CHAPTER 14 Mixtures and Solutions

BIG Idea Nearly all of the gases, liquids, and solids that make up our world are mixtures.

14.1 Types of Mixtures

MAIN Idea Mixtures can be either heterogeneous or homogeneous.

14.2 Solution Concentration

MAIN Idea Concentration can be expressed in terms of percent or in terms of moles.

14.3 Factors Affecting Solvation

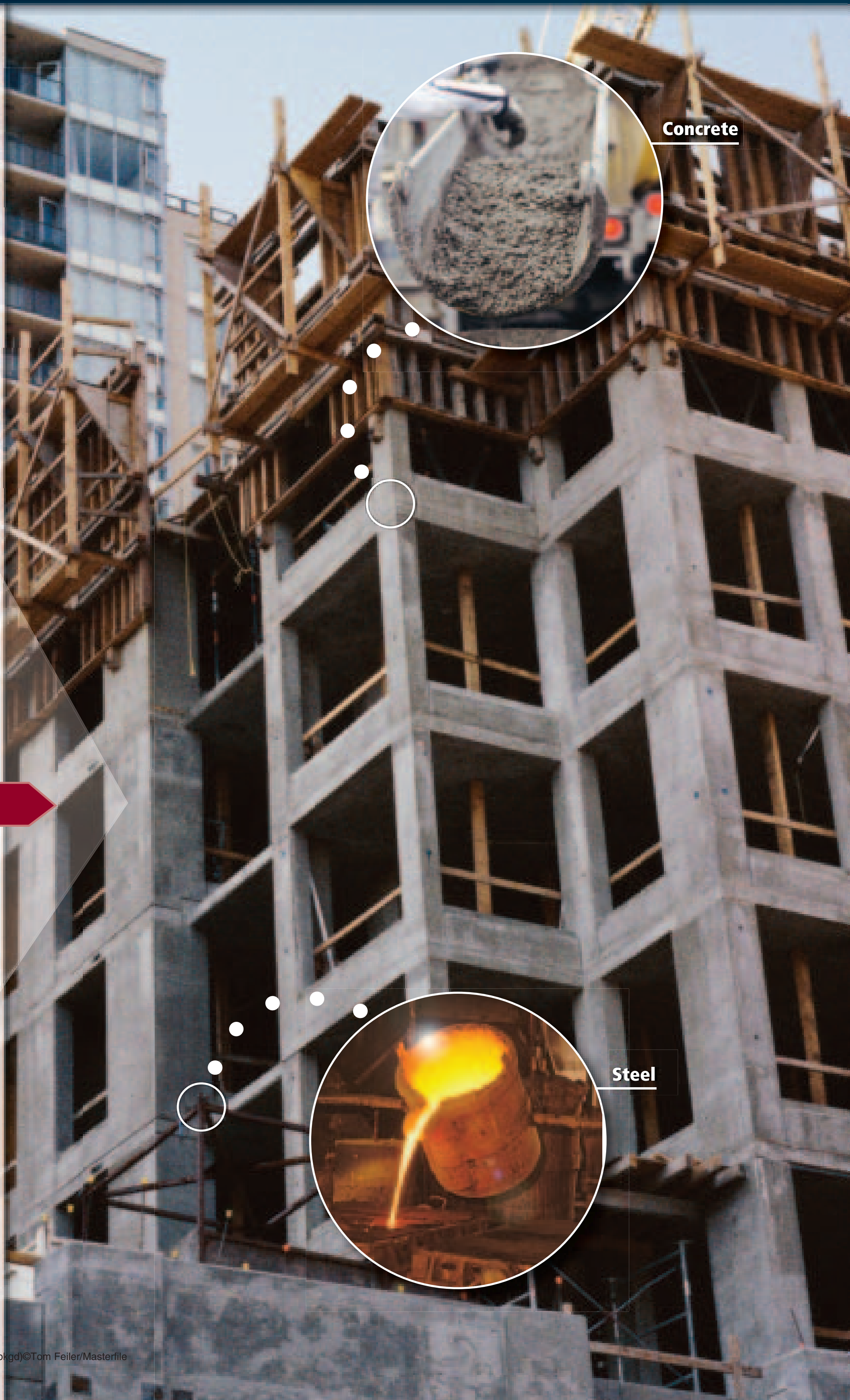
MAIN Idea Factors such as temperature, pressure, and polarity affect the formation of solutions.

14.4 Colligative Properties of Solutions

MAIN Idea Colligative properties depend on the number of solute particles in a solution.

ChemFacts

- Approximately 42.3% of the steel produced each year comes from recycled material.
- The major component of steel is iron, but other elements such as nickel, manganese, chromium, vanadium, and tungsten might be added depending on the desired use.
- Cement is used to make concrete and mortar to form building materials that are strong and can withstand normal environmental effects.
- About 6 billion cubic meters of concrete—the equivalent of 1 cubic meter per person—is produced each year.



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Start-Up Activities

LAUNCH Lab

How does energy change when solutions form?

When a solution is formed, there is an energy change that results from the interaction of two forces—the intermolecular forces among dissolving particles and the attractive forces between solute and solvent particles. How can this change be observed?

Procedure

1. Read and complete the lab safety form.
2. Measure 10 g of **ammonium chloride (NH_4Cl)** using a **balance,** and place it in a **100-mL beaker.**
3. Measure 30 mL of **water** with a **50-mL graduated cylinder,** and add to the NH_4Cl, stirring with your **stirring rod.**
4. Feel the bottom of the beaker, and record your observations.
5. Repeat Steps 2–4 using **calcium chloride ($CaCl_2$).**
6. Dispose of the solutions by flushing them down a drain with water.

Analysis

1. **Compare** Which dissolving process is exothermic, and which is endothermic?
2. **Infer** What are everyday applications for dissolving processes that are exothermic? Endothermic?

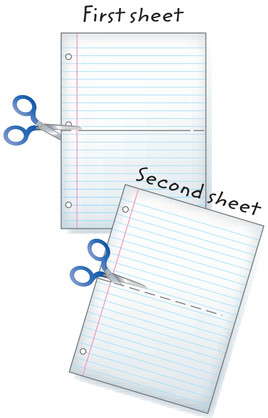
Inquiry If you wanted a greater temperature change, would you add more solute or more solvent? Explain.

FOLDABLES™ Study Organizer

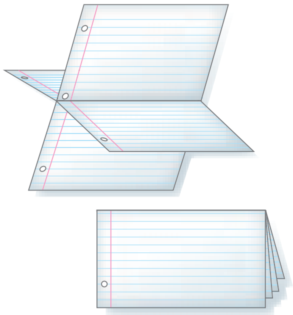
Concentration Make the following Foldable to help you organize information about the concentration of solutions.

STEP 1 Fold two sheets of notebook paper in half horizontally.

STEP 2 On the first sheet, make a 3-cm cut on the fold line on each side of the paper. On the second sheet, cut along the fold line to within 3 cm of each edge.



STEP 3 Slip the first sheet through the cut in the second sheet to make a four-page book.



FOLDABLES **Use this Foldable with Section 14.2.** As you read this section, use your book to record what you learn about how the concentrations of solutions are expressed. Include sample calculations.

Chemistry Online

Visit glencoe.com to:

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- find the Try at Home Lab, Identifying Colloids

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Section 14.1

Objectives

- **Compare** the properties of suspensions, colloids, and solutions.
- **Identify** types of colloids and types of solutions.
- **Describe** the electrostatic forces in colloids.

Review Vocabulary

solute: a substance dissolved in a solution

New Vocabulary

suspension
colloid
Brownian motion
Tyndall effect
soluble
miscible
insoluble
immiscible

Types of Mixtures

MAIN Idea **Mixtures can be either heterogeneous or homogeneous.**

Real-World Reading Link If you have ever filled a pail with ocean water, you might have noticed that some of the sediment settles to the bottom of the pail. However, the water will be salty no matter how long you let the pail sit. Why do some substances settle out but others do not?

Heterogeneous Mixtures

Recall from Chapter 3 that a mixture is a combination of two or more pure substances in which each pure substance retains its individual chemical properties. Heterogeneous mixtures do not blend smoothly throughout, and the individual substances remain distinct. Two types of heterogeneous mixtures are suspensions and colloids.

Suspensions A **suspension** is a mixture containing particles that settle out if left undisturbed. The muddy water shown in **Figure 14.1** is a suspension. Pouring a liquid suspension through a filter will also separate out the suspended particles. Some suspensions will separate into two distinct layers if left undisturbed for awhile—a solidlike substance on the bottom and water on the top. However, when stirred, the solidlike substance quickly begins flowing like a liquid.

Substances that behave in this way are called thixotropic (thik suh TROH pik). Some clays are thixotropic suspensions, of particular significance in the construction of buildings in earthquake zones. These clays can form liquids in response to the agitation of an earthquake, which can result in the collapse of structures built on the clay.

Figure 14.1 A suspension can be separated by allowing it to sit for a period of time. A liquid suspension can also be separated by pouring it through a filter.

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Table 14.1 Types of Colloids

Category	Example	Dispersed Particles	Dispersing Medium
Solid sol	colored gems	solid	solid
Sol	blood, gelatin	solid	liquid
Solid emulsion	butter, cheese	liquid	solid
Emulsion	milk, mayonnaise	liquid	liquid
Solid foam	marshmallow, soaps that float	gas	solid
Foam	whipped cream, beaten egg white	gas	liquid
Solid aerosol	smoke, dust in air	solid	gas
Liquid aerosol	spray deodorant, fog, clouds	liquid	gas

Colloids Particles in a suspension are much larger than atoms and can settle out of solution. A heterogeneous mixture of intermediate-sized particles (between atomic-scale size of solution particles and the size of suspension particles) is a **colloid.** Colloid particles are between 1 nm and 1000 nm in diameter and do not settle out. Milk is a colloid. The components of homogenized milk cannot be separated by settling or by filtration.

The most abundant substance in the mixture is the dispersion medium. Colloids are categorized according to the phases of their dispersed particles and dispersing mediums. Milk is a colloidal emulsion because liquid particles are dispersed in a liquid medium. Other types of colloids are described in **Table 14.1.**

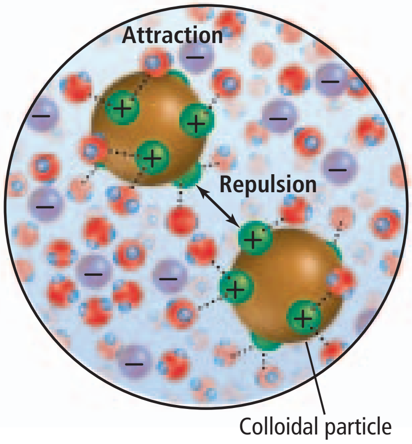
The dispersed particles in a colloid are prevented from settling out because they often have polar or charged atomic groups on their surfaces. These areas on their surfaces attract the positively or negatively charged areas of the dispersing-medium particles. This results in the formation of electrostatic layers around the particles, as shown in **Figure 14.2.** The layers repel each other when the dispersed particles collide; thus, the particles remain in the colloid.

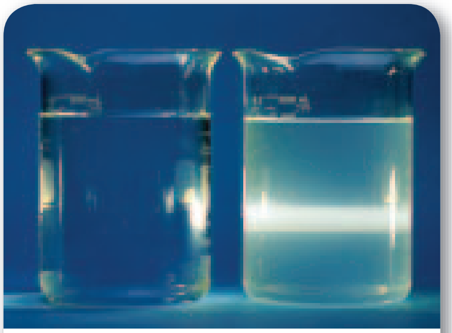
If you interfere with the electrostatic layering, colloid particles will settle out of the mixture. For example, if you stir an electrolyte into a colloid, the dispersed particles clump together, destroying the colloid. Heating also destroys a colloid because it gives colliding particles enough kinetic energy to overcome the electrostatic forces and settle out.

Brownian motion The dispersed particles of liquid colloids make jerky, random movements. This erratic movement of colloid particles is called **Brownian motion.** It was first observed by, and later named for, the Scottish botanist Robert Brown (1773–1858), who noticed the random movements of pollen grains dispersed in water. Brownian motion results from collisions of particles of the dispersion medium with the dispersed particles. These collisions help to prevent the colloid particles from settling out of the mixture.

Reading Check **Describe** two reasons why particles in a colloid do not settle out.

■ **Figure 14.2** The dispersing medium particles form charged layers around the colloid particles. These charged layers repel each other and keep the particles from settling out.





■ **Figure 14.3** Particles in a colloid scatter light, unlike particles in a solution. Called the Tyndall effect, the beam of light is visible in the colloid because of light scattering.

Determine *which mixture is the colloid.*

Tyndall effect Concentrated colloids are often cloudy or opaque. Dilute colloids sometimes appear as clear as solutions. Dilute colloids appear to be homogeneous solutions because their dispersed particles are so small. However, dispersed colloid particles scatter light, a phenomenon known as the **Tyndall effect.** In **Figure 14.3,** a beam of light is shone through two unknown mixtures. You can observe that dispersed colloid particles scatter the light, unlike particles in the solution. Suspensions also exhibit the Tyndall effect, but solutions never exhibit the Tyndall effect. You have observed the Tyndall effect if you have observed rays of sunlight passing through smoke-filled air, or viewed lights through fog. The Tyndall effect can be used to determine the amount of colloid particles in suspension.

Homogeneous Mixtures

Cell solutions, ocean water, and steel might appear dissimilar, but they share certain characteristics. In Chapter 3, you learned that solutions are homogeneous mixtures that contain two or more substances called the solute and the solvent. The solute is the substance that dissolves. The solvent is the dissolving medium. When you look at a solution, it is not possible to distinguish the solute from the solvent.

Types of solutions A solution might exist as a gas, a liquid, or a solid, depending on the state of its solvent, as shown in **Table 14.2.** Air is a gaseous solution, and its solvent is nitrogen gas. Braces that you wear on your teeth might be made of nitinol, a solid solution of titanium in nickel. Most solutions, however, are liquids. You read in Chapter 9 that reactions can take place in aqueous solutions, or solutions in which the solvent is water. Water is the most common solvent among liquid solutions.

DATA ANALYSIS LAB

Based on Real Data*

Design an Experiment

How can you measure turbidity? The National Primary Drinking Water Regulations set the standards for public water systems. Turbidity—a measure of the cloudiness of water that results from the suspension of solids in the water—is often associated contamination from viruses, parasites, and bacteria. Most of these colloid particles come from erosion, industrial and human waste, algae blooms from fertilizers, and decaying organic matter.

Data and Observation

The Tyndall effect can be used to measure the turbidity of water. Your goal is to plan a procedure and develop a scale to interpret data.

Think Critically

1. **Identify** the variables that can be used to relate the ability of light to pass through the liquid and the number of the colloid particles present. What will you use as a control?
2. **Relate** the variables used in the experiment to the actual number of colloid particles that are present.
3. **Analyze** What safety precautions must be considered?
4. **Determine** the materials you need to measure the Tyndall effect. Select technology to collect or interpret data.

*Data obtained from U.S. Environmental Protection Agency. 2006. *The Office of Groundwater and Drinking Water.*

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Concepts In Motion

Interactive Table Explore the gas laws at glencoe.com.

Table 14.2 Types and Examples of Solutions

Type of Solution	Example	Solvent	Solute
Gas	air	nitrogen (gas)	oxygen (gas)
Liquid	carbonated water	water (liquid)	carbon dioxide (gas)
	ocean water	water (liquid)	oxygen gas (gas)
	antifreeze	water (liquid)	ethylene glycol (liquid)
	vinegar	water (liquid)	acetic acid (liquid)
	ocean water	water (liquid)	sodium chloride (solid)
Solid	dental amalgam	silver (solid)	mercury (liquid)
	steel	iron (solid)	carbon (solid)

Just as solutions can exist in different forms, the solutes in the solutions can be gases, liquids, or solids, also shown in **Table 14.2.** Solutions, such as ocean water, can contain more than one solute.

Forming solutions Some combinations of substances readily form solutions, and others do not. A substance that dissolves in a solvent is said to be **soluble** in that solvent. For example, sugar is soluble in water—a fact you might have learned by dissolving sugar in flavored water to make a sweetened beverage, such as tea or lemonade. Two liquids that are soluble in each other in any proportion, such as those that form the antifreeze listed in **Table 14.2,** are said to be **miscible.** A substance that does not dissolve in a solvent is said to be **insoluble** in that solvent. Sand is insoluble in water. The liquids in a bottle of oil and vinegar separate shortly after they are mixed. Oil is insoluble in vinegar. Two liquids that can be mixed together but separate shortly after are said to be **immiscible.**

Section 14.1 Assessment

Section Summary

- The individual substances in a heterogeneous mixture remain distinct.
- Two types of heterogeneous mixtures are suspensions and colloids.
- Brownian motion is the erratic movement of colloid particles.
- Colloids exhibit the Tyndall effect.
- A solution can exist as a gas, a liquid, or a solid, depending on the solvent.
- Solutes in a solution can be gases, liquids, or solids.

1. **MAIN Idea** **Explain** Use the properties of seawater to describe the characteristics of mixtures.
2. **Distinguish** between suspensions and colloids.
3. **Identify** the various types of solutions. Describe the characteristics of each type of solution.
4. **Explain** Use the Tyndall effect to explain why it is more difficult to drive through fog using high beams than using low beams.
5. **Describe** different types of colloids.
6. **Explain** Why do dispersed colloid particles stay dispersed?
7. **Summarize** What causes Brownian motion?
8. **Compare and Contrast** Make a table that compares the properties of suspensions, colloids, and solutions.

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Section 14.2

Solution Concentration

Objectives

- **Describe** concentration using different units.
- **Determine** the concentrations of solutions.
- **Calculate** the molarity of a solution.

Review Vocabulary

solvent: the substance that dissolves a solute to form a solution

New Vocabulary

concentration
molarity
molality
mole fraction

MAIN Idea **Concentration can be expressed in terms of percent or in terms of moles.**

Real-World Reading Link Have you ever tasted a glass of iced tea and found it too strong or too bitter? To adjust the taste, you could add sugar to sweeten the tea, or you could add water to dilute it. Either way, you are changing the concentration of the particles dissolved in the water.

Expressing Concentration

The **concentration** of a solution is a measure of how much solute is dissolved in a specific amount of solvent or solution. Concentration can be described qualitatively using the words *concentrated* or *dilute*. Notice the pots of tea in **Figure 14.4.** One of the tea solutions is more concentrated than the other. In general, a concentrated solution contains a large amount of solute. The darker tea has more tea particles than the lighter tea. Conversely, a dilute solution contains a small amount of solute. The lighter tea in **Figure 14.4** is dilute and contains less tea particles than the darker tea.

Although qualitative descriptions of concentration can be useful, solutions are more often described quantitatively. Some commonly used quantitative descriptions are percent by mass, percent by volume, molarity, and molality. These descriptions express concentration as a ratio of measured amounts of solute and solvent or solution. **Table 14.3** lists each ratio's description.

Which qualitative description should be used? The description used depends on the type of solution analyzed and the reason for describing it. For example, a chemist working with a reaction in an aqueous solution most likely refers to the molarity of the solution, because he or she need to know the number of particles involved in the reaction.

Figure 14.4 The strength of the tea corresponds to its concentration. The darker pot of tea is more concentrated than the lighter pot.

Table 14.3 Concentration Ratios

Concentration Description	Ratio
Percent by mass	$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Percent by volume	$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$
Molarity	$\frac{\text{moles of solute}}{\text{liter of solution}}$
Molality	$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$
Mole fraction	$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$

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Percent by mass The percent by mass is the ratio of the solute's mass to the solution's mass expressed as a percent. The mass of the solution equals the sum of the masses of the solute and the solvent.

Percent by Mass

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Percent by mass equals the mass of the solute divided by the mass of the whole solution, multiplied by 100.

FOLDABLES Incorporate information from this section into your Foldable.

EXAMPLE Problem 14.1

Math Handbook
Percents
pages 964–965

Calculate Percent by Mass In order to maintain a sodium chloride (NaCl) concentration similar to ocean water, an aquarium must contain 3.6 g NaCl per 100.0 g of water. What is the percent by mass of NaCl in the solution?

1 Analyze the Problem

You are given the amount of sodium chloride dissolved in 100.0 g of water. The percent by mass of a solute is the ratio of the solute's mass to the solution's mass, which is the sum of the masses of the solute and the solvent.

Known	**Unknown**
mass of solute = 3.6 g NaCl	**percent by mass** = **?**
mass of solvent = 100.0 g H_2O	

2 Solve for the Unknown

Find the mass of the solution.

mass of solution = grams of solute + grams of solvent

mass of solution = 3.6 g + 100.0 g = 103.6 g **Substitute mass of solute = 3.6 g, and mass of solvent = 100.0 g.**

Calculate the percent by mass.

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$ **State the equation for percent by mass.**

$$\text{percent by mass} = \frac{3.6\text{ g}}{103.6\text{ g}} \times 100 = 3.5\%$$ **Substitute mass of solute = 3.6 g, and mass of solution = 103.6 g.**

3 Evaluate the Answer

Because only a small mass of sodium chloride is dissolved per 100.0 g of water, the percent by mass should be a small value, which it is. The mass of sodium chloride was given with two significant figures; therefore, the answer is also expressed with two significant figures.

PRACTICE Problems

Extra Practice Pages 985–986 and glencoe.com

9. What is the percent by mass of $NaHCO_3$ in a solution containing 20.0 g of $NaHCO_3$ dissolved in 600.0 mL of H_2O?

10. You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite (NaOCl) is 3.62%. How many grams of NaOCl are in the solution?

11. In Question 10, how many grams of solvent are in the solution?

12. **Challenge** The percent by mass of calcium chloride in a solution is found to be 2.65%. If 50.0 g of calcium chloride is used, what is the mass of the solution?

■ **Figure 14.5** B20 is 20% by volume biodiesel and 80% by volume petroleum diesel. Biodiesel is a alternative fuel that can be produced from renewable resources, such as vegetable oil.

Percent by volume Percent by volume usually describes solutions in which both solute and solvent are liquids. The percent by volume is the ratio of the volume of the solute to the volume of the solution, expressed as a percent. The volume of the solution is the sum of the volumes of the solute and the solvent. Calculations of percent by volume are similar to those involving percent by mass.

Percent by Volume

$$\text{percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Percent by volume equals the volume of solute divided by the volume of the solution, multiplied by 100.

Biodiesel, shown in **Figure 14.5,** is a clean-burning alternative fuel that is produced from renewable resources. Biodiesel can be used in diesel engines with little or no modifications. Biodiesel is simple to use, biodegradable, nontoxic, and it does not contain sulfur or aromatics. It does not contain petroleum, but it can be blended with petroleum diesel to create a biodiesel blend. B20 is 20% by volume biodiesel, 80% by volume petroleum diesel.

✓ **Reading Check** **Compare** percent mass and percent volume.

PRACTICE Problems **Extra Practice** Pages 985–986 and glencoe.com

13. What is the percent by volume of ethanol in a solution that contains 35 mL of ethanol dissolved in 155 mL of water?

14. What is the percent by volume of isopropyl alcohol in a solution that contains 24 mL of isopropyl alcohol in 1.1 L of water?

15. **Challenge** If 18 mL of methanol is used to make an aqueous solution that is 15% methanol by volume, how many milliliters of solution is produced?

Molarity Percent by volume and percent by mass are only two of the commonly used ways to quantitatively describe the concentrations of liquid solutions. One of the most common units of solution concentration is molarity. **Molarity** (M) is the number of moles of solute dissolved per liter of solution. Molarity is also known as molar concentration, and the unit M is read as molar. A liter of solution containing 1 mol of solute is a $1M$ solution, which is read as a one-molar solution. A liter of solution containing 0.1 mol of solute is a $0.1M$ solution. To calculate a solution's molarity, you must know the volume of the solution in liters and the amount of dissolved solute in moles.

Molarity

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

The molarity of a solution equals the moles of solute divided by the liters of solution.

✓ **Reading Check** **Determine** What is the molar concentration of a liter solution with 0.5 mol of solute?

AP Photo/L.G. Patterson

EXAMPLE Problem 14.2

Calculating Molarity A 100.5-mL intravenous (IV) solution contains 5.10 g of glucose ($C_6H_{12}O_6$). What is the molarity of this solution? The molar mass of glucose is 180.16 g/mol.

1 Analyze the Problem

You are given the mass of glucose dissolved in a volume of water. The molarity of the solution is the ratio of moles of solute per liter of solution.

Known

mass of solute = 5.10 g $C_6H_{12}O_6$

molar mass of $C_6H_{12}O_6$ = 180.16 g/mol

volume of solution = 100.5 mL

Math Handbook

Solving Algebraic Equations
page 954

Unknown

solution concentration = ? *M*

2 Solve for the Unknown

Calculate the number of moles of $C_6H_{12}O_6$.

$$(5.10\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \text{mol } C_6H_{12}O_6}{180.16\ \cancel{\text{g } C_6H_{12}O_6}}\right)$$

$$= 0.0283\ \text{mol } C_6H_{12}O_6$$

Multiply grams of $C_6H_{12}O_6$ by the molar mass of $C_6H_{12}O_6$.

Convert the volume of H_2O to liters.

$$(100.5\ \cancel{\text{mL}}\ \text{solution})\left(\frac{1\ \text{L}}{1000\ \cancel{\text{mL}}}\right) = 0.1005\ \text{L solution}$$

Use the conversion factor 1 L/1000 mL.

Solve for the molarity.

$$M = \frac{\text{moles of solute}}{\text{liters of solutions}}$$

State the molarity equation.

$$M = \left(\frac{0.0283\ \text{mol } C_6H_{12}O_6}{0.1005\ \text{L solution}}\right)$$

Substitute moles of $C_6H_{12}O_6$ = 0.0283 and volume of solution = liters of solution = 0.1005 L.

$$M = \left(\frac{0.0282\ \text{mol } C_6H_{12}O_6}{1\ \text{L solution}}\right) = 0.282M$$

Divide numbers and units.

3 Evaluate the Answer

The molarity value will be small because only a small mass of glucose was dissolved in the solution. The mass of glucose used in the problem has three significant figures; therefore, the value of the molarity also has three significant figures.

PRACTICE Problems

Extra Practice Pages 985–986 and glencoe.com

16. What is the molarity of an aqueous solution containing 40.0 g of glucose ($C_6H_{12}O_6$) in 1.5 L of solution?

17. Calculate the molarity of 1.60 L of a solution containing 1.55 g of dissolved KBr.

18. What is the molarity of a bleach solution containing 9.5 g of NaOCl per liter of bleach?

19. **Challenge** How much calcium hydroxide ($Ca(OH)_2$), in grams, is needed to produce 1.5 L of a 0.25*M* solution?

CAREERS IN CHEMISTRY

Pharmacy Technician Most pharmacists rely on pharmacy technicians to prepare the proper medications to fill prescriptions. These technicians read patient charts and prescriptions in order to prepare the proper concentration, or dose, of medication that is to be administered to patients. For more information on chemistry careers, visit glencoe.com.

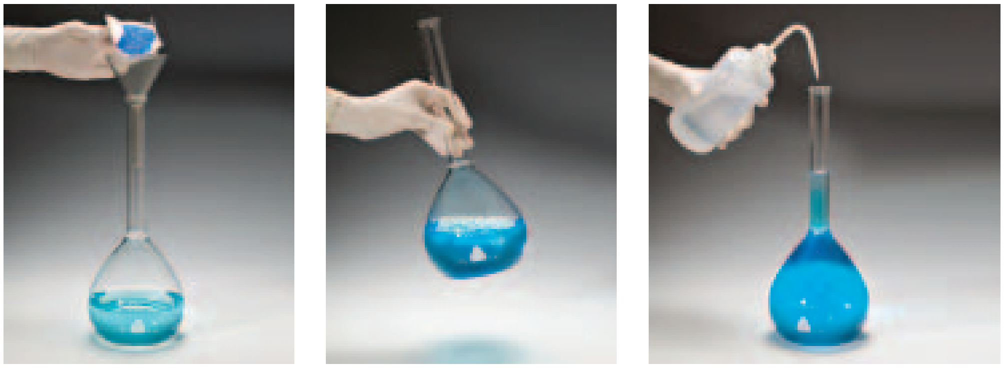


Figure 14.6 Accurately preparing a solution of copper sulfate involves several steps.
Explain *why you cannot add 375 g of copper sulfate directly to 1 L of water to make a 1.5M solution.*

Step 1: The mass of the solute is measured.

Step 2: The solute is placed in a volumetric flask of the correct volume.

Step 3: Distilled water is added to the flask to bring the solution level up to the calibration mark.

Preparing molar solutions Now that you know how to calculate the molarity of a solution, how do you think you would prepare 1 L of a 1.50*M* aqueous solution of copper (II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$)? A 1.50*M* aqueous solution of ($CuSO_4 \cdot 5H_2O$) contains 1.50 mol of $CuSO_4 \cdot 5H_2O$ dissolved in 1 L of solution. The molar mass of $CuSO_4 \cdot 5H_2O$ is about 249.70 g. Thus, 1.50 mol of $CuSO_4 \cdot 5H_2O$ has a mass of 375 g, an amount that you can measure on a balance.

$$\frac{1.50 \text{ mol } CuSO_4 \cdot 5H_2O}{1 \text{ L solution}} \times \frac{249.7 \text{ g } CuSO_4 \cdot 5H_2O}{1 \text{ mol } CuSO_4 \cdot 5H_2O} = \frac{375 \text{ g } CuSO_4 \cdot 5H_2O}{1 \text{ L solution}}$$

You cannot simply add 375 g of $CuSO_4 \cdot 5H_2O$ to 1 L of water to make the 1.50*M* solution. Like all substances, $CuSO_4 \cdot 5H_2O$ takes up space and will add volume to the solution. Therefore, you must use slightly less than 1 L of water to make 1 L of solution, as shown in **Figure 14.6.**

You will often do experiments that call for small quantities of solution. For example, you might need only 100 mL of a 1.50*M* $CuSO_4 \cdot 5H_2O$ solution for an experiment. Look again at the definition of molarity. As calculated above, a 1.50*M* solution of $CuSO_4 \cdot 5H_2O$ contains 1.50 mol of $CuSO_4 \cdot 5H_2O$ per 1 L of solution. Therefore, 1 L of solution contains 375 g of $CuSO_4 \cdot 5H_2O$.

This relationship can be used as a conversion factor to calculate how much solute you need for your experiment.

$$100 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{375 \text{ g } CuSO_4 \cdot 5H_2O}{1 \text{ L solution}} = 37.5 \text{ g } CuSO_4 \cdot 5H_2O$$

Thus, you would need to measure out 37.5 g of $CuSO_4 \cdot 5H_2O$ to make 100 mL of a 1.50*M* solution.

PRACTICE Problems

Extra Practice Pages 985–986 and glencoe.com

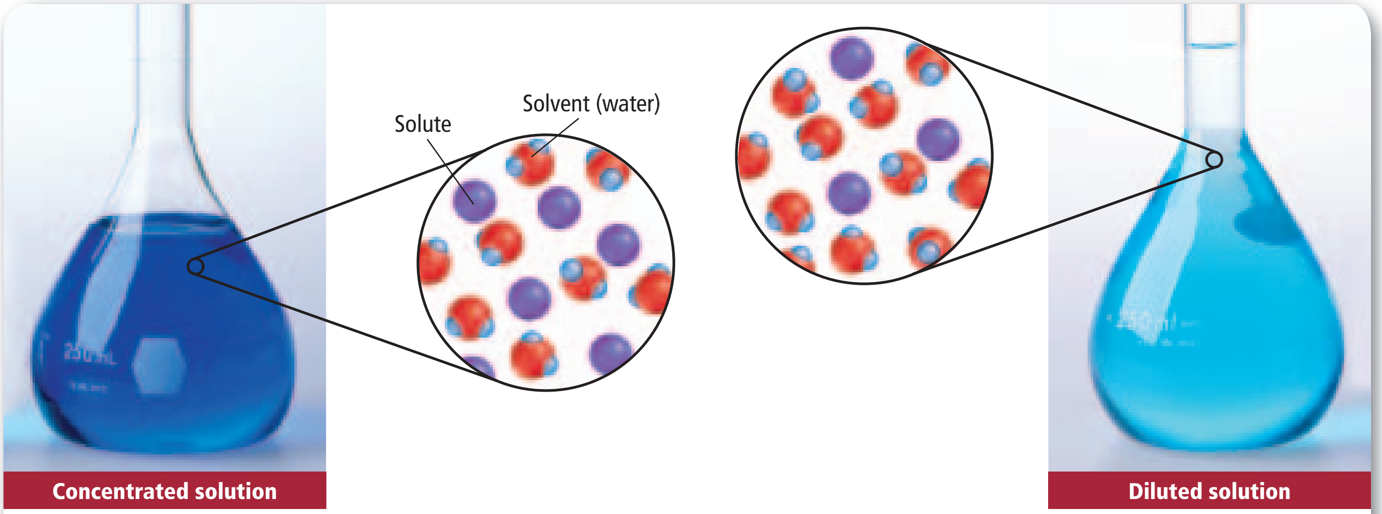
20. How many grams of $CaCl_2$ would be dissolved in 1.0 L of a 0.10*M* solution of $CaCl_2$?

21. How many grams of $CaCl_2$ should be dissolved in 500.0 mL of water to make a 0.20*M* solution of $CaCl_2$?

22. How much NaOH are in 250 mL of a 3.0*M* NaOH solution?

23. Challenge What volume of ethanol (C_2H_3OH) is in 100.0 mL of 0.15*M* solution? The density of ethanol is 0.7893 g/mL.

Matt Meadows



■ **Figure 14.7** A concentrated solution can be diluted by adding solvent. The number of moles of solute does not change when a concentrated solution is diluted.

Diluting molar solutions In the laboratory, you might use concentrated solutions of standard molarities, called stock solutions. For example, concentrated hydrochloric acid (HCl) is 12*M*. Recall that a concentrated solution has a large amount of solute. You can prepare a less-concentrated solution by diluting the stock solution with additional solvent. When you add solvent, you increase the number of solvent particles among which the solute particles move, as shown in **Figure 14.7,** thereby decreasing the solution's concentration.

How do you determine the volume of stock solution you must dilute? You can rearrange the expression of molarity to solve for moles of solute.

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{moles of solute} = \text{molarity} \times \text{liters of solution}$$

Because the total number of moles of solute does not change during dilution,

$$\text{moles of solute in the stock solution} = \text{moles of solute after dilution.}$$

Substituting moles of solute with molarity times liters of solution, the relationship can be expressed in the dilution equation.

Dilution Equation

$$M_1V_1 = M_2V_2$$

M **represents molarity.**
V **represents volume.**

For a given amount of solute, the product of the molarity and volume of the stock solution equals the product of the molarity and the volume of the dilute solution.

M_1 and V_1 represent the molarity and volume of the stock solution, and M_2 and V_2 represent the molarity and volume of the dilute solution. Before dilution, a concentrated solution contains a fairly high ratio of solute particles to solvent particles. After adding more solvent, the ratio of solute particles to solvent particles has decreased.

VOCABULARY

ACADEMIC VOCABULARY

Concentrated
less dilute or diffuse
We added more water to the lemonade because it was too concentrated.

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EXAMPLE Problem 14.3

Diluting Stock Solutions If you want to know the concentration and volume of the solution you want to prepare, you can calculate the volume of stock solution you will need. What volume, in milliliters, of 2.00M calcium chloride ($CaCl_2$) stock solution would you use to make 0.50 L of 0.300M calcium chloride solution?

1 Analyze the Problem

You are given the molarity of a stock solution of $CaCl_2$ and the volume and molarity of a dilute solution of $CaCl_2$. Use the relationship between molarities and volumes to find the volume, in liters, of the stock solution required. Then, convert the volume to milliliters.

Known	Unknown
$M_1 = 2.00M\ CaCl_2$	$V_1 = ?\ mL\ 2.00M\ CaCl_2$
$M_2 = 0.300M$	
$V_2 = 0.50\ L$	

Math Handbook

Solving Algebraic
Equations
page 954

2 Solve for the Unknown

Solve the molarity-volume relationship for the volume of the stock solution V_1.

$$M_1 V_1 = M_2 V_2$$

State the dilution equation.

$$V_1 = V_2\left(\frac{M_2}{M_1}\right)$$

Solve for V_1.

$$V_1 = (0.50\ L)\left(\frac{0.300M}{2.00M}\right)$$

Substitute $M_1 = 2.00M$, $M_2 = 0.300M$, and $V_2 = 0.50$ L.

$$V_1 = (0.50\ L)\left(\frac{0.300\cancel{M}}{2.00\cancel{M}}\right) = 0.075\ L$$

Multiply and divide numbers and units.

$$V_1 = (0.075\ L)\left(\frac{1000\ mL}{1\ L}\right) = 75\ mL$$

Convert to milliliters using the conversion factor 1000 mL/1 L.

To make the dilution, measure out 75 mL of the stock solution and dilute it with enough water to make the final volume 0.50 L.

3 Evaluate the Answer

The volume V_1 was calculated, and then its value was converted to milliliters. This volume should be less than the final volume of the dilute solution, and it is. Of the given information, V_2 had the fewest number of significant figures, with two. Thus, the volume V_1 should also have two significant figures, and it does.

PRACTICE Problems

Extra Practice Pages 985–986 and glencoe.com

24. What volume of a 3.00M KI stock solution would you use to make 0.300 L of a 1.25M KI solution?

25. How many milliliters of a 5.0M H_2SO_4 stock solution would you need to prepare 100.0 mL of 0.25M H_2SO_4?

26. **Challenge** If 0.5 L of 5M stock solution of HCl is diluted to make 2 L of solution, how much HCl, in grams, was in the solution?

Molality The volume of a solution changes with temperature as it expands or contracts. This change in volume alters the molarity of the solution. Masses, however, do not change with temperature. It is sometimes more useful to describe solutions in terms of how many moles of solute are dissolved in a specific mass of solvent. Such a description is called **molality—**the ratio of the number of moles of solute dissolved in 1 kg of solvent. The unit *m* is read as molal. A solution containing 1 mol of solute per kilogram of solvent is a one-molal solution.

Molality

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

The molality of a solution equals the number of moles of solute divided by kg of solvent.

Chemistry Online

Personal Tutor For an online tutorial on calculating molarity and molality, visit glencoe.com.

EXAMPLE Problem 14.4

Math Handbook
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page 954

Calculating Molality In the lab, a student adds 4.5 g of sodium chloride (NaCl) to 100.0 g of water. Calculate the molality of the solution.

1 Analyze the Problem

You are given the mass of solute and solvent. Determine the number of moles of solute. Then, you can calculate the molality.

Known	**Unknown**
mass of water (H_2O) = 100.0 g	m = ? mol/kg
mass of sodium chloride (NaCl) = 4.5 g	

2 Solve for the Unknown

$$4.5 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.077 \text{ mol NaCl}$$

Calculate the number of moles of solute.

$$100.0 \text{ g } H_2O \times \frac{1 \text{ kg } H_2}{1000 \text{ g } H_2O} = 0.1000 \text{ kg } H_2O$$

Convert the mass of H_2O from grams to kilograms using the factor 1 kg/1000 g.

Substitute the known values into the expression for molality, and solve.

$$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Write the equation for molality.

$$m = \frac{0.077 \text{ mol NaCl}}{0.1000 \text{ kg } H_2O} = \mathbf{0.77 \text{ mol/kg}}$$

Substitute moles of solute = 0.077 mol NaCl, kilograms of solvent = 0.1000 kg H_2).

3 Evaluate the Answer

Because there was less than one-tenth of a mole of solute present in one-tenth of a kilogram of water, the molality should be less than one, and it is. The mass of sodium chloride was given with two significant figures; therefore, the molality is also expressed with two significant figures.

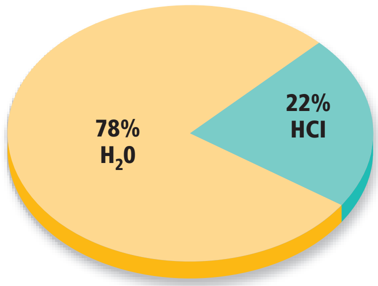
PRACTICE Problems

Extra Practice Pages 985–986 and glencoe.com

27. What is the molality of a solution containing 10.0 g of Na_2SO_4 dissolved in 1000.0 g of water?

28. Challenge How much ($Ba(OH)_2$), in grams, is needed to make a 1.00*m* aqueous solution?

Hydrochloric Acid in Aqueous Solution



$$X_{HCl} + X_{H_20} = 1.00$$
$$0.22 + 0.78 = 1.00$$

Figure 14.8 The mole fraction expresses the number of moles of solvent and solute relative to the total number of moles of solution. Each mole fraction can be thought of as a percent. For example, the mole fraction of water (X_{H_20}) is 0.78, which is equivalent to saying the solution contains 78% water (on a mole basis).

Mole fraction If you know the number of moles of solute and solvent, you can also express the concentration of a solution in what is known as a **mole fraction**—the ratio of the number of moles of solute in solution to the total number of moles of solute and solvent.

The symbol X is commonly used for mole fraction, with a subscript to indicate the solvent or solute. The mole fraction for the solvent (X_A) and the mole fraction for the solute (X_B) can be expressed as follows.

Mole Fraction

$$X_A = \frac{n_A}{n_A + n_B} \qquad X_B = \frac{n_A}{n_A + n_B}$$

***X_A and X_b* represent the mole fractions of each substance.**

n_A and n_B represent the number of moles of each substance.

A mole fraction equals the number of moles of solute in a solution divided by the total number of moles of solute and solvent.

For example, 100 g of a hydrochloric acid solution contains 36 g of HCl and 64 g of H_2O, as shown in **Figure 14.8.** To convert these masses to moles, you would use the molar masses as conversion factors.

$$n_{HCl} = 36\cancel{\text{ g HCl}} \times \frac{1\text{ mol HCl}}{36.5\cancel{\text{ g HCl}}} = 0.99\text{ mol HCl}$$

$$n_{H_2O} = 64\cancel{\text{ g H}_2\text{O}} \times \frac{1\text{ mol H}_2\text{O}}{18.0\cancel{\text{ g H}_2\text{O}}} = 3.60\text{ mol H}_2\text{O}$$

The mole fractions of HCl and water can be expressed as follows.

$$X_{HCl} = \frac{n_{HCl}}{n_{HCl} + n_{H_2O}} = \frac{0.99\text{ mol HCl}}{0.99\text{ mol HCl} + 3.60\text{ mol H}_2\text{O}} = 0.22$$

$$X_{H_2O} = \frac{n_{H_2O}}{n_{HCl} + n_{H_2O}} = \frac{3.60\text{ mol H}_2\text{O}}{0.99\text{ mol HCl} + 3.60\text{ mol H}_2\text{O}} = 0.78$$

PRACTICE Problems

Extra Practice Pages 985–986 and glencoe.com

29. What is the mole fraction of NaOH in an aqueous solution that contains 22.8% NaOH by mass?

30. Challenge If the mole fraction of sulfuric acid (H_2SO_4) in an aqueous solution is 0.325, how much water, in grams, is in 100 mL of the solution?

Section 14.2 Assessment

Section Summary

- Concentrations can be measured qualitatively and quantitatively.
- Molarity is the number of moles of solute dissolved per liter of solution.
- Molality is the ratio of the number of moles of solute dissolved in 1 kg of solvent.
- The number of moles of solute does not change during a dilution.

31. MAIN Idea **Compare and contrast** five quantitative ways to describe the composition of solutions.

32. Explain the similarities and differences between a 1*M* solution of NaOH and a 1*m* solution of NaOH.

33. Calculate A can of chicken broth contains 450 mg of sodium chloride in 240.0 g of broth. What is the percent by mass of sodium chloride in the broth?

34. Solve How much ammonium chloride (NH_4Cl), in grams, is needed to produce 2.5 L of a 0.5*M* aqueous solution?

35. Outline the laboratory procedure for preparing a specific volume of a dilute solution from a concentrated stock solution.

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Section 14.3

Objectives

- **Describe** how intermolecular forces affect solvation.
- **Define** solubility.
- **Understand** what factors affect solubility.

Review Vocabulary

exothermic: a chemical reaction in which more energy is released than is required to break bonds in the initial reactants

New Vocabulary

solvation
heat of solution
unsaturated solution
saturated solution
supersaturated solution
Henry's law

Factors Affecting Solvation

MAIN Idea **Factors such as temperature, pressure, and polarity affect the formation of solutions.**

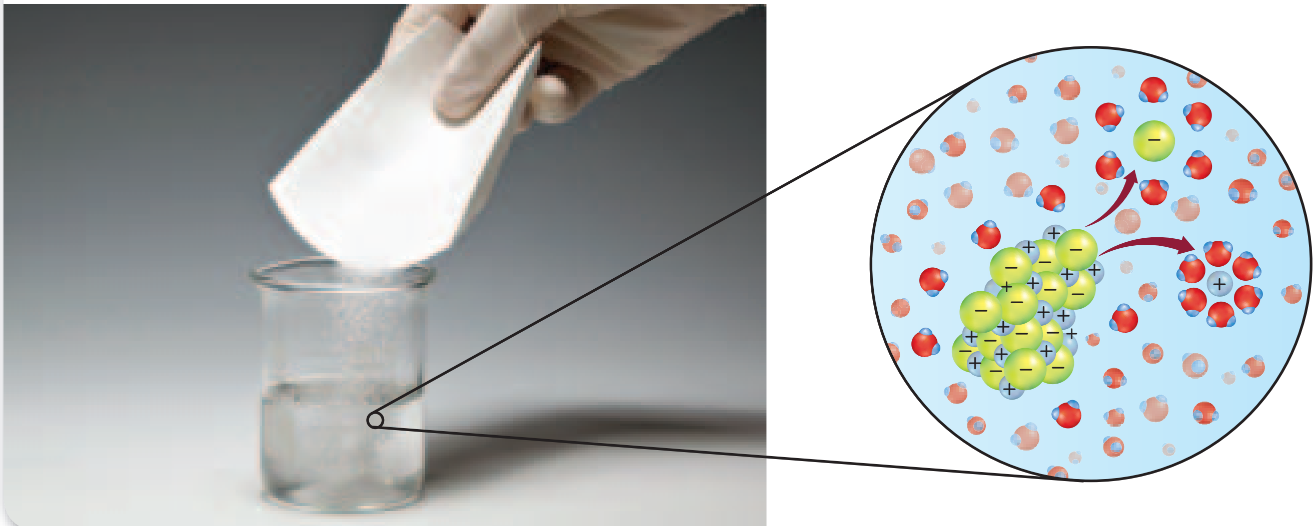
Real-World Reading Link If you have ever made microwavable soup from a dry mix, you added cold water to the dry mix and stirred. At first, only a small amount of the powdered mix dissolves in the cold water. After heating it in the microwave and stirring again, all of the powdered mix dissolves and you have soup.

The Solvation Process

Why are some substances soluble in each other, while others are not? To form a solution, solute particles must separate from one another and the solute and solvent particles must mix. Recall from Chapter 12 that attractive forces exist among the particles of all substances. Attractive forces exist between the pure solute particles, between the pure solvent particles, and between the solute and solvent particles. When a solid solute is placed in a solvent, the solvent particles completely surround the surface of the solid solute. If the attractive forces between the solvent and solute particles are greater than the attractive forces holding the solute particles together, the solvent particles pull the solute particles apart and surround them. These surrounded solute particles then move away from the solid solute and out into the solution.

The process of surrounding solute particles with solvent particles to form a solution is called **solvation,** as shown in **Figure 14.9.** Solvation in water is called hydration. "Like dissolves like" is the general rule used to determine whether solvation will occur in a specific solvent. To determine whether a solvent and solute are alike, you must examine the bonding and polarity of the particles and the intermolecular forces among particles.

■ **Figure 14.9** Salt begins to separate when it is dropped into water. The solute particles are pulled from the solid and surrounded by solvent particles.



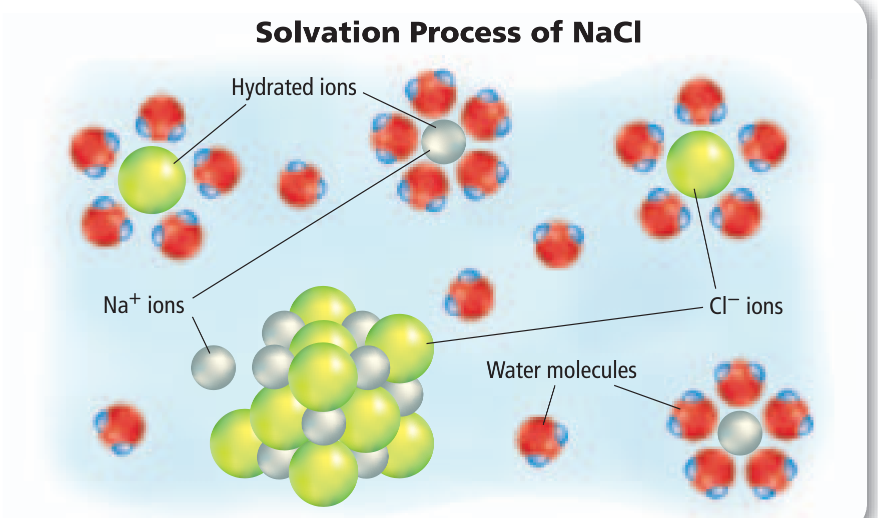
Matt Meadows

Figure 14.10 Sodium chloride dissolves in water as the water molecules surround the sodium and chloride ions. Note how the polar water molecules orient themselves differently around the positive and negative ions.

CLICK HERE

Concepts In Motion

Interactive Figure To see an animation of the dissolution of compounds, visit glencoe.com.



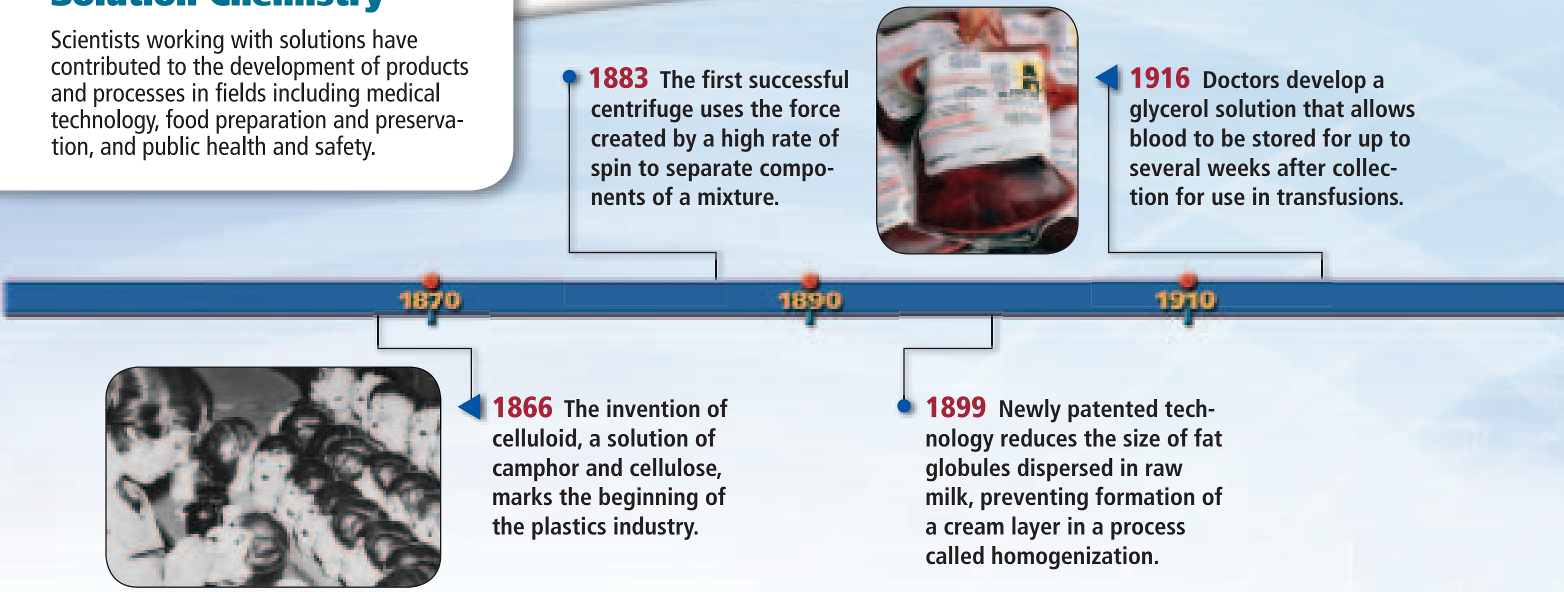
Aqueous solutions of ionic compounds Recall that water molecules are polar molecules and are in constant motion, as described by the kinetic-molecular theory. When a crystal of an ionic compound, such as sodium chloride (NaCl), is placed in water, the water molecules collide with the surface of the crystal. The charged ends of the water molecules attract the positive sodium ions and negative chloride ions. This attraction between the dipoles and the ions is greater than the attraction among the ions in the crystal, so the ions break away from the surface. The water molecules surround the ions, and the solvated ions move into the solution, shown in **Figure 14.10,** exposing more ions on the surface of the crystal. Solvation continues until the entire crystal has dissolved.

Not all ionic substances are solvated by water molecules. Gypsum is insoluble in water because the attractive forces between the ions in gypsum are so strong that they cannot be overcome by the attractive forces of the water molecules. As shown in **Figure 14.11,** the discoveries of specific solutions and mixtures, such as plaster made out of gypsum, have contributed to the development of many products and processes.

Figure 14.11

Milestones in Solution Chemistry

Scientists working with solutions have contributed to the development of products and processes in fields including medical technology, food preparation and preservation, and public health and safety.



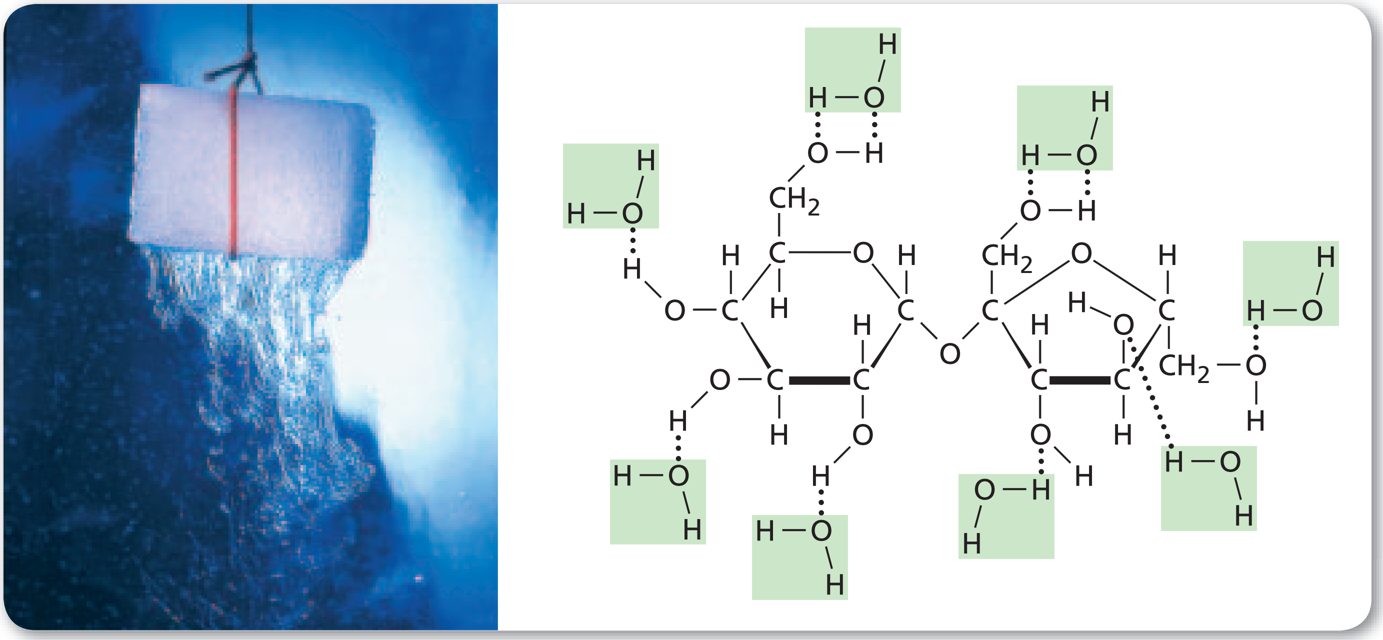
1883 The first successful centrifuge uses the force created by a high rate of spin to separate components of a mixture.

1916 Doctors develop a glycerol solution that allows blood to be stored for up to several weeks after collection for use in transfusions.

1866 The invention of celluloid, a solution of camphor and cellulose, marks the beginning of the plastics industry.

1899 Newly patented technology reduces the size of fat globules dispersed in raw milk, preventing formation of a cream layer in a process called homogenization.

(l)©Hulton-Deutsch Collection/CORBIS, (r)©SuperStock, Inc./SuperStock



■ **Figure 14.12** Sucrose molecules contain eight O–H bonds and are polar. Polar water molecules form hydrogen bonds with the O–H bonds, which pulls the sucrose into solution.

Aqueous solutions of molecular compounds Water is also a good solvent for many molecular compounds. Table sugar is the molecular compound sucrose. As shown in **Figure 14.12,** sucrose molecules are polar and have several O–H bonds. As soon as the sugar crystals contact the water, water molecules collide with the outer surface of the crystal. Each O–H bond becomes a site for hydrogen bonding with water. The attractive forces among sucrose molecules are overcome by the attractive forces between polar water molecules and polar sucrose molecules. Sucrose molecules leave the crystal and become solvated by water molecules.

Oil is a substance made up primarily of carbon and hydrogen. It does not form a solution with water. There is little attraction between the polar water molecules and the nonpolar oil molecules. However, oil spills can be cleaned up with a nonpolar solvent because nonpolar solutes are more readily dissolved in nonpolar solvents.

1964 Stephanie Kwolek discovers a synthetic fiber, formed from liquid crystals in solution, that is stronger than steel and lighter than fiberglass.

2003 Scientists develop chemical packets that remove toxic metals and pesticides and kill pathogens in drinking water. They can be distributed to survivors of natural disasters.

1943 The first artificial kidney removes toxins dissolved in a patient's blood.

1980 Gypsum board is developed as a firewall system to separate townhome and condominium units.

Concepts in Motion

Interactive Time Line To learn more about these discoveries and others, visit glencoe.com.

Chemistry Online

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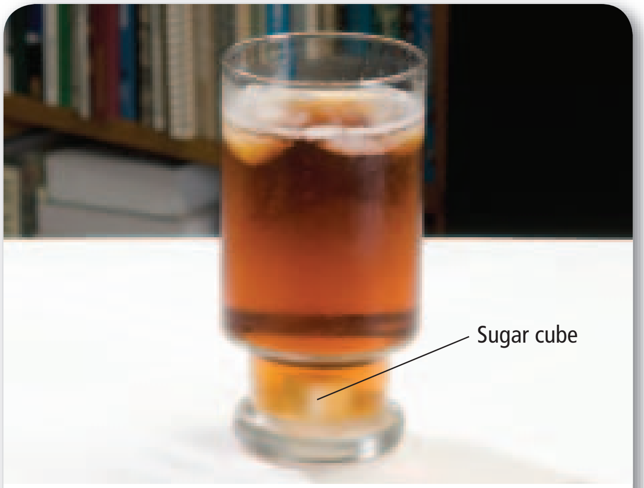

A sugar cube in iced tea will dissolve slowly, but stirring will make the sugar cube dissolve more quickly.

Granulated sugar dissolves more quickly in iced tea than a sugar cube, and stirring will make the granulated sugar dissolve even more quickly.

Granulated sugar dissolves very quickly in hot tea.

■ **Figure 14.13** Agitation, structure, and temperature affect the rate of solvation.

Heat of solution During the process of solvation, the solute must separate into particles. Solvent particles must also move apart in order to allow solute particles to come between them. Energy is required to overcome the attractive forces within the solute and within the solvent, so both steps are endothermic. When solute and solvent particles mix, the particles attract each other and energy is released. This step in the solvation process is exothermic. The overall energy change that occurs during the solution formation process is called the **heat of solution.**

As you observed in the Launch Lab at the beginning of this chapter, some solutions release energy as they form, whereas others absorb energy during formation. For example, after ammonium nitrate dissolves in water, its container feels cool. In contrast, after calcium chloride dissolves in water, its container feels warm.

✓ **Reading Check** **Explain** why some solutions absorb energy during formation, while others release energy during formation.

Factors That Affect Solvation

Solvation occurs only when the solute and solvent particles come in contact with each other. There are three common ways, shown in **Figure 14.13,** to increase the collisions between solute and solvent particles and thus increase the rate at which the solute dissolves: agitation, increasing the surface area of the solute, and increasing the temperature of the solvent.

Agitation Stirring or shaking—agitation of the mixture—moves dissolved solute particles away from the contact surfaces more quickly and thereby allows new collisions between solute and solvent particles to occur. Without agitation, solvated particles move away from the contact areas slowly.

Surface area Breaking the solute into small pieces increases its surface area. A greater surface area allows more collisions to occur. This is why a teaspoon of granulated sugar dissolves more quickly than an equal amount of sugar in cube form.

Temperature The rate of solvation is affected by temperature. For example, sugar dissolves more quickly in hot tea, shown in **Figure 14.13,** than it does in iced tea. Additionally, hotter solvents generally can dissolve more solid solute. Hot tea can hold more dissolved sugar than the iced tea. Most solids act in the same way as sugar—as temperature increases, the rate of solvation also increases. Solvation of other substances, such as gases, decreases at higher temperatures. For example, a carbonated soft drink will lose its fizz (carbon dioxide) faster at room temperature than when cold.

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Solubility

Just as solvation can be understood at the particle level, so can solubility. The solubility of a solute also depends on the nature of the solute and solvent. When a solute is added to a solvent, solvent particles collide with the solute's surface particles; solute particles begin to mix randomly among the solvent particles. At first, the solute particles are carried away from the crystal. However, as the number of solvated particles increases, the same random mixing results in increasingly frequent collisions between solvated solute particles and the remaining crystal. Some colliding solute particles rejoin the crystal, or crystallize, as illustrated in **Figure 14.14.** As solvation continues, the crystallization rate increases, while the solvation rate remains constant. As long as the solvation rate is greater than the crystallization rate, the net effect is continuing solvation.

Depending on the amount of solute present, the rates of solvation and crystallization might eventually equalize. No more solute appears to dissolve and a state of dynamic equilibrium exists between crystallization and solvation (as long as the temperature remains constant).

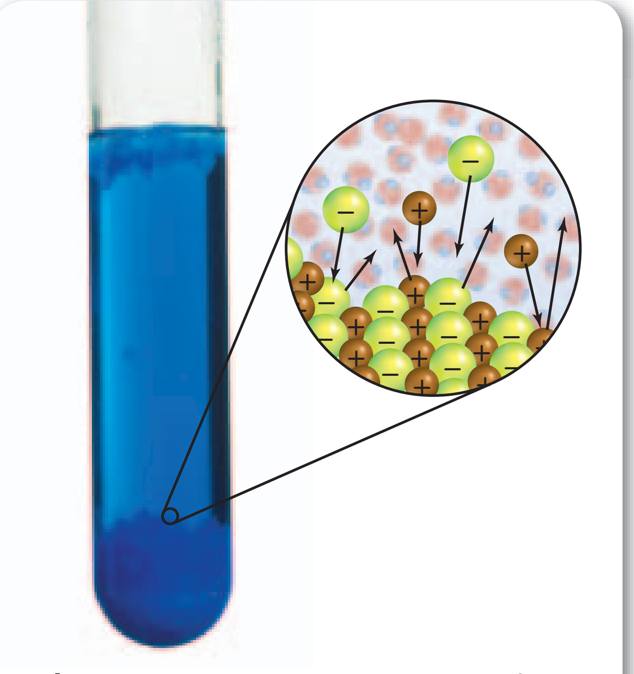


Figure 14.14 In a saturated solution, the rate of solvation equals the rate of crystallization. The amount of dissolved solute does not change.

Unsaturated solutions

An **unsaturated solution** is one that contains less dissolved solute for a given temperature and pressure than a saturated solution. In other words, more solute can be dissolved in an unsaturated solution.

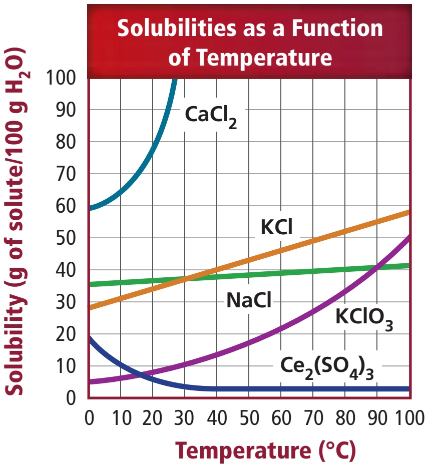
Saturated solutions

Although solute particles continue to dissolve and crystallize in solutions that reach equilibrium, the overall amount of dissolved solute in the solution remains constant. Such a solution, illustrated in **Figure 14.14,** is said to be a **saturated solution;** it contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure.

Temperature and supersaturated solutions

Solubility is affected by raising the temperature of the solvent because the kinetic energy of its particles is increased, resulting in more-frequent collisions and collisions with greater energy than those that occur at lower temperatures. The fact that many substances are more soluble at high temperatures is demonstrated in **Figure 14.15.** For example, calcium chloride ($CaCl_2$) has a solubility of about 64 g $CaCl_2$ per 100 g H_2O at 10°C. Increasing the temperature to approximately 27°C increases the solubility by almost 50%, to 100 g $CaCl_2$ per 100 g H_2O. Other substances, such as cerium sulfate, have decreasing solubility as temperature increases, and then remains constant after a specific temperature is reached.

Figure 14.15 The solubilities of several substances as a function of temperature are shown in graph.



Graph Check

Interpret What is the solubility of NaCl at 80°C?

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Table 14.4 Solubilities of Solutes in Water at Various Temperatures

Substance	Formula	Solubility (g/100 g H_2O)*			
		0°C	20°C	60°C	100°C
Aluminum sulfate	$Al_2(SO_4)_3$	31.2	36.4	59.2	89.0
Barium hydroxide	$Ba(OH)_2$	1.67	3.89	20.94	--
Calcium hydroxide	$Ca(OH)_2$	0.189	0.173	0.121	0.076
Lithium sulfate	Li_2SO_4	36.1	34.8	32.6	--
Potassium chloride	KCl	28.0	34.2	45.8	56.3
Sodium chloride	NaCl	35.7	35.9	37.1	39.2
Silver nitrate	$AgNO_3$	122	216	440	733
Sucrose	$C_{12}H_{22}O_{11}$	179.2	203.9	287.3	487.2
Ammonia*	NH_3	1130	680	200	--
Carbon dioxide*	CO_2	1.713	0.878	0.359	--
Oxygen*	O_2	0.048	0.031	0.019	--

* L/1 L H_2O of gas at standard pressure (101 kPa)

VOCABULARY

WORD ORIGIN

Saturated
comes from the Latin *saturatus* meaning *to fill*

The effect of temperature on solubility is also illustrated by the data in **Table 14.4.** Notice in **Table 14.4** that at 20°C, 203.9 g of sucrose ($C_{12}H_{22}O_{11}$) dissolves in 100 g of water. At 100°C, 487.2 g of sucrose dissolves in 100 g of water, a nearly 140% increase in solubility.

The fact that solubility changes with temperature and that some substances become more soluble with increasing temperature is the key to forming supersaturated solutions. A **supersaturated solution** contains more dissolved solute than a saturated solution at the same temperature. To make a supersaturated solution, a saturated solution is formed at a high temperature and then cooled slowly. The slow cooling allows the excess solute to remain dissolved in solution at the lower temperature, as shown in **Figure 14.16.**

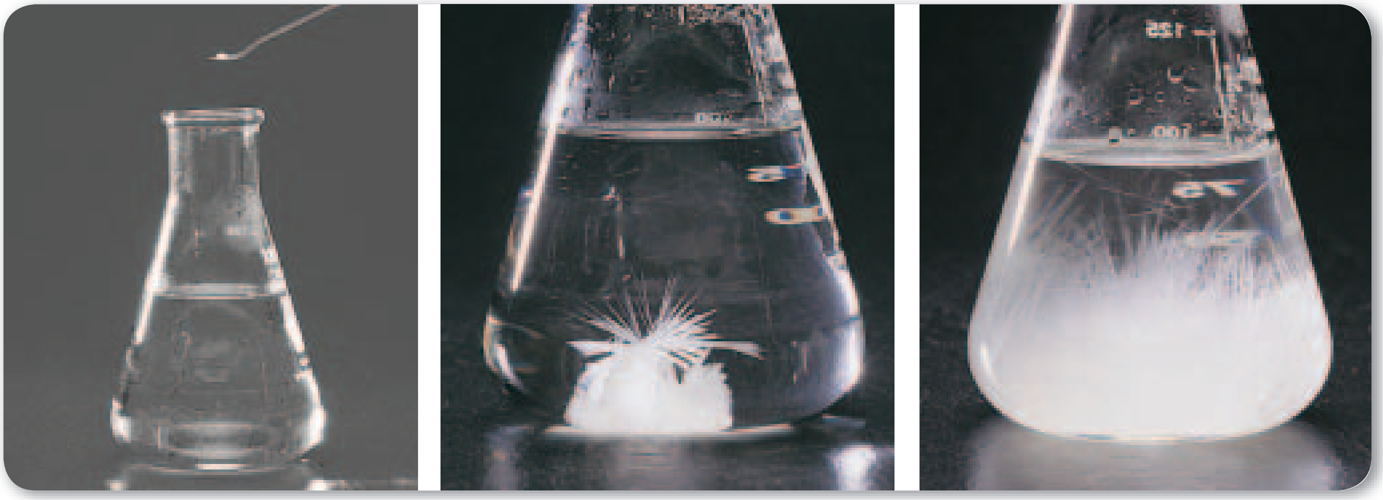

Figure 14.16 When a seed crystal is added to a supersaturated solution, the excess solute crystallizes out of the solution.

The McGraw-Hill Companies, Inc./Stephen Frisch, photographer

■ **Figure 14.17** Hot spring mineral deposits are an example of crystals that formed from supersaturated solutions.

Supersaturated solutions are unstable. If a tiny amount of solute, called a seed crystal, is added to a supersaturated solution, the excess solute precipitates quickly, as illustrated in **Figure 14.16.** Crystallization can also occur if the inside of the container is scratched or the supersaturated solution undergoes a physical shock, such as stirring or tapping the container. Using crystals of silver iodide (AgI) to seed air that is supersaturated with water vapor causes the water particles to come together and form droplets that might fall to Earth as rain. This technique is called cloud seeding. Rock candy and mineral deposits at the edges of mineral springs, such as those shown in **Figure 14.17,** are both formed from supersaturated solutions.

Solubility of gases The gases oxygen and carbon dioxide are less soluble at higher temperatures than at lower temperatures. This is a predictable trend for all gaseous solutes in liquid solvents. Can you explain why? Recall from Chapter 12 that the kinetic energy of gas particles allows them to escape from a solution more readily at higher temperatures. Thus, as a solution's temperature increases, the solubility of a gaseous solute decreases.

Pressure and Henry's law Pressure affects the solubility of gaseous solutes in solutions. The solubility of a gas in any solvent increases as its external pressure (the pressure above the solution) increases. Carbonated beverages depend on this fact. Carbonated beverages contain carbon dioxide gas dissolved in an aqueous solution. In bottling or canning the beverage, carbon dioxide is dissolved in the solution at a pressure higher than atmospheric pressure. When the beverage container is opened, the pressure of the carbon dioxide gas in the space above the liquid decreases. As a result, bubbles of carbon dioxide gas form in the solution, rise to the top, and escape. Unless the container is sealed, the process will continue until the solution loses almost all of its carbon dioxide gas and goes flat. The decreased solubility of the carbon dioxide contained in the beverage after it is opened can be described by Henry's law.

VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Pressure

Science usage: the force exerted over an area

As carbon dioxide escapes the solution, the pressure in the closed bottle increases.

Common usage: The burden of physical or mental stress

There is a lot of pressure to do well on exams.

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Henry's law states that at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid. When the bottle of soda is closed, as illustrated in **Figure 14.18,** the pressure above the solution keeps carbon dioxide from escaping the solution. You can express this relationship in the following way.

Henry's Law

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S represents solubility.
P represents pressure.

At a given temperature, the quotient of solubility of a gas and its pressure is constant.

You will often use Henry's law to determine the solubility S_2 at a new pressure P_2, where P_2 is known. The basic rules of algebra can be used to solve Henry's law for any one specific variable. To solve for S_2, begin with the standard form of Henry's law.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Cross multiplying yields the following expression.

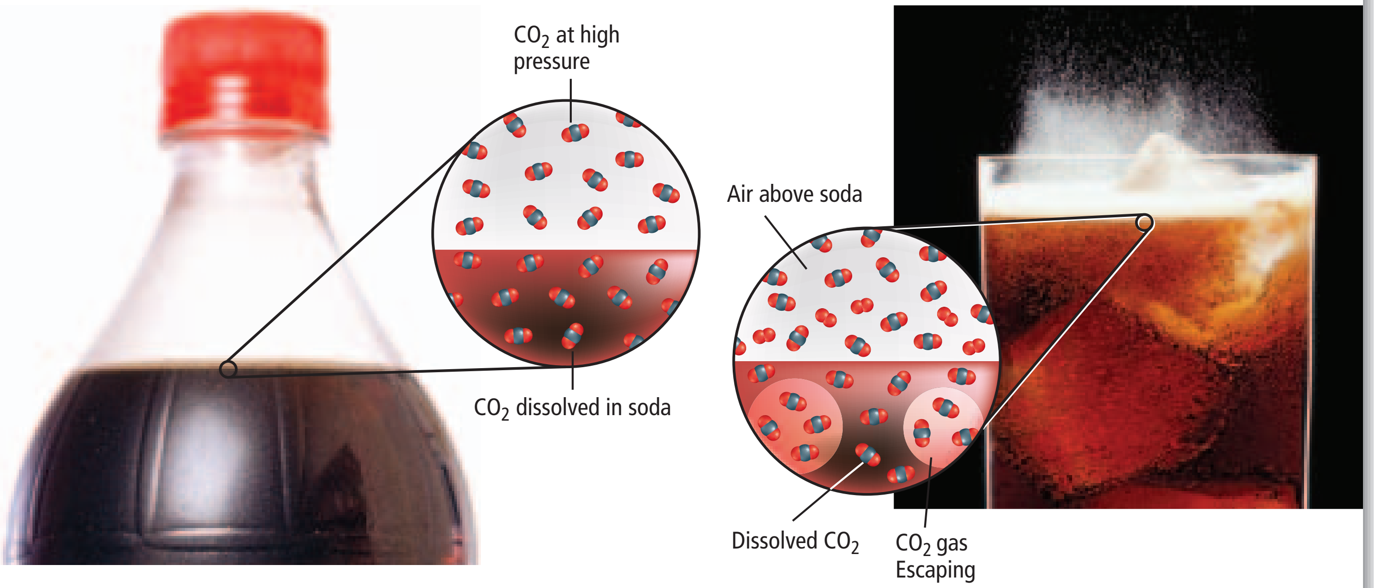
$$S_1P_2 = P_1S_2$$

Dividing both sides of the equation by P_1 yields the desired result—the equation solved for S_2.

$$\frac{S_1P_2}{P_1} = \frac{\cancel{P_1}S_2}{\cancel{P_1}} \qquad S_2 = \frac{S_1P_2}{P_1}$$

Figure 14.18 Carbon dioxide (CO_2) is dissolved in soda. Some CO_2 also is found in the gas above the liquid.

Explain ***Why does the carbon dioxide escape from the solution when the cap is removed?***



The pressure above the solution of a closed soda bottle keeps excess carbon dioxide from escaping the solution.

The pressure above the solution decreases when the cap is removed, which decreases the solubility of the carbon dioxide.

(t)©Marilyn Genter/The Image Works, (bl)©Rachel Epstein/Photo Edit, (br)©Royalty-Free/Corbis

EXAMPLE Problem 14.5

Henry's Law If 0.85 g of a gas at 4.0 atm of pressure dissolves in 1.0 L of water at 25°C, how much will dissolve in 1.0 L of water at 1.0 atm of pressure and the same temperature?

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1 Analyze the Problem

You are given the solubility of a gas at an initial pressure. The temperature of the gas remains constant as the pressure changes. Because decreasing pressure reduces a gas's solubility, less gas should dissolve at the lower pressure.

Known	**Unknown**
$S_1 = 0.85$ g/L	$S_2 = ?$ **g/L**
$P_1 = 4.0$ atm	
$P_2 = 1.0$ atm	

2 Solve for the Unknown

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$ **State Henry's law.**

$$S_2 = S_1\left(\frac{P_2}{P_1}\right)$$ **Solve Henry's law to solve for S_2.**

$$S_2 = \left(\frac{0.85\text{ g}}{1.0\text{ L}}\right)\left(\frac{1.0\text{ atm}}{4.0\text{ atm}}\right) = 0.21\text{ g/L}$$ **Substitute $S_1 = 0.85$ g/L, $P_1 = 4.0$ atm, and $P_2 = 1.0$ atm. Multiply and divide numbers and units.**

3 Evaluate the Answer

The solubility decreased as expected. The pressure on the solution was reduced from 4.0 atm to 1.0 atm, so the solubility should be reduced to one-fourth its original value, which it is. The unit g/L is a solubility unit, and there are two significant figures.

PRACTICE Problems

Extra Practice Page 986 and glencoe.com

36. If 0.55 g of a gas dissolves in 1.0 L of water at 20.0 kPa of pressure, how much will dissolve at 110.0 kPa of pressure?

37. A gas has a solubility of 0.66 g/L at 10.0 atm of pressure. What is the pressure on a 1.0-L sample that contains 1.5 g of gas?

38. Challenge The solubility of a gas at 7 atm of pressure is 0.52 g/L. How many grams of the gas would be dissolved per 1 L if the pressure was raised to 10 atm?

Section 14.3 Assessment

Section Summary

- The process of solvation involves solute particles surrounded by solvent particles.
- Solutions can be unsaturated, saturated, or supersaturated.
- Henry's law states that at a given temperature, the solubility *(S)* of a gas in a liquid is directly proportional to the pressure *(P)* of the gas above the liquid.

39. MAIN Idea **Describe** factors that affect the formation of solutions.

40. Define *solubility*.

41. Describe how intermolecular forces affect solvation.

42. Explain on a particle basis why the vapor pressure of a solution is lower than a pure solvent.

43. Sumarize If a seed crystal was added to a supersaturated solution, how would you characterize the resulting solution?

44. Make and Use Graphs Use the information in **Table 14.3** to graph the solubilities of aluminum sulfate, lithium sulfate, and potassium chloride at 0°C, 20°C, 60°C, and 100°C. Which substance's solubility is most affected by increasing temperature?

Chemistry Online **Self-Check Quiz** glencoe.com

Section 14.4

Objectives

- **Describe** colligative properties.
- **Identify** four colligative properties of solutions.
- **Determine** the boiling point elevation and freezing point depression of a solution.

Review Vocabulary

ion: an atom that is electrically charged

New Vocabulary

colligative property
vapor pressure lowering
boiling point elevation
freezing point depression
osmosis
osmotic pressure

Colligative Properties of Solutions

MAIN Idea Colligative properties depend on the number of solute particles in a solution.

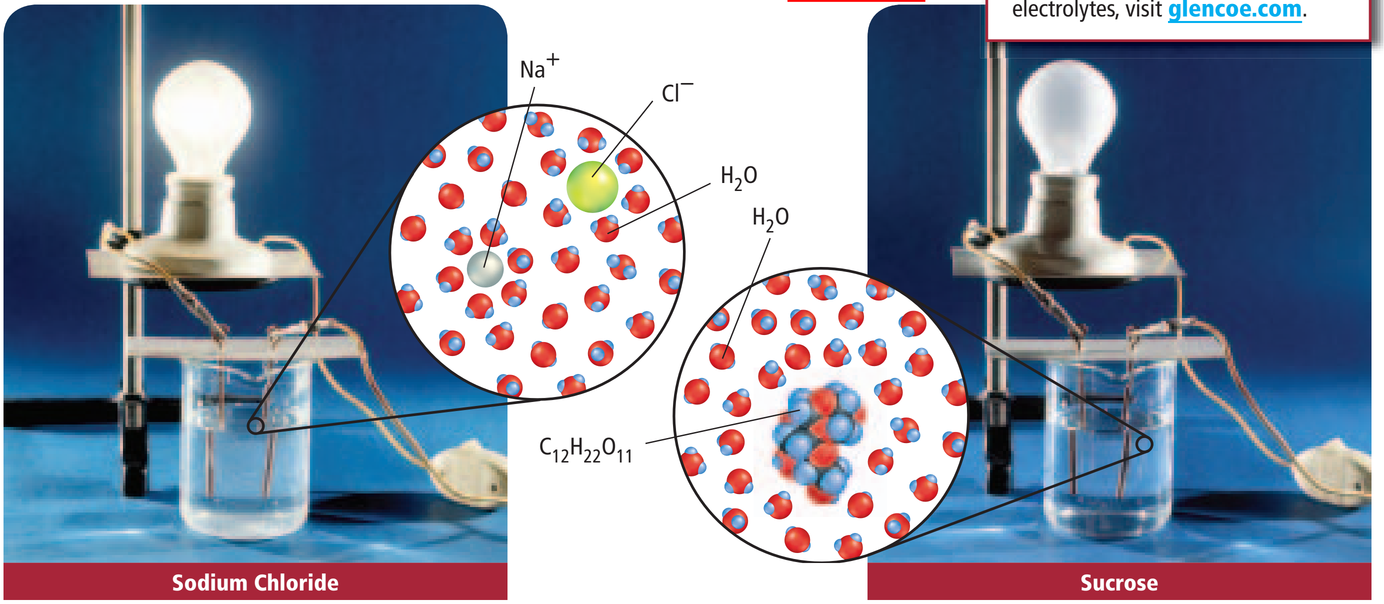
Real-World Reading Link If you live in an area that experiences cold winters, you have probably noticed people spreading salt to melt icy sidewalks and roads. How does salt help make a winter's drive safer?

Electrolytes and Colligative Properties

Solutes affect some of the physical properties of their solvents. Early researchers were puzzled to discover that the effects of a solute on a solvent depended only on how many solute particles were in the solution, not on the specific solute dissolved. Physical properties of solutions that are affected by the number of particles but not by the identity of dissolved solute particles are called **colligative properties.** The word *colligative* means *depending on the collection.* Colligative properties include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.

Electrolytes in aqueous solution In Chapter 8, you read that ionic compounds are called electrolytes because they dissociate in water to form a solution that conducts electric current, as shown in **Figure 14.19.** Some molecular compounds ionize in water and are also electrolytes. Electrolytes that produce many ions in a solution are called strong electrolytes; those that produce only a few ions in a solution are called weak electrolytes.

Figure 14.19 Sodium chloride conducts electricity well because it is an electrolyte. Sucrose does not conduct electricity because it is not an electrolyte.



CLICK HERE

Concepts In Motion

Interactive Figure To see an animation of strong, weak, and nonelectrolytes, visit glencoe.com.

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Sodium chloride is a strong electrolyte. It dissociates in solution, producing Na^+ and Cl^- ions.

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

Dissolving 1 mol of NaCl in 1 kg of water would not yield a 1*m* solution of ions. Rather, there would be 2 mol of solute particles in solution—1 mol each of Na^+ and Cl^- ions.

Nonelectrolytes in aqueous solution Many molecular compounds dissolve in solvents but do not ionize. Such solutions do not conduct an electric current, as shown in **Figure 14.18,** and the solutes are called nonelectrolytes. Sucrose is an example of a nonelectrolyte. A 1*m* sucrose solution contains only 1 mol of sucrose particles.

Reading Check **Infer** Which compound would have the greater effect on colligative properties, sodium chloride or sucrose?

Vapor Pressure Lowering

In Chapter 12, you learned that vapor pressure is the pressure exerted in a closed container by liquid particles that have escaped the liquid's surface and entered the gaseous state. In a closed container at constant temperature and pressure, the solvent particles reach a state of dynamic equilibrium, escaping and reentering the liquid state at the same rate.

Experiments show that adding a nonvolatile solute (one that has little tendency to become a gas) to a solvent lowers the solvent's vapor pressure. The particles that produce vapor pressure escape the liquid phase at its surface. When a solvent is pure, as shown in **Figure 14.20,** its particles occupy the entire surface area. When the solvent contains solute, as also shown in **Figure 14.20,** a mix of solute and solvent particles occupies the surface area. With fewer solvent particles at the surface, fewer particles enter the gaseous state, and the vapor pressure is lowered. The greater the number of solute particles in a solvent, the lower the resulting vapor pressure. Thus, **vapor pressure lowering** is due to the number of solute particles in solution and is a colligative property of solutions.

You can predict the relative effect of a solute on vapor pressure based on whether the solute is an electrolyte or a nonelectrolyte. For example, 1 mol each of the solvated nonelectrolytes glucose, sucrose, and ethanol molecules has the same relative effect on the vapor pressure. However, 1 mol each of the solvated electrolytes sodium chloride (NaCl), sodium sulfate (Na_2SO_4), and aluminum chloride ($AlCl_3$) has an increasingly greater effect on vapor pressure because of the increasing number of ions each produces in solution.

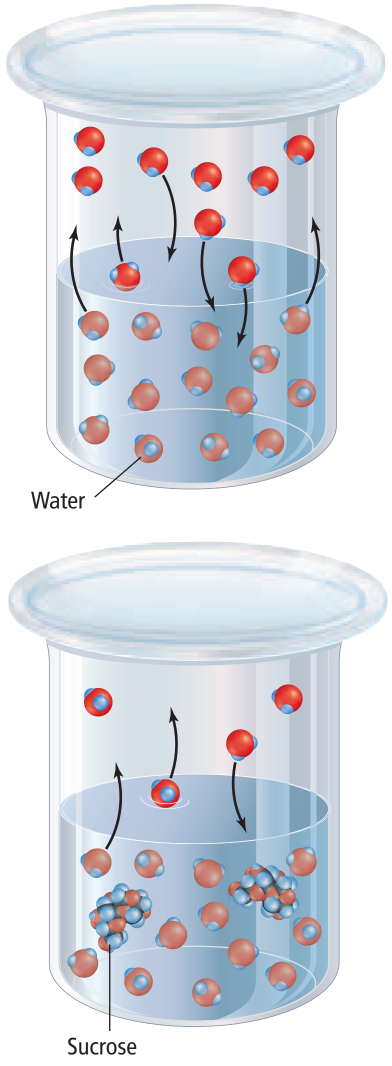

Figure 14.20 The vapor pressure of a pure solvent is greater than the vapor pressure of a nonvolatile solution.

Boiling Point Elevation

Because a nonvolatile solute lowers a solvent's vapor pressure, it also affects the boiling point of the solvent. Recall from Chapter 12 that liquid in a pot on a stove boils when its vapor pressure equals the atmospheric pressure. When the temperature of a solution containing a nonvolatile solute is raised to the boiling point of the pure solvent, the resulting vapor pressure is still less than the atmospheric pressure and the solution will not boil. Thus, the solution must be heated to a higher temperature to supply the additional kinetic energy needed to raise the vapor pressure to atmospheric pressure. The temperature difference between a solution's boiling point and a pure solvent's boiling point is called the **boiling point elevation.**

For nonelectrolytes, the value of the boiling point elevation, which is symbolized ΔT_b, is directly proportional to the solution's molality.

Boiling Point Elevation

$$\Delta T_b = K_b m$$

ΔT_b represents the boiling point elevation.
K_b represents the molal boiling elevation constant.
m represents molality.

The temperature difference is equal to the molal boiling point elevation constant multiplied by the solution's molality.

The molal boiling point elevation constant, K_b, is the difference in boiling points between a 1*m* nonvolatile, nonelectrolyte solution and a pure solvent. Boiling point elevation is expressed in units of °C/*m* and varies for different solvents. Values of K_b for several common solvents are found in **Table 14.5.** Note that water's K_b value is 0.512°C/*m*. This means that a 1*m* aqueous solution containing a nonvolatile, nonelectrolyte solute boils at 100.512°C—a temperature just 0.512°C higher than pure water's boiling point of 100.0°C.

Like vapor pressure lowering, boiling point elevation is a colligative property. The value of the boiling point elevation is directly proportional to the solution's solute molality; that is, the greater the number of solute particles in the solution, the greater the boiling point elevation. Because it is related to mole fraction, which involves the number of solute particles, molality is used as the concentration. Molality also uses mass of solvent rather than volume, and therefore is not affected by temperature changes. Examine **Figure 14.21** and notice that the curve for a solution lies below the curve for the pure solvent at any temperature.

Table 14.5 Molal Boiling Point Elevation Constants (K_b)

Solvent	Boiling Point (°C)	K_b (°C/*m*)
Water	100.0	0.512
Benzene	80.1	2.53
Carbon tetrachloride	76.7	5.03
Ethanol	78.5	1.22
Chloroform	61.7	3.63

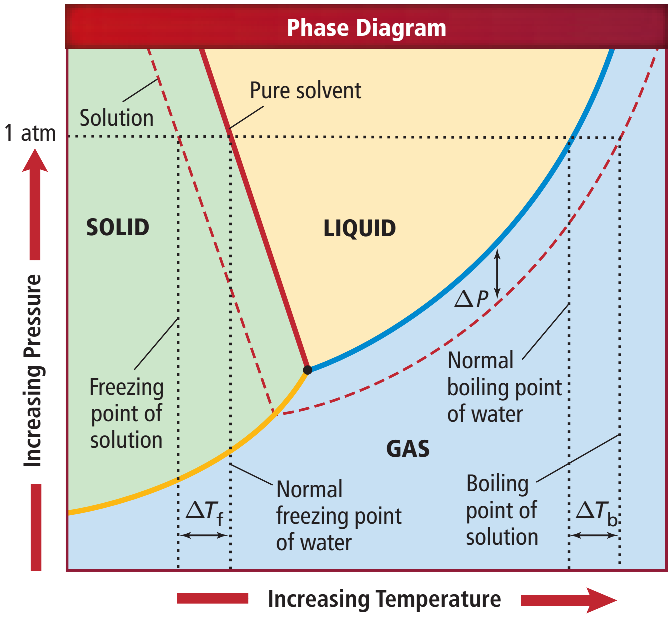

Figure 14.21 Temperature and pressure affect solid, liquid, and gas phases of a pure solvent (solid lines) and a solution (dashed line).

Graph Check

Describe how the difference between the solid lines and dashed line corresponds to vapor pressure lowering, boiling point elevation, and freezing point depression. Use specific data from the graph to support your answer.

Freezing Point Depression

At a solvent's freezing point temperature, the particles no longer have sufficient kinetic energy to overcome the interparticle attractive forces; the particles form into a more organized structure in the solid state. In a solution, the solute particles interfere with the attractive forces among the solvent particles. This prevents the solvent from entering the solid state at its normal freezing point.

The freezing point of a solution is always lower than that of a pure solvent. **Figure 14.21** shows the differences in boiling and melting points of pure water and an aqueous solution. By comparing the solid and dashed lines, you can see that the temperature range over which the aqueous solution exists as a liquid is greater than that of pure water. Two common applications of freezing point depression, shown in **Figure 14.22,** use salt to lower the freezing point of a water solution.

Figure 14.22 By adding salts to the ice on a road, the freezing point of the ice is lowered, which results in the ice melting. Adding salt to ice when making ice cream lowers the freezing point of the ice, allowing the resulting water to freeze the ice cream.

(l)AP Photo/Gerry Broome, (r)©Tom Pantages

Table 14.6 Molal Freezing Point Depression Constants (K_f)

Solvent	Freezing Point (°C)	K_f (°C/m)
Water	0.0	1.86
Benzene	5.5	5.12
Carbon tetrachloride	−23.0	29.8
Ethanol	−114.1	1.99
Chloroform	−63.5	4.68

A solution's **freezing point depression,** ΔT_f, is the difference in temperature between its freezing point and the freezing point of its pure solvent. Molal freezing point depression constants (K_f) for several solvents are shown in **Table 14.6.** For nonelectrolytes, the value of the freezing point depression is directly proportional to the solution's molality.

Freezing Point Depression

$$\Delta T_f = K_f m$$

ΔT_f represents temperature.
K_f is the freezing point depression constant.
m represents molality.

The temperature difference is equal to the freezing point depression constant multiplied by the solution's molality.

As with K_b values, K_f values are specific to their solvents. With water's K_f value of 1.86°C/m, a 1m aqueous solution containing a nonvolatile, nonelectrolye solute freezes at −1.86ºC rather than at pure water's freezing point of 0.0°C. Glycerol is a nonelectrolyte solute produced by many fish and insects to keep their blood from freezing during cold winters. Antifreeze and the de-icer contain the nonelectrolyte solute ethylene glycol.

Notice that the equations for boiling point elevation and freezing point depression specify the molality of a nonelectrolyte. For electrolytes, you must make sure to use the effective molality of the solution. Example Problem 14.6 illustrates this point.

MiniLab

Examine Freezing Point Depression

How do you measure freezing point depression?

Procedure

1. Read and complete the lab safety form.
2. Fill two **400-mL beakers** with **crushed ice.** Add 50 mL of **cold tap water** to each beaker.
3. Measure the temperature of each beaker using a **nonmercury thermometer.**
4. Stir the contents of each beaker with a **stirring rod** until both beakers are at a constant temperature—approximately 1 min. Record the temperature.
5. Add 75 g of **rock salt (NaCl)** to one of the beakers. Continue stirring both beakers. Some of the salt will dissolve.
6. When the temperature in each beaker is constant, record the final readings.
7. To clean up, flush the contents of each beaker down the drain with excess water.

Analysis

1. **Compare** your readings taken for the ice water and the salt water. How do you explain the observed temperature change?
2. **Explain** why salt was added to only one of the beakers.
3. **Explain** Salt is a strong electrolyte that produces two ions, Na^+ and Cl^-, when it dissociates in water. Explain why this is important to consider when calculating the colligative property of freezing point depression.
4. **Predict** whether it would be better to use coarse rock salt or fine table salt when making homemade ice cream. Explain.

EXAMPLE Problem 14.6

Changes in Boiling and Freezing Points Sodium chloride (NaCl) often is used to prevent icy roads and to freeze ice cream. What are the boiling point and freezing point of a 0.029*m* aqueous solution of sodium chloride?

1 Analyze the Problem

You are given the molality of an aqueous sodium chloride solution. First, calculate ΔT_b and ΔT_f based on the number of particles in the solution. Then, to determine the elevated boiling point and the depressed freezing point, add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point.

Known	**Unknown**
solute = sodium chloride (NaCl)	**boiling point = ?°C**
molality of solution = 0.029*m*	**freezing point = ?°C**

2 Solve for the Unknown

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Determine the molality of the particles.

particle molality = 2 × 0.029*m* = 0.058*m*

$\Delta T_b = K_b m$
$\Delta T_f = K_f m$

State the boiling point elevation and freezing point depression formulas.

Determine ΔT_b and ΔT_f.

$\Delta T_b = (0.512°\text{C}/\cancel{m})(0.058\cancel{m}) = 0.030°\text{C}$
$\Delta T_f = (1.86°\text{C}/\cancel{m})(0.058\cancel{m}) = 0.11°\text{C}$

Substitute $K_b = 0.512°\text{C}/m$, $K_f = 1.86°\text{C}/m$, and $m = 0.058m$.

Determine the elevated boiling point and depressed freezing point of the solution.

boiling point = 100.000°C + 0.030°C
= **100.030°C**

freezing point = 0.00°C − 0.11°C
= **−0.11°C**

Add ΔT_b to the normal boiling point and subtract ΔT_f from the normal freezing point.

3 Evaluate the Answer

The boiling point is higher and the freezing point is lower, as expected. Because the molality of the solution has two significant figures, both ΔT_bd Δ have two significant figures. Because the normal boiling point and freezing point are exact values, they do not affect the number of significant figures in the final answer.

PRACTICE Problems

Extra Practice Page 986 and glencoe.com

45. What are the boiling point and freezing point of a 0.625*m* aqueous solution of any nonvolatile, nonelectrolyte solute?

46. What are the boiling point and freezing point of a 0.40*m* solution of sucrose in ethanol?

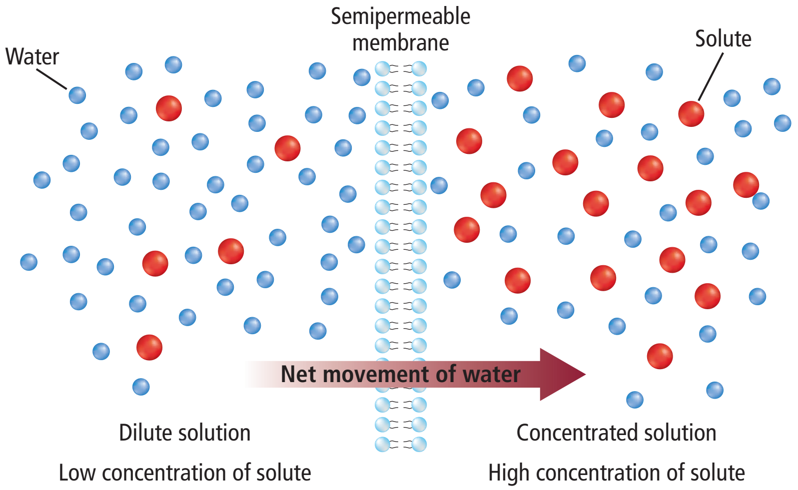
47. Challenge A 0.045m solution (consisting of a nonvolatile, nonelectrolyte solute) is experimentally found to have a freezing point depression of 0.08°C. What is the freezing point depression constant (K_f). Which is most likely to be the solvent: water, ethanol, or chloroform?

Real-World Chemistry

Freezing Point Depression

Saltwater fish Maintaining the proper saline (salt) concentration is important to the health of saltwater fish. In the ocean, the presence of salt in arctic areas keeps the water from freezing, allowing aquatic life to be sustained.

Figure 14.23 Due to osmosis, solvents diffuse from a higher concentration to a lower concentration through semipermeable membranes.



CLICK HERE

Concepts In Motion

Interactive Figure To see an animation of osmosis, visit glencoe.com.

Osmotic Pressure

Connection to Biology Recall from Chapter 12 that diffusion is the mixing of gases or liquids resulting from their random motions. **Osmosis** is the diffusion of a solvent through a semipermeable membrane. Semipermeable membranes are barriers that allow some, but not all, particles to cross. The membranes surrounding all living cells are semipermeable membranes. Osmosis plays an important role in many biological systems , such as the uptake of nutrients by plants.

Examine a system in which a dilute solution is separated from a concentrated solution by a semipermeable membrane, illustrated in **Figure 14.23.** During osmosis, water molecules move in both directions across the membrane, but the solute molecules cannot cross it. Water molecules diffuse across the membrane from the dilute solution to the concentrated solution. The amount of additional pressure caused by the water molecules that moved into the concentrated solution is called the **osmotic pressure.** Osmotic pressure depends on the number of solute particles in a given volume of solution and is a colligative property of solutions.

Section 14.4 Assessment

Section Summary

- Nonvolatile solutes lower the vapor pressure of a solution.
- Boiling point elevation is directly related to the solution's molality.
- A solution's freezing point depression is always lower than that of the pure solvent.
- Osmotic pressure depends on the number of solute particles in a given volume.

48. **MAIN Idea** **Explain** the nature of colligative properties.
49. **Describe** four colligative properties of solutions.
50. **Explain** why a solution has a lower boiling point than that of the pure solvent.
51. **Solve** An aqueous solution of calcium chloride ($CaCl_2$) boils at 101.3°C. How many kilograms of calcium chloride were dissolved in 1000.0 g of the solvent?
52. **Calculate** the boiling point elevation of a solution containing 50.0 g of glucose ($C_6H_{12}O_6$) dissolved in 500.0 g of water. Calculate the freezing point depression for the same solution.
53. **Investigate** A lab technician determines the boiling point elevation of an aqueous solution of a nonvolatile, nonelectrolyte to be 1.12°C. What is the solution's molality?

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In the Field

Career: Environmental Chemist

A CO_2 Solution

Geologic records indicate that the levels of atmospheric carbon dioxide (CO_2) are likely higher today than in the past 20 million years. Anthropogenic (an thruh pah JEN ihk) CO_2, which means CO_2 from human-made sources, has contributed to this high level. CO_2 does not remain in the atmosphere indefinitely. Oceans naturally contain CO_2 that comes from the atmosphere and from living organisms. Oceans have absorbed nearly 50% of anthro-pogenic CO_2. Some scientists think that over the next thousand years, as much as 90% of anthropogenic CO_2 will dissolve in the oceans.

Collecting CO_2 data The rate at which CO_2 dissolves into the oceans is influenced by many factors including temperature, concentration of CO_2 in the air and in the water, and the mixing of air and water due to wind and waves. A team of researchers spent years collecting and analyzing CO_2 data from thousdands of collection points throughout the world's oceans. The data, shown in **Figure 1,** indicate that the North Atlantic has the most anthropogenic CO_2 per square meter of ocean surface. The combination of temperature, depth, and current make the North Atlantic an efficient absorber of anthropogenic CO_2.

CO_2 capture and storage One way to reduce the amount of CO_2 released into the atmosphere would be to capture and store the CO_2 produced when fossil fuels are burned. Researchers are investigating the possibility of directly injecting captured CO_2 into the ocean to speed up the dissolution process. This could reduce the greenhouse effect of CO_2 gas. However, upsetting the natural balance of dissolved CO_2 can have profound effects on water chemistry, which can harm or even kill marine life. For example, coral reefs throughout the world already show signs of stress due to increasing levels of dissolved CO_2.

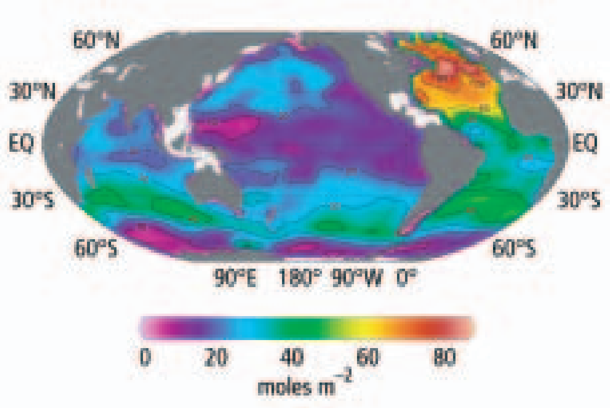


Figure 1 The red, yellow, and green regions represent areas where high levels of anthropogenic CO_2 are dissolved in the water.

Data obtained from: Sabine et al. 2004. The oceanic sink for anthropogenic CO_2. *Science* 305: 367–371.

Deep ocean sequestration A proposal that might reduce atmospheric CO_2 and protect life in the upper ocean is to liquefy the CO_2 and pump it deep under water, a process known as deep ocean sequestration. It is thought that the extreme pressure at depths greater than 3000 m will cause the CO_2 to form a hydrate. The hydrate will dissolve into the deep ocean water, but the CO_2 will remain trapped for hundreds of years far from the upper ocean and atmosphere.

Ongoing research Scientists are working on many of the unanswered questions about deep ocean sequestration, such as the effect of CO_2 on deep-sea animals. There are still many technological problems involving capturing, storing, and transporting large quantities of liquid CO_2. If the technological problems can be solved, the public as well as government officials will have to consider the relative dangers of releasing CO_2 into the air and into the ocean.

WRITING in Chemistry

Brainstorm a list of questions that must be addressed through research before deep ocean sequestration is attempted. Visit glencoe.com for more information on CO_2 sequestration.

Courtesy of Dr. Christopher L. Sabine, National Oceanic and Atmospheric Administration.

CHEMLAB

INVESTIGATE FACTORS AFFECTING SOLUBILITY

Background The process of making a solution involves the solvent coming in contact with the solute particles. When you add a soluble compound to water, several factors affect the rate of solution formation.

Question *How do factors affect the rate of solution formation?*

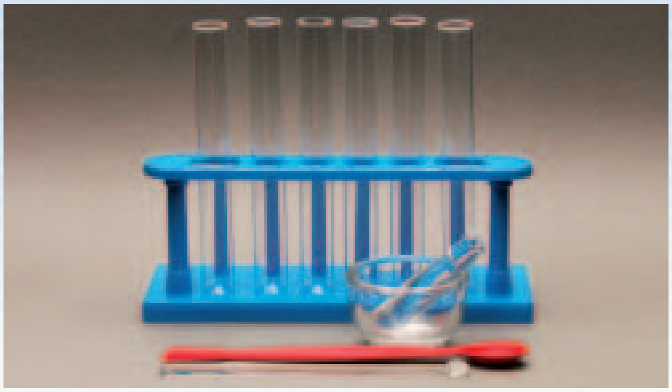
Materials

copper (II) sulfate pentahydrate
distilled water
test tubes (6)
25-mL graduated cylinder
glass stirring rod
tweezers
test tube rack
mortar and pestle
spatula
clock

Safety Precautions

Procedure

1. Read and complete the lab safety form.
2. Create a table to record your data.
3. Write a hypothesis that uses what you know about reaction rates to explain what you might observe during the procedure.
4. Place the 6 test tubes in the test tube rack.
5. Place one crystal of copper (II) sulfate pentahydrate in each of the first two test tubes.
6. For the remaining test tubes, use the mortar and pestle to crush a crystal. Use the spatula to scrape it into the third test tube.
7. Measure 15-mL of room-temperature distilled water. Pour the water into the first test tube and record the time.
8. Observe the solution in the test tube just after adding the water and after 15 min.
9. Leave the first test tube undisturbed in the rack.
10. Repeat Steps 7 and 8 for the third and fourth test tubes.
11. Use the glass stirring rod to agitate the second test tube for 1 to 2 min.
12. Leave the third test tube undisturbed.
13. Agitate the fourth test tube with the glass stirring rod for 1 to 2 min.



14. Repeat Steps 7 and 8 for the fifth test tube using cold water. Leave the fifth test tube undisturbed.
15. Repeat Steps 7 and 8 for the sixth test tube using hot water. Leave the sixth test tube undisturbed.
16. **Cleanup and Disposal** Dispose of the remaining solids and solutions as directed by your teacher. Wash and return all lab equipment to its designated location.

Analyze and Conclude

1. **Compare and Contrast** What effect did you observe due to the agitation of the second and fourth test tubes versus the solutions in the first and third test tubes?
2. **Observe and Infer** What factor caused the more rapid solution formation in the fourth test tube in comparison to the second test tube?
3. **Recognize Cause and Effect** Why do you think the results for the third, fifth, and sixth test tubes were different?
4. **Discuss** whether or not your data supported your hypothesis.
5. **Error Analysis** Identify a major potential source of error for this lab, and suggest an easy method to correct it.

INQUIRY EXTENSION

Think Critically The observations in this lab were macroscopic in nature. Propose a submicroscopic explanation to account for these factors that affected the rate of solution formation. At the molecular level, what is occurring to speed solution formation in each case?

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CHAPTER 14 Study Guide

BIG Idea Nearly all of the gases, liquids, and solids that make up our world are mixtures.

Section 14.1 Types of Mixtures

MAIN Idea Mixtures can be either heterogeneous or homogeneous.

Vocabulary

- Brownian motion (p. 477)
- colloid (p. 477)
- immiscible (p. 479)
- insoluble (p. 479)
- miscible (p. 479)
- soluble (p. 479)
- suspension (p. 476)
- Tyndall effect (p. 478)

Key Concepts

- The individual substances in a heterogeneous mixture remain distinct.
- Two types of heterogeneous mixtures are suspensions and colloids.
- Brownian motion is the erratic movement of colloid particles.
- Colloids exhibit the Tyndall effect.
- A solution can exist as a gas, a liquid, or a solid, depending on the solvent.
- Solutes in a solution can be gases, liquids, or solids.

Section 14.2 Solution Concentration

MAIN Idea Concentration can be expressed in terms of percent or in terms of moles.

Vocabulary

- concentration (p. 480)
- molality (p. 487)
- molarity (p. 482)
- mole fraction (p. 488)

Key Concepts

- Concentrations can be measured qualitatively and quantitatively.
- Molarity is the number of moles of solute dissolved per liter of solution.

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

- Molality is the ratio of the number of moles of solute dissolved in 1 kg of solvent.

$$\text{molality } (m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

- The number of moles of solute does not change during a dilution.

$$M_1V_1 = M_2V_2$$

Section 14.3 Factors Affecting Solvation

MAIN Idea Factors such as temperature, pressure, and polarity affect the formation of solutions.

Vocabulary

- heat of solution (p. 492)
- Henry's law (p. 496)
- saturated solution (p. 493)
- solvation (p. 489)
- supersaturated solution (p. 494)
- unsaturated solution (p. 493)

Key Concepts

- The process of solvation involves solute particles surrounded by solvent particles.
- Solutions can be unsaturated, saturated, or supersaturated.
- Henry's law states that at a given temperature, the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid.

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

Section 14.4 Colligative Properties of Solutions

MAIN Idea Colligative properties depend on the number of solute particles in a solution.

Vocabulary

- boiling point elevation (p. 500)
- colligative property (p. 498)
- freezing point depression (p. 502)
- osmosis (p. 504)
- osmotic pressure (p. 504)
- vapor pressure lowering (p. 499)

Key Concepts

- Nonvolatile solutes lower the vapor pressure of a solution.
- Boiling point elevation is directly related to the solution's molality.

$$\Delta T_b = K_b m$$

- A solution's freezing point depression is always lower than that of the pure solvent.

$$\Delta T_f = K_f m$$

- Osmotic pressure depends on the number of solute particles in a given volume.

CHAPTER 14 Assessment

Section 14.1

Mastering Concepts

54. Explain what is meant by the statement "not all mixtures are solutions."

55. What is the difference between a solute and a solvent?

56. What is a suspension, and how does it differ from a colloid?

57. How can the Tyndall effect be used to distinguish between a colloid and a solution? Why?

58. Name a colloid formed from a gas dispersed in a liquid.

■ **Figure 14.24**

59. Salad dressing What type of heterogenous mixture is shown in **Figure 14.24?** What characteristic is most useful in classifying the mixture?

60. What causes the Brownian motion observed in liquid colloids?

61. Aerosol sprays are categorized as colloids. Identify the phases of an aerosol spray.

Section 14.2

Mastering Concepts

62. What is the difference between percent by mass and percent by volume?

63. What is the difference between molarity and molality?

64. What factors must be considered when creating a dilute solution from a stock solution?

65. How do 0.5*M* and 2.0*M* aqueous solutions of NaCl differ?

66. Under what conditions might a chemist describe a solution in terms of molality? Why?

Mastering Problems

67. According to lab procedure, you stir 25.0 g of $MgCl_2$ into 550 mL of water. What is the percent by mass of $MgCl_2$ in the solution?

68. How many grams of LiCl are in 275 g of a 15% aqueous solution of LiCl?

69. You need to make a large quantity of a 5% solution of HCl but have only 25 mL HCl. What volume of 5% solution can be made from this volume of HCl?

70. Calculate the percent by volume of a solution created by adding 75 mL of acetic acid to 725 mL of water.

71. Calculate the molarity of a solution that contain 15.7 g of $CaCO_3$ dissolved in 275 mL of water.

72. What is the volume of a 3.00*M* solution made with 122 g of LiF?

73. How many moles of BaS would be used to make 1.5×10^3 mL of a 10.0*M* solution?

74. How much $CaCl_2$, in grams, is needed to make 2.0 L of a 3.5*M* solution?

75. Stock solutions of HCl with various molarities are frequently prepared. Complete **Table 14.7** by calculating the volume of concentrated, or 12*M*, hydrochloric acid that should be used to make 1.0 L of HCl solution with each molarity listed.

Table 14.7 HCl Solutions

Molarity of HCl Desired	Volume of 12*M* HCl Stock Solution Needed (mL)
0.5	
1.0	
1.5	
2.0	
5.0	

76. How much of 5.0M nitric acid (HNO_3), in milliliters, is needed to make 225 mL of 1.0M HNO_3?

77. Experiment In the lab, you dilute 55 mL of a 4.0*M* solution to make 250 mL of solution. Calculate the molarity of the new solution.

78. How many milliliters of 3.0*M* phosphoric acid (H_3PO_4) can be made from 95 mL of a 5.0*M* H_3PO_4 solution?

79. If you dilute 20.0 mL of a 3.5*M* solution to make 100.0 mL of solution, what is the molarity of the dilute solution?

80. What is the molality of a solution that contain 75.3 g of KCl dissolved in 95.0 g of water?

81. How many grams of Na_2CO_3 must be dissolved into 155 g of water to create a solution with a molality of 8.20 mol/kg?

82. What is the molality of a solution containing 30.0 g of naphthalene ($C_{10}H_8$) dissolved in 500.0 g of toluene?

83. What are the molality and mole fraction of solute in a 35.5 percent by mass aqueous solution of formic acid (HCOOH)?

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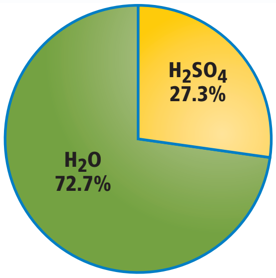


■ **Figure 14.25**

84. What is the mole fraction of H_2SO_4 in a solution containing the percentage of sulfuric acid and water shown in **Figure 14.25?**

85. Calculate the mole fraction of $MgCl_2$ in a solution created by dissolving 132.1 g of $MgCl_2$ in 175 mL of water.

Section 14.3

Mastering Concepts

86. Describe the process of solvation.

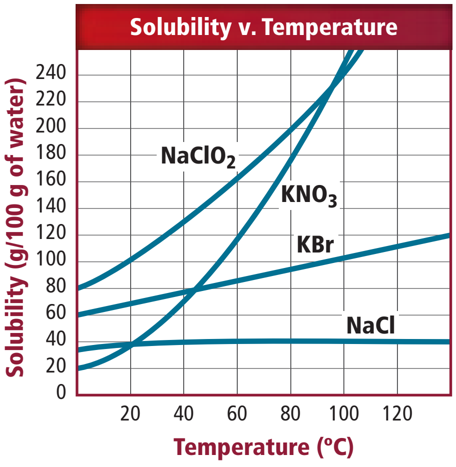
87. What are three ways to increase the rate of solvation?

88. Explain the difference between saturated and unsaturated solutions.

Mastering Problems

89. At a pressure of 1.5 atm, the solubility of a gas is 0.54 g/L. Calculate the solubility when the pressure is doubled.

90. At 4.5 atm of pressure, the solubility of a gas is 9.5 g/L. How much gas, in grams, will dissolve in 1 L if the pressure is reduced by 3.5 atm?



■ **Figure 14.26**

91. Using **Figure 14.26,** compare the solubility of potassium bromide (KBr) and potassium nitrate (KNO_3) at 80°C.

92. The solubility of a gas at 37.0 kPa is 1.80 g/L. At what pressure will the solubility reach 9.00 g/L?

93. Use Henry's law to complete **Table 14.8.**

Table 14.8 Solubility and Pressure

Solubility (g/L)	Pressure (kPa)
2.9	?
3.7	32
?	39

94. Soft Drinks The partial pressure of CO_2 inside a bottle of soft drink is 4.0 atm at 25°C. The solubility of CO_2 is 0.12 mol/L. When the bottle is opened, the partial pressure drops to 3.0×10^{-4} atm. What is the solubility of CO_2 in the open drink? Express your answer in grams per liter.

Section 14.4

Mastering Concepts

95. Define the term *colligative property*.

96. Use the terms *dilute* and *concentrated* to compare the solution on both sides of a membrane.

97. Identify each variable in the following formula:

$$\Delta T_b = K_b m$$

98. Define the term *osmotic pressure,* and explain why it is considered a colligative property.

Mastering Problems

99. Calculate the freezing point of a solution of 12.1 g of naphthalene ($C_{10}H_8$) dissolved in 0.175 kg of benzene (C_6H_6). Refer to **Table 14.6** needed data.

100. In the lab, you dissolve 179 g of $MgCl_2$ into 1.00 L of water. Use **Table 14.6** to find the freezing point of the solution.

101. Cooking A cook prepares a solution for boiling by adding 12.5 g of NaCl to a pot holding 0.750 L of water. At what temperature should the solution in the pot boil? Use **Table 14.5** for needed data.

102. The boiling point of ethanol (C_2H_5OH) changes from 78.5°C to 85.2°C when an amount of naphthalene ($C_{10}H_8$) is added to 1.00 kg of ethanol. How much naphthalene, in grams, is required to cause this change? Refer to **Table 14.5** for needed data.

103. Ice Cream A rock salt (NaCl), ice, and water mixture is used to cool milk and cream to make homemade ice cream. How many grams of rock salt must be added to water to lower the freezing point by 10.0°C?

Chapter 14 Assessment

Mixed Review

104. Apply your knowledge of polarity and solubility to predict whether solvation is possible in each situation shown in **Table 14.9.** Explain your answers.

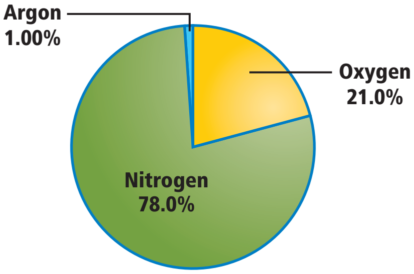
Table 14.9 Is solvation possible?

Solute	Solvent
solid $MgCl_2$	liquid H_2O
liquid NH_3	liquid C_6H_6
gaseous H_2	liquid H_2O
liquid I_2	liquid Br_2

105. Household Paint Some types of paint are colloids composed of pigment particles dispersed in oil. Based on what you know about colloids, recommend an appropriate location for storing cans of leftover household paint. Justify your recommendation.

106. Which solute has the greatest effect on the boiling point of 1.00 kg of water: 50.0 g of strontium chloride ($SrCl_2$) or 150.0 g of carbon tetrachloride (CCl_4)? Justify your answer.

107. Study **Table 14.4.** Analyze solubility and temperature data to determine the general trend followed by the gases (NH_3, CO_2, O_2) in the chart. Compare this trend to the trend followed by most of the solids in the chart. Identify the solids listed that do not follow the general trend followed by most of the solids in the chart.



■ **Figure 14.27**

108. An air sample yields the percent composition shown in **Figure 14.27.** Calculate the mole fraction for each gas present in the sample.

109. If you prepared a saturated aqueous solution of potassium chloride at 25°C and then heated it to 50°C, would you describe the solution as unsaturated, saturated, or supersaturated? Explain.

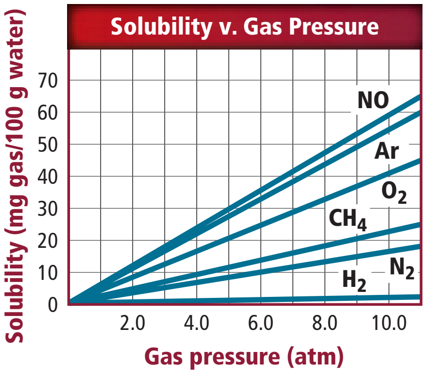
110. How many grams of calcium nitrate ($Ca(NO_3)_2$) would you need to prepare 3.00 L of a 0.500M solution?

111. What would be the molality of the solution described in the previous problem?

Think Critically

112. Develop a plan for making 1000 mL of a 5% by volume solution of hydrochloric acid in water. Your plan should describe the amounts of solute and solvent necessary, as well as the steps involved in making the solution.

113. Compare and Infer Study the phase diagram in **Figure 14.21** on page 501. Compare the dotted lines surrounding ΔT_f and ΔT_b, and describe the differences you observe. How might these lines be positioned differently for solutions of electrolytes and nonelectrolytes? Why?



■ **Figure 14.28**

114. Extrapolate The solubility of argon in water at various pressures is shown in **Figure 14.28.** Extrapolate the data to 15 atm. Use Henry's law to verify the solubility determined by your extrapolation.

115. Infer Dehydration occurs when more fluid is lost from the body than is taken in. Scuba divers are advised to hydrate their bodies before diving. Use your knowledge of the relationship between pressure and gas solubility to explain the importance of hydration prior to a dive.

116. Graph **Table 14.10** shows solubility data that was collected in an experiment. Plot a graph of the molarity of KI versus temperature. What is the solubility of KI at 55°C?

Table 14.10 Solubility of KI

Temperature (°C)	Grams of KI per 100.0 g Solution
20	144
40	162
60	176
80	192
100	206

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117. Design an Experiment You are given a sample of a solid solute and three aqueous solutions containing that solute. How would you determine which solution is saturated, unsaturated, and supersaturated?

118. Compare Which of the following solutions has the highest concentration? Rank the solutions from the greatest to the smallest boiling point depression. Explain your answer.
- **a.** 0.10 mol NaBr in 100.0 mL solution
- **b.** 2.1 mol KOH in 1.00 L solution
- **c.** 1.2 mol $KMnO_4$ in 3.00 L solution

Challenge Problems

119. Interpret the solubility data in **Table 14.11** using the concept of Henry's law.

Table 14.11 Measurements of Solubility of a Gas

Measurement	Solubility
1	0.225
2	0.45
3	0.9
4	1.8
5	3.6

120. You have a solution containing 135.2 g of dissolved KBr in 2.3 L of water. What volume of this solution, in mL, would you use to make 1.5 L of a 0.1*M* KBr solution? What is the boiling point of this new solution?

Cumulative Review

121. The radius of an argon atom is 94 pm. Assuming the atom is spherical, what is the volume of an argon atom in cubic nanometers? $V = 4/3\pi r^3$ (*Chapter 3*)

122. Identify which molecule is polar. (*Chapter 8*)
- **a.** SiH_4
- **b.** NO_2
- **c.** H_2S
- **d.** NCl_3

123. Name the following compounds. (*Chapter 7*)
- **a.** NaBr
- **b.** $Pb(CH_3COO)_2$
- **c.** $(NH_4)_2CO_3$

124. A 12.0-g sample of an element contains 5.94×10^{22} atoms. What is the unknown element? (*Chapter 10*)

125. Pure bismuth can be produced by the reaction of bismuth oxide with carbon at high temperatures.

$$2Bi_2O_3 + 3C \longrightarrow 4Bi + 3CO_2$$

How many moles of Bi_2O_3 reacted to produce 12.6 mol of CO_2? (*Chapter 11*)

Additional Assessment

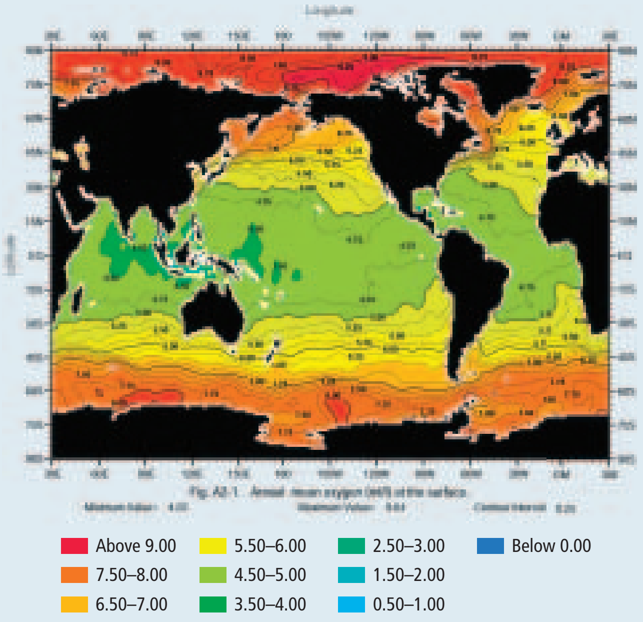
WRITING in Chemistry

126. Homogenized Milk The first homogenized milk was sold in the United States around 1919. Today, almost all milk sold in this country is homogenized in the form of a colloidal emulsion. Research the homogenization process. Write a brief article describing the process. The article should include a flowchart or diagram of the process, as well as a discussion of the reputed benefits and drawbacks associated with drinking homogenized milk.

DBQ Document-Based Questions

Annual Mean Dissolved Oxygen *The data in* **Figure 14.29** *shows the average dissolved oxygen values, in milliliters per liter, in ocean-surface waters during a one-month period in 2001. Longitude is indicated horizontally, and latitude is indicated vertically.*

Data obtained from: National Oceanographic Data Center. 2002. *World Ocean Atlas 2001 Figures.*



■ **Figure 14.29**

127. Are dissolved oxygen values most closely related to latitude or longitude? Why do you think this is true?

128. At what latitude are average dissolved oxygen values the lowest?

129. Describe the general trend defined by the data. Relate the trend to the relationship between gas solubility and temperature.

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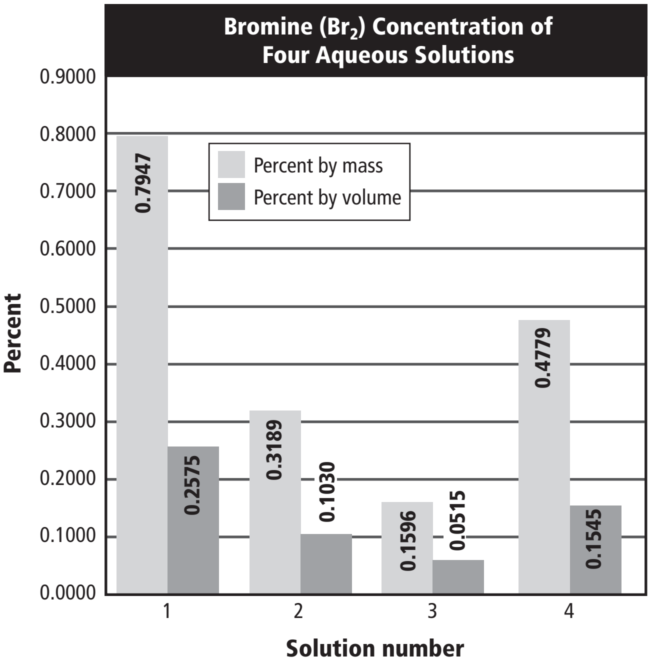
Cumulative
Standardized Test Practice

Multiple Choice

Use the graph below to answer Questions 1 and 2.



1. What is the volume of bromine (Br_2) in 7.000 L of Solution 1?
   - **A.** 55.63 mL
   - **B.** 8.808 mL
   - **C.** 18.03 mL
   - **D.** 27.18 mL

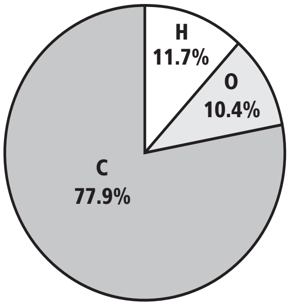
2. How many grams of Br_2 are in 55.00 g of Solution 4?
   - **A.** 3.560 g
   - **B.** 0.084 98 g
   - **C.** 1.151 g
   - **D.** 0.2628 g

3. Which is an intensive physical property?
   - **A.** volume
   - **B.** length
   - **C.** hardness
   - **D.** mass

4. What is the product of this synthesis reaction?
   $Cl_2(g) + 2NO(g) \longrightarrow$?
   - **A.** NCl_2
   - **B.** 2NOCl
   - **C.** N_2O_2
   - **D.** 2ClO

5. If 1 mol of each of the solutes listed below is dissolved in 1 L of water, which solute will have the greatest effect on the vapor pressure of its respective solution?
   - **A.** KBr
   - **B.** $C_6H_{12}O_6$
   - **C.** $MgCl_2$
   - **D.** $CaSO_4$

Use the diagram below to answer Question 6.



6. What is the empirical formula for this substance?
   - **A.** CH_2O
   - **B.** C_8HO
   - **C.** $C_{10}H_{18}O$
   - **D.** $C_7H_{12}O$

7. What is the correct chemical formula for the ionic compound formed by the calcium ion (Ca^{2+}) and the acetate ion ($C_2H_3O_2^-$)?
   - **A.** $CaC_2H_3O_2$
   - **B.** $CaC_4H_6O_3$
   - **C.** $(Ca)_2C_2H_3O_2$
   - **D.** $Ca(C_2H_3O_2)_2$

Use the reaction below to answer Questions 8 and 9.

$$Fe_3O_4\ (s) + 4H_2\ (g) \longrightarrow 3\ Fe\ (s) + 4\ H_2O\ (l)$$

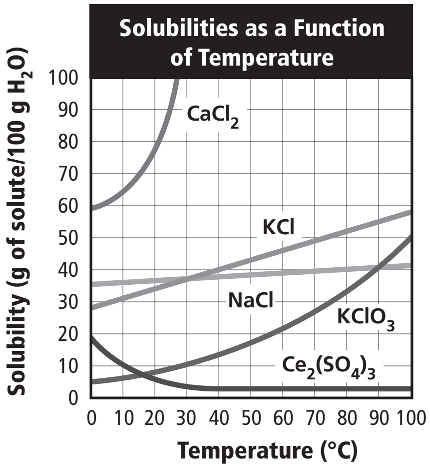
8. If 16 mol of H_2 are used, how many moles of Fe will be produced?
   - **A.** 6
   - **B.** 3
   - **C.** 12
   - **D.** 9

9. If 7 mol of Fe_3O_4 are mixed with 30 mol of H_2, what will be true?
   - **A.** There will be no reactants left.
   - **B.** 2 mol of hydrogen gas will be left over.
   - **C.** 30 mol of water will be produced.
   - **D.** 7 mol of Fe will be produced.

10. What is the molar mass of Fe_3O_4?
    - **A.** 231.56 g/mol
    - **B.** 71.85 g/mol
    - **C.** 287.40 g/mol
    - **D.** 215.56 g/mol

Short Answer

Use the graph below to answer Questions 11 to 13.



11. How many moles of $KClO_3$ can be dissolved in 100 g of water at 60°C?

12. Which can hold more solute at 20°C: NaCl or KCl? How does this compare to their solubilities at 80°C?

13. How many moles of $KClO_3$ would be required to make 1 L of a saturated solution of $KClO_3$ at 75°C?

Extended Response

Use information below to answer Questions 14 and 15.

The electron configuration for silicon is $1s^22s^22p^63s^23p^2$.

14. Explain how this configuration demonstrates the Aufbau principle.

15. Draw the orbital diagram for silicon. Explain how Hund's rule and the Pauli exclusion principle are used in constructing the orbital diagram.

SAT Subject Test: Chemistry

16. What volume of a 0.125M $NiCl_2$ solution contains 3.25 g of $NiCl_2$?

A. 406 mL
B. 32.5 mL
C. 38.5 mL
D. 26.0 mL
E. 201 mL

17. Which is NOT a colligative property?

A. boiling point elevation
B. freezing point depression
C. vapor pressure increase
D. osmotic pressure
E. heat of solution

Use the data table below to answer Questions 18 and 19.

Electronegativities of Selected Elements

H						
2.20						
Li	Be	B	C	N	O	F
0.98	1.57	2.04	2.55	3.04	3.44	3.98
Na	Mg	Al	Si	P	S	Cl
0.93	1.31	1.61	1.90	2.19	2.58	3.16

18. What is the electronegativity difference in Li_2O?

A. 1.48
B. 2.46
C. 3.4
D. 4.42
E. 5.19

19. Which bond has the greatest polarity?

A. C–H
B. Si–O
C. Mg–Cl
D. Al–N
E. H–Cl

NEED EXTRA HELP?																			
If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19
Review Section . . .	14.2	14.2	3.1	9.2	14.4	10.4	7.3	11.2	11.3	10.3	14.3	14.3	14.3	5.3	5.3	14.2	14.4	8.5	8.5