of the following solutions:
 (a) 0.75 L of 1.50 *M* HNO₃
 (b) 10.0 mL of 0.75 *M* NaClO₃
 (c) 175 mL of 0.50 *M* LiBr
- 18. Calculate the grams of solute in each of the following solutions:(a) 1.20 L of 18 M H₂SO₄
 - (a) $1.20 \pm 0110 \text{ M} \text{ H}_2\text{SO}_4$
 - (b) 27.5 mL of 1.50 *M* KMnO₄
 - (c) 120 mL of 0.025 $M \operatorname{Fe}_2(SO_4)_3$
- 20. How many milliliters of 0.250 *M* NH₄Cl will contain the following?
 (a) 0.85 mol NH₄Cl
 (b) 25.2 g NH₄Cl

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Dilution

- **21.** What will be the molarity of the resulting solutions made by mixing the following? Assume that volumes are additive.
 - (a) 125 mL of 5.0 M H₃PO₄ with 775 mL of H₂O
 - (b) 250 mL of 0.25 M Na₂SO₄ with 750 mL of H₂O
 - (c) 75 mL of 0.50 M HNO₃ with 75 mL of 1.5 M HNO₃
- 23. Calculate the volume of concentrated reagent required to prepare the diluted solutions indicated:
 (a) 15 *M* H₃PO₄ to prepare 750 mL of 3.0 *M* H₃PO₄
 (b) 16 *M* HNO₃ to prepare 250 mL of 0.50 *M* HNO₃
- 25. Calculate the molarity of the solutions made by mixing 125 mL of 6.0 *M* HC₂H₃O₂ with the following:
 (a) 525 mL of H₂O
 (b) 175 mL of 1.5 *M* HC₂H₃O₂

Stoichiometry

27. Use the equation to calculate the following:

 $3 \operatorname{Ca(NO_3)}_2(aq) + 2 \operatorname{Na_3PO_4}(aq) \rightarrow$

 $\operatorname{Ca}_{3}(\operatorname{PO}_{4})_{2}(s) + 6 \operatorname{NaNO}_{3}(aq)$

- (a) the moles $Ca_3(PO_4)_2$ produced from 2.7 mol Na_3PO_4
- (b) the moles NaNO₃ produced from 0.75 mol Ca(NO₃)₂
- (c) the moles Na₃PO₄ required to react with 1.45 L of 0.225 *M* Ca(NO₃)₂
- (d) the grams of Ca₃(PO₄)₂ that can be obtained from 125 mL of 0.500 *M* Ca(NO₃)₂
- (e) the volume of 0.25 *M* Na₃PO₄ needed to react with 15.0 mL of 0.50 *M* Ca(NO₃)₂
- (f) the molarity (M) of the $Ca(NO_3)_2$ solution when 50.0 mL react with 50.0 mL of 2.0 M Na_3PO_4
- **29.** Use the equation to calculate the following:
 - $2 \text{ KMnO}_4(aq) + 16 \text{ HCl}(aq) \rightarrow$

 $2 \operatorname{MnCl}_2(aq) + 5 \operatorname{Cl}_2(g) + 8 \operatorname{H}_2O(l) + 2 \operatorname{KCl}(aq)$

- (a) the moles of H₂O that can be obtained from 15.0 mL of 0.250 *M* HCl
- (b) the volume of 0.150 $M~{\rm KMnO_4}$ needed to produce 1.85 mol ${\rm MnCl_2}$
- (c) the volume of 2.50 M HCl needed to produce 125 mL of 0.525 M KCl
- (d) the molarity (*M*) of the HCl solution when 22.20 mL react with 15.60 mL of 0.250 *M* KMnO₄
- (e) the liters of Cl_2 gas at STP produced by the reaction of 125 mL of 2.5 *M* HCl
- (f) the liters of Cl₂ gas at STP produced by the reaction of 15.0 mL of 0.750 *M* HCl and 12.0 mL of 0.550 *M* KMnO₄

Molality and Colligative Properties

31. Calculate the molality of these solutions:
(a) 14.0 g CH₃OH in 100.0 g of H₂O
(b) 2.50 mol of benzene (C₆H₆) in 250 g of hexane (C₆H₁₄)

22. What will be the molarity of the resulting solutions made by mixing the following? Assume that volumes are additive.

CHAPTER 14 SOLUTIONS

- (a) 175 mL of 3.0 $M\,{\rm H_2SO_4}$ with 275 mL of ${\rm H_2O}$
- (b) 350 mL of 0.10 M CuSO₄ with 150 mL of H₂O
- (c) 50.0 mL of 0.250 *M* HCl with 25.0 mL of 0.500 *M* HCl
- 24. Calculate the volume of concentrated reagent required to prepare the diluted solutions indicated:
 (a) 18 M H₂SO₄ to prepare 225 mL of 2.0 M H₂SO₄
 (b) 15 M NH₃ to prepare 75 mL of 1.0 M NH₃
- 26. Calculate the molarity of the solutions made by mixing 175 mL of 3.0 *M* HCl with the following:
 (a) 250 mL of H₂O
 (b) 115 mL of GO MUSI
 - (b) 115 mL of 6.0 *M* HCl
- **28.** Use the equation to calculate the following:

 $2 \operatorname{NaOH}(aq) + \operatorname{H}_2 \operatorname{SO}_4(aq) \rightarrow \operatorname{Na}_2 \operatorname{SO}_4(aq) + 2 \operatorname{H}_2 \operatorname{O}(l),$

- (a) the moles Na_2SO_4 produced from 3.6 mol H_2SO_4
- (b) the moles H₂O produced from 0.025 mol NaOH
- (c) the moles NaOH required to react with 2.50 L of $0.125 \ M \ H_2 SO_4$
- (d) the grams of Na_2SO_4 that can be obtained from 25 mL of 0.050 M NaOH
- (e) the volume of 0.250 M H₂SO₄ needed to react with 25.5 mL of 0.750 M NaOH
- (f) the molarity (M) of the NaOH solution when 48.20 mL react with 35.72 mL of 0.125 M H₂SO₄
- **30.** Use the equation to calculate the following:

 $K_2CO_3(aq) + 2 HC_2H_3O_2(aq) \rightarrow$

 $2 \text{ KC}_2 \text{H}_3 \text{O}_2(aq) + \text{H}_2 \text{O}(l) + \text{CO}_2(g)$

- (a) the moles of H_2O that can be obtained from 25.0 mL of 0.150 M HC₂H₃O₂
- (b) the volume of 0.210 *M* K₂CO₃ needed to produce 17.5 mol KC₂H₃O₂
- (c) the volume of 1.25 M HC₂H₃O₂ needed to react with 75.2 mL 0.750 M K₂CO₃
- (d) the molarity (M) of the HC₂H₃O₂ solution when 10.15 mL react with 18.50 mL of 0.250 M K₂CO₃
- (e) the liters of CO₂ gas at STP produced by the reaction of 105 mL of $1.5 M HC_2H_3O_2$
- (f) the liters of CO₂ gas at STP produced by the reaction of 25.0 mL of 0.350 M K₂CO₃ and 25.0 mL of 0.250 M HC₂H₃O₂
- 32. Calculate the molality of these solutions:
 (a) 1.0 g C₆H₁₂O₆ in 1.0 g H₂O
 - (b) 0.250 mol iodine in 1.0 kg H_2O

ADDITIONAL EXERCISES

- **33.** What is the (a) molality, (b) freezing point, and (c) boiling point of a solution containing 2.68 g of naphthalene $(C_{10}H_8)$ in 38.4 g of benzene (C_6H_6) ?
- *35. The freezing point of a solution of 8.00 g of an unknown *36. compound dissolved in 60.0 g of acetic acid is 13.2°C. Calculate the molar mass of the compound.

Additional Exercises

All exercises with blue numbers have answers in the appendix of the text.

- **37.** What happens to salt (NaCl) crystals when they are dissolved in water?
- **38.** What happens to sugar molecules $(C_{12}H_{22}O_{11})$ when they are dissolved in water?
- **39.** Why do sugar and salt behave differently when dissolved in water?
- **40.** Why don't blood cells shrink or swell in an isotonic sodium *** 52.** chloride solution (0.9% saline)?
- **41.** In the picture of dissolving KMnO₄ found in Section 14.1, the compound is forming purple streaks as it dissolves. Why?
- **42.** In Figure 14.4, observe the line for KNO₃. Explain why it slopes up from left to right. How does the slope compare to the slopes of the other substances? What does this mean?
- **43.** How many grams of solution, 10.0% NaOH by mass, are required to neutralize 150 mL of a 1.0 *M* HCl solution?
- *44. How many grams of solution, 10.0% NaOH by mass, are required to neutralize 250.0 g of a 1.0 *m* solution of HCI?
- * **45.** A sugar syrup solution contains 15.0% sugar, C₁₂H₂₂O₁₁, by mass and has a density of 1.06 g/mL.
 - (a) How many grams of sugar are in 1.0 L of this syrup?
 - (b) What is the molarity of this solution?
 - (c) What is the molality of this solution?
- *46. A solution of 3.84 g C_4H_2N (empirical formula) in 250.0 g of benzene depresses the freezing point of benzene 0.614°C. What is the molecular formula for the compound?
- **47.** Hydrochloric acid (HCl) is sold as a concentrated aqueous solution (12.0 mol/L). If the density of the solution is 1.18 g/mL, determine the molality of the solution.
- *48. How many grams of KNO₃ are needed to make 450 mL of a solution that is to contain 5.5 mg/mL of potassium ion? Calculate the molarity of the solution.
- **49.** What mass of 5.50% solution can be prepared from 25.0 g KCl?
- **50.** Given a solution containing $16.10 \text{ g } \text{C}_2\text{H}_6\text{O}_2$ in 82.0 g H₂O that has a boiling point of 101.62°C , verify that the boiling point elevation constant K_f for water is 0.512°C kg H₂O/mole solute.

- 34. What is the (a) molality, (b) freezing point, and (c) boiling point of a solution containing 100.0 g of ethylene glycol (C₂H₆O₂) in 150.0 g of water?
 - What is the molar mass of a compound if 4.80 g of the compound dissolved in 22.0 g of H_2O give a solution that freezes at -2.50°C?
- **51.** Physiological saline (NaCl) solutions used in intravenous injections have a concentration of 0.90% NaCl (mass/volume).
 - (a) How many grams of NaCl are needed to prepare 500.0 mL of this solution?
 - (b) How much water must evaporate from this solution to give a solution that is 9.0% NaCl (mass/volume)?
 - A solution is made from 50.0 g KNO₃ and 175 g H₂O. How many grams of water must evaporate to give a saturated solution of KNO₃ in water at 20°C? (See Figure 14.4.)
- **53.** What volume of 70.0% rubbing alcohol can you prepare if you have only 150 mL of pure isopropyl alcohol on hand?
- **54.** At 20°C, an aqueous solution of HNO₃ that is 35.0% HNO₃ by mass has a density of 1.21 g/mL.
 - (a) How many grams of HNO₃ are present in 1.00 L of this solution?
 - (b) What volume of this solution will contain 500. g HNO₃?
- * **55.** What is the molarity of a nitric acid solution if the solution is 35.0% HNO₃ by mass and has a density of 1.21 g/mL?
- **56.** To what volume must a solution of $80.0 \text{ g H}_2\text{SO}_4$ in 500.0 mL of solution be diluted to give a 0.10 *M* solution?
- **57.** A 10.0-mL sample of 16 *M* HNO₃ is diluted to 500.0 mL. What is the molarity of the final solution?
- 58. (a) How many moles of hydrogen will be liberated from 200.0 mL of 3.00 M HCl reacting with an excess of magnesium? The equation is

 $Mg(s) + 2 HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$

- (b) How many liters of hydrogen gas (H₂) measured at 27°C and 720 torr will be obtained? (*Hint*: Use the ideal gas law.)
- **59.** Which will be more effective in neutralizing stomach acid, HCI: a tablet containing 1.20 g Mg(OH)₂ or a tablet containing 1.00g Al(OH)₃? Show evidence for your answer.
- **60.** Which would be more effective as an antifreeze in an automobile radiator? A solution containing
 - (a) 10 kg of methyl alcohol (CH₃OH) or 10 kg of ethyl alcohol (C₂H₅OH)?
 - (b) 10 *m* solution of methyl alcohol or 10 *m* solution of ethyl alcohol?

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- * 61. Automobile battery acid is 38% H₂SO₄ and has a density * 67. of 1.29 g/mL. Calculate the molality and the molarity of this solution.
- * 62. What is the (a) molality and (b) boiling point of an aque- * 68. ous sugar, $C_{12}H_{22}O_{11}$, solution that freezes at $-5.4^{\circ}C$?
- **63.** A solution of 6.20 g $C_2H_6O_2$ in water has a freezing point of -0.372°C. How many grams of H_2O are in the solution?
- 64. What (a) mass and (b) volume of ethylene glycol $(C_2H_6O_2, \text{density} = 1.11 \text{ g/mL})$ should be added to 12.0 L of water in an automobile radiator to protect it from freezing at $-20^{\circ}C$? (c) To what temperature Fahrenheit will the radiator be protected?
- * 65. If 150 mL of 0.055 *M* HNO₃ are needed to completely neutralize 1.48 g of an *impure* sample of sodium hydrogen carbonate (baking soda), what percent of the sample is baking soda?
- 66. (a) How much water must be added to concentrated sulfuric acid (H₂SO₄) (17.8 *M*) to prepare 8.4 L of 1.5 *M* sulfuric acid solution?
 - (b) How many moles of H₂SO₄ are in each milliliter of the original concentrate?
 - (c) How many moles are in each milliliter of the diluted solution?

- **67.** How would you prepare a 6.00 *M* HNO₃ solution if only 3.00 *M* and 12.0 *M* solutions of the acid are available for mixing?
- **68.** A 20.0-mL portion of an HBr solution of unknown strength is diluted to exactly 240 mL. If 100.0 mL of this diluted solution requires 88.4 mL of 0.37 *M* NaOH to achieve complete neutralization, what was the strength of the original HBr solution?
- **69.** When 80.5 mL of 0.642 *M* Ba(NO₃)₂ are mixed with 44.5 mL of 0.743 *M* KOH, a precipitate of Ba(OH)₂ forms. How many grams of Ba(OH)₂ do you expect?
- **70.** A 0.25 *M* solution of lithium carbonate (Li₂CO₃), a drug used to treat manic depression, is prepared.
 - (a) How many moles of Li_2CO_3 are present in 45.8 mL of the solution?
 - (b) How many grams of Li_2CO_3 are in 750 mL of the same solution?
 - (c) How many milliliters of the solution would be needed to supply 6.0 g of the solute?
 - (d) If the solution has a density of 1.22 g/mL, what is its mass percent?

Challenge Exercises

All exercises with blue numbers have answers in the appendix of the text.

71. When solutions of hydrochloric acid and sodium sulfite ***72.** react, a salt, water, and sulfur dioxide gas are produced. How many liters of sulfur dioxide gas at 775 torr and 22°C can be produced when 125 mL of 2.50 *M* hydrochloric acid react with 75.0 mL of 1.75 *M* sodium sulfite?

Consider a saturated solution at 20°C made from 5.549 moles of water and an unknown solute. You determine the mass of the container containing the solution to be 563 g. The mass of the empty container is 375 g. Identify the solute.

Answers to Practice Exercises

- 14.1 unsaturated
- **14.2** 10.0% Na₂SO₄ solution
- **14.3** 2.02 *M*
- 14.4 0.27 g NaCl
- **14.5** 0.84 L (840 mL)
- 14.6 75 mL NaOH solution

14.7 $5.00 \times 10^{-2} M$

- **14.8** 43 g
- **14.9** 1.387 m
- **14.10** freezing point = -2.58°C, boiling point = 100.71°C

PUTTING IT TOGETHER: Review for Chapters 12–14

Multiple Choice:

Choose the correct answer to each of the following.

- **1.** Which of these statements is *not* one of the principal assumptions of the kinetic-molecular theory for an ideal gas?
 - (a) All collisions of gaseous molecules are perfectly elastic.(b) A mole of any gas occupies 22.4 L at STP.
 - (c) Gas molecules have no attraction for one another.
 - (d) The average kinetic energy for molecules is the same
 - for all gases at the same temperature.
- **2.** Which of the following is not equal to 1.00 atm?

(a) 760. cm Hg (c) 760. mm	ı Hg
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(b)	29.9	in. Hg	(d)) 760.	torr
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3. If the pressure on 45 mL of gas is changed from 600. torr to 800. torr, the new volume will be

(a) 60 mL	(c) 0.045 L
(b) 34 mL	(d) 22.4 L

4. The volume of a gas is 300. mL at 740. torr and 25°C. If the pressure remains constant and the temperature is raised to 100.°C, the new volume will be (a) 240. mL (c) 376 mL

(~)	2.0.		(\mathbf{v})	0,0 1111
(b)	1.20	L	(d)	75.0 mL

- 5. The volume of a dry gas is 4.00 L at 15.0°C and 745 torr. What volume will the gas occupy at 40.0°C and 700. torr?
 (a) 4.63 L
 (b) 3.46 L
 (c) 3.92 L
 (d) 4.08 L
- 6. A sample of Cl₂ occupies 8.50 L at 80.0°C and 740. mm Hg. What volume will the Cl₂ occupy at STP?
 (a) 10.7 L
 (c) 11.3 L

(b) 6.75 L (d) 6.40 L

- 7. What volume will 8.00 g O₂ occupy at 45°C and 2.00 atm?
 (a) 0.462 L
 (b) 104 L
 (c) 9.62 L
 (d) 3.26 L
- 8. The density of NH₃ gas at STP is

 (a) 0.760 g/mL
 (b) 0.760 g/L
 (c) 1.32 g/mL
 (d) 1.32 g/L
- **9.** The ratio of the relative rate of effusion of methane (CH₄) to sulfur dioxide (SO₂) is (a) ${}^{64}\!\!\!/_{16}$ (c) ${}^{1}\!\!/_{4}$

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(b)	¹⁶ / ₆₄	(d)	$^{2}/_{1}$

- 10. Measured at 65°C and 500. torr, the mass of 3.21 L of a gas is 3.5 g. The molar mass of this gas is
 (a) 21 g/mole
 (b) 46 g/mole
 (c) 24 g/mole
 (d) 130 g/mole
- **11.** Box A contains O_2 (molar mass = 32.0) at a pressure of 200 torr. Box B, which is identical to box A in volume, contains twice as many molecules of CH₄ (molar mass = 16.0) as the molecules of O_2 in box A. The temperatures of the gases are identical. The pressure in box B is

(a) 100 torr	(c) 400 torr
(b) 200 torr	(d) 800 torr

- 12. A 300.-mL sample of oxygen (O₂) is collected over water at 23°C and 725 torr. If the vapor pressure of water at 23°C is 21.0 torr, the volume of dry O₂ at STP is
 (a) 256 mL
 (b) 351 mL
 (c) 341 mL
 (d) 264 mL
- 13. A tank containing 0.01 mol of neon and 0.04 mol of helium shows a pressure of 1 atm. What is the partial pressure of neon in the tank?
 (a) 0.8 atm
 (b) 0.01 atm
 (c) 0.2 atm
 (d) 0.5 atm
- 14. How many liters of NO₂ (at STP) can be produced from 25.0 g Cu reacting with concentrated nitric acid?
 Cu(s) + 4 HNO₃(aq) →

 $Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$

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15. How many liters of butane vapor are required to produce 2.0 L CO₂ at STP?

16. What volume of CO_2 (at STP) can be produced when 15.0 g C_2H_6 and 50.0 g O_2 are reacted?

 $2 C_2 H_6(g) + 7 O_2(g) \longrightarrow 4 CO_2(g) + 6 H_2 O(g)$

- 17. Which of these gases has the highest density at STP?
 (a) N₂O
 (b) NO₂
 (c) Cl₂
 (d) SO₂
- 18. What is the density of CO2 at 25°C and 0.954 atm?

 (a) 1.72 g/L
 (c) 0.985 g/L

 (b) 2.04 g/L
 (d) 1.52 g/L
- **19.** How many molecules are present in 0.025 mol of H_2 gas? (a) 1.5×10^{22} molecules
 - (b) 3.37×10^{23} molecules
 - (c) 2.40×10^{25} molecules
 - (d) 1.50×10^{22} molecules
- 20. 5.60 L of a gas at STP have a mass of 13.0 g. What is the molar mass of the gas?
 (a) 33.2 g/mol
 (c) 66.4 g/mol

(b) 26.0 g/mol (c) 60.4 g/mol(d) 52.0 g/mol

- **21.** The heat of fusion of water is (a) 4.184 J/g (c) 2.26 kJ/g (b) 335 J/g (d) 2.26 kJ/mol
- **22.** The heat of vaporization of water is (a) 4.184 J/g (c) 2.26 kJ/g (b) 335 J/g (d) 2.26 kJ/mol
- **23.** The specific heat of water is (a) 4.184 J/g°C (c) 2.26 kJ/g°C (b) 335 J/g°C (d) 18 J/g°C

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- **24.** The density of water at 4°C is (a) 1.0 g/mL (c) 18.0 g/mL (d) 14.7 lb/in.³ (b) 80 g/mL
- **25.** SO₂ can be properly classified as a(n)(a) basic anhydride (c) anhydrous salt (b) hydrate (d) acid anhydride
- **26.** When compared to H_2S , H_2Se , and H_2Te , water is found to have the highest boiling point because it (a) has the lowest molar mass
 - (b) is the smallest molecule

 - (c) has the highest bonding
 - (d) forms hydrogen bonds better than the others
- 27. In which of the following molecules will hydrogen bonding be important?

(b)
$$S-H$$
 (d) $H-C-O-C-H$
 $|H$ H H

28. Which of the following is an incorrect equation? (a) $H_2SO_4 + 2 NaOH \longrightarrow Na_2SO_4 + 2 H_2O$ (b) $C_2H_c + O_2 \longrightarrow 2CO_2 + 3H$

(b)
$$C_2 H_6 + O_2 \longrightarrow 2 CO_2 + 3 H_2$$

(c) $C_2 H_6 + O_2 \longrightarrow 2 CO_2 + 3 H_2$

(c)
$$2 H_2 O \xrightarrow{H_2 SO_4} 2 H_2 + O_2$$

(d) $Ca + 2 H_2O \longrightarrow H_2 + Ca(OH)_2$

29. Which of the following is an incorrect equation?

 $\begin{array}{ccc} (a) \ C \ + \ H_2O(g) \xrightarrow{1000^\circ C} CO(g) \ + \ H_2(g) \\ (b) \ CaO \ + \ H_2O \longrightarrow Ca(OH)_2 \end{array}$

- (c) $2 \text{ NO}_2 + \text{H}_2\text{O} \longrightarrow 2 \text{ HNO}_2$
- (d) $Cl_2 + H_2O \longrightarrow HCl + HOCl$
- **30.** How many kilojoules are required to change 85 g of water at 25°C to steam at 100.°C? (a) 219 kJ (c) 590 kJ
 - (b) 27 kJ (d) 192 kJ
- **31.** A chunk of 0°C ice, mass 145 g, is dropped into 75 g of water at 62°C. The heat of fusion of water is 335 J/g. The result, after thermal equilibrium is attained, will be (a) 87 g ice and 133 g liquid water, all at 0°C (b) 58 g ice and 162 g liquid water, all at 0°C (c) 220 g water at 7°C
 - (d) 220 g water at 17°C
- **32.** The formula for iron(II) sulfate heptahydrate is (a) $Fe_2SO_4 \cdot 7 H_2O$ (c) $FeSO_4 \cdot 7 H_2O$ (b) $Fe(SO_4)_2 \cdot 6 H_2O$ (d) $Fe_2(SO_4)_3 \cdot 7 H_2O$
- **33.** The process by which a solid changes directly to a vapor is called

(a)	vaporization	(c)	sublimation
(b)	evaporation	(d)	condensation
~ ~			

- **34.** Hydrogen bonding
 - (a) occurs only between water molecules
 - (b) is stronger than covalent bonding
 - (c) can occur between NH_3 and H_2O
 - (d) results from strong attractive forces in ionic compounds

- **35.** A liquid boils when
 - (a) the vapor pressure of the liquid equals the external pressure above the liquid
 - (b) the heat of vaporization exceeds the vapor pressure
 - (c) the vapor pressure equals 1 atm
 - (d) the normal boiling temperature is reached
- 36. Consider two beakers, one containing 50 mL of liquid A and the other 50 mL of liquid B. The boiling point of A is 90°C and that of B is 72°C. Which of these statements is correct?
 - (a) A will evaporate faster than B.
 - (b) B will evaporate faster than A.
 - (c) Both A and B evaporate at the same rate.
 - (d) Insufficient data to answer the question.
- **37.** 95.0 g of 0.0°C ice is added to exactly 100. g of water at 60.0°C. When the temperature of the mixture first reaches 0.0°C, the mass of ice still present is
 - (a) 0.0 g (c) 10.0 g (b) 20.0 g (d) 75.0 g
- 38. Which of the following is not a general property of solutions?
 - (a) a homogeneous mixture of two or more substances
 - (b) variable composition
 - (c) dissolved solute breaks down to individual molecules (d) the same chemical composition, the same chemical
 - properties, and the same physical properties in every part
- **39.** If NaCl is soluble in water to the extent of 36.0 g NaCl/100 g H₂O at 20°C, then a solution at 20°C containing 45 g NaCl/150 g H_2O would be

(a)	dilute	(c)	supersaturated
(b)	saturated	(d)	unsaturated

40. If 5.00 g NaCl are dissolved in 25.0 g of water, the percent of NaCl by mass is

(a) 16.7%	(c) 0.20%
(b) 20.0%	(d) no correct answer

41. How many grams of 9.0% AgNO₃ solution will contain 5.3 g AgNO₃?

given

- (a) 47.7 g (c) 59 g (b) 0.58 g (d) no correct answer given
- 42. The molarity of a solution containing 2.5 mol of acetic acid (HC₂H₃O₂) in 400. mL of solution is (a) 0.063 M (c) 0.103 M
 - (b) 1.0 M (d) 6.3 M
- 43. What volume of 0.300 M KCl will contain 15.3 g KCl? (a) 1.46 L (c) 61.5 mL (b) 683 mL (d) 4.60 L
- 44. What mass of $BaCl_2$ will be required to prepare 200. mL of 0.150 *M* solution?

(a) 0.750 g	(c) 6.25 g
(b) 156 g	(d) 31.2 g

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PUTTING IT TOGETHER

Problems 45-47 relate to the reaction

 $CaCO_3 + 2 HCl \longrightarrow CaCl_2 + H_2O + CO_2$

45. What volume of 6.0 *M* HCl will be needed to react with 0.350 mol of CaCO₃?
(a) 42.0 mL
(c) 117 mL

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(\mathbf{h})	1 17	Ĭ.	(d)	583 mI

- 46. If 400. mL of 2.0 *M* HCl react with excess CaCO₃, the volume of CO₂ produced, measured at STP, is
 (a) 18 L
 (b) 5 6 L
 - (b) 5.6 L (d) 56 L
- **47.** If 5.3 g CaCl₂ are produced in the reaction, what is the molarity of the HCl used if 25 mL of it reacted with excess CaCO₃?

(a) 3.8 <i>M</i>	(c) 0.38 M
(b) 0.19 <i>M</i>	(d) 0.42 M

48. If 20.0 g of the nonelectrolyte urea $(CO(NH_2)_2)$ is dissolved in 25.0 g of water, the freezing point of the solution will be

(a) −2.47°C	(c) −24.7°C
(b) −1.40°C	(d) −3.72°C

49. When 256 g of a nonvolatile, nonelectrolyte unknown were dissolved in 500. g H_2O , the freezing point was found to be $-2.79^{\circ}C$. The molar mass of the unknown solute is

(a) 357	(c) 768
(b) 62.0	(d) 341

- 50. How many milliliters of 6.0 M H₂SO₄ must you use to prepare 500. mL of 0.20 M sulfuric acid solution?
 (a) 30
 (b) 17
 (c) 12
 (d) 100
- 51. How many milliliters of water must be added to 200. mL of 1.40 M HCl to make a solution that is 0.500 M HCl?
 (a) 360. mL
 (c) 140. mL
 - (b) 560. mL (d) 280. mL
- **52.** Which procedure is most likely to increase the solubility of most solids in liquids?

(a) stirring

- (b) pulverizing the solid
- (c) heating the solution
- (d) increasing the pressure
- **53.** The addition of a crystal of $NaClO_3$ to a solution of $NaClO_3$ causes additional crystals to precipitate. The original solution was

(a) unsaturated	(c) saturated
(b) dilute	(d) supersaturated

54. Which of these anions will not form a precipitate with silver ions, Ag⁺?

(a) Cl ⁻	(c) Br ⁻
(b) NO ₃	(d) CO_3^2

55. Which of these salts are considered to be soluble in water?

(a) BaSO ₄	(c) AgI
(b) NH ₄ Cl	(d) PbS

- (a) The solution contains 40 mL of alcohol in 100 mL of solution.
- (b) The solution contains 60 mL of benzene in 100 mL of solution.
- (c) The solution contains 40 mL of alcohol in 100 g of solution.
- (d) The solution is made by dissolving 40 mL of alcohol in 60 mL of benzene.
- **57.** Which of the following is not a colligative property?
 - (a) boiling point elevation
 - (b) freezing point depression
 - (c) osmotic pressure
 - (d) surface tension
- 58. When a solute is dissolved in a solvent
 - (a) the freezing point of the solution increases.
 - (b) the vapor pressure of the solution increases.
 - (c) the boiling point of the solution increases.
 - (d) the concentration of the solvent increases.
- **59.** Which of the following solutions will have the lowest freezing point where X is any element or nonelectrolytic compound?
 - (a) 1.0 mol X in 1 kg H_2O
 - (b) 2.0 mol X in 1 kg H_2O
 - (c) 1.2 mol X in 1 kg H_2O
 - (d) 0.80 mol X in 1 kg H_2O
- **60.** In the process of osmosis, water passes through a semipermeable membrane
 - (a) from a more concentrated solution to a dilute solution
 - (b) from a dilute solution to a more concentrated solution
 - (c) in order to remove a solute from a solution
 - (d) so that a sugar solution can become sweeter

Free Response Questions

Answer each of the following. Be sure to include your work and explanations in a clear, logical form.

- **1.** Which solution should have a higher boiling point: 215 mL of a 10.0% (m/v) aqueous KCl solution or 224 mL of a 1.10 *M* aqueous NaCl solution?
- A glass containing 345 mL of a soft drink (a carbonated beverage) was left sitting out on a kitchen counter. If the CO₂ released at room temperature (25°C) and pressure (1 atm) occupies 1.40 L, at a minimum, what is the concentration (in ppm) of the CO₂ in the original soft drink (assume the density of the original soft drink is 0.965 g/mL).
- **3.** Dina and Murphy were trying to react 100. mL of a 0.10 *M* HCl solution with KOH. The procedure called for a 10% KOH solution. Dina made a 10% mass/volume solution while Murphy made a 10% by mass solution. (Assume there is no volume change upon dissolving KOH.) Which solution required less volume to fully react with 100. mL of the HCl solution?

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4. A flask containing 825 mL of a solution containing 0.355 mole of CuSO₄ was left open overnight. The next morning the flask only contained 755 mL of solution.



- 5. Three students at Jamston High—Zack, Gaye, and Lamont—each had the opportunity to travel over spring break. As part of a project, each of them measured the boiling point of pure water at their vacation spot. Zack found a boiling point of 93.9°C, Gaye measured 101.1°C, and Lamont read 100.°C. Which student most likely went snow skiing near Ely, Nevada, and which student most likely went water skiing in Honolulu? From the boiling point information, what can you surmise about the Dead Sea region, the location of the third student's vacation? Explain.
- **6.** Why does a change in pressure of a gas significantly affect its volume whereas a change in pressure on a solid or liquid has negligible effect on their respective volumes? If the accompanying picture represents a liquid at the molecular level, draw what you might expect a solid and a gas to look like.



- **PUTTING IT TOGETHER**
- (a) What is the concentration (molarity) of the CuSO₄ solution remaining in the flask?
- (b) Which of the pathways shown below best represents the evaporation of water and why are the others wrong?

$$H-H(g) + O \equiv O(g)$$

 $H^+(g) + OH^-(g)$

Η

H(g)

H(g)

Η

- (b) If you filled up three balloons with equal masses of nitrogen, oxygen, and neon, all to the same volume at the same temperature, which would have the lowest pressure?
- 8. Ray ran a double-displacement reaction using 0.050 mol $CuCl_2$ and 0.10 mol AgNO₃. The resulting white precipitate was removed by filtration. The filtrate was accidentally left open on the lab bench for over a week, and when Ray returned the flask contained solid, blue crystals. Ray weighed the crystals and found they had a mass of 14.775 g. Was Ray expecting this number? If not, what did he expect?
- **9.** Why is it often advantageous or even necessary to run reactions in solution rather than mixing two solids? Would you expect reactions run in the gas phase to be more similar to solutions or to solids? Why?
- **10.** A solution of 5.36 g of a molecular compound dissolved in 76.8 g benzene (C_6H_6) and has a boiling point of 81.48°C. What is the molar mass of the compound?

Boiling point for benzene = 80.1° C K_b for benzene = 2.53° C kg solvent/mol solute