16 Solutions

INSIDE:

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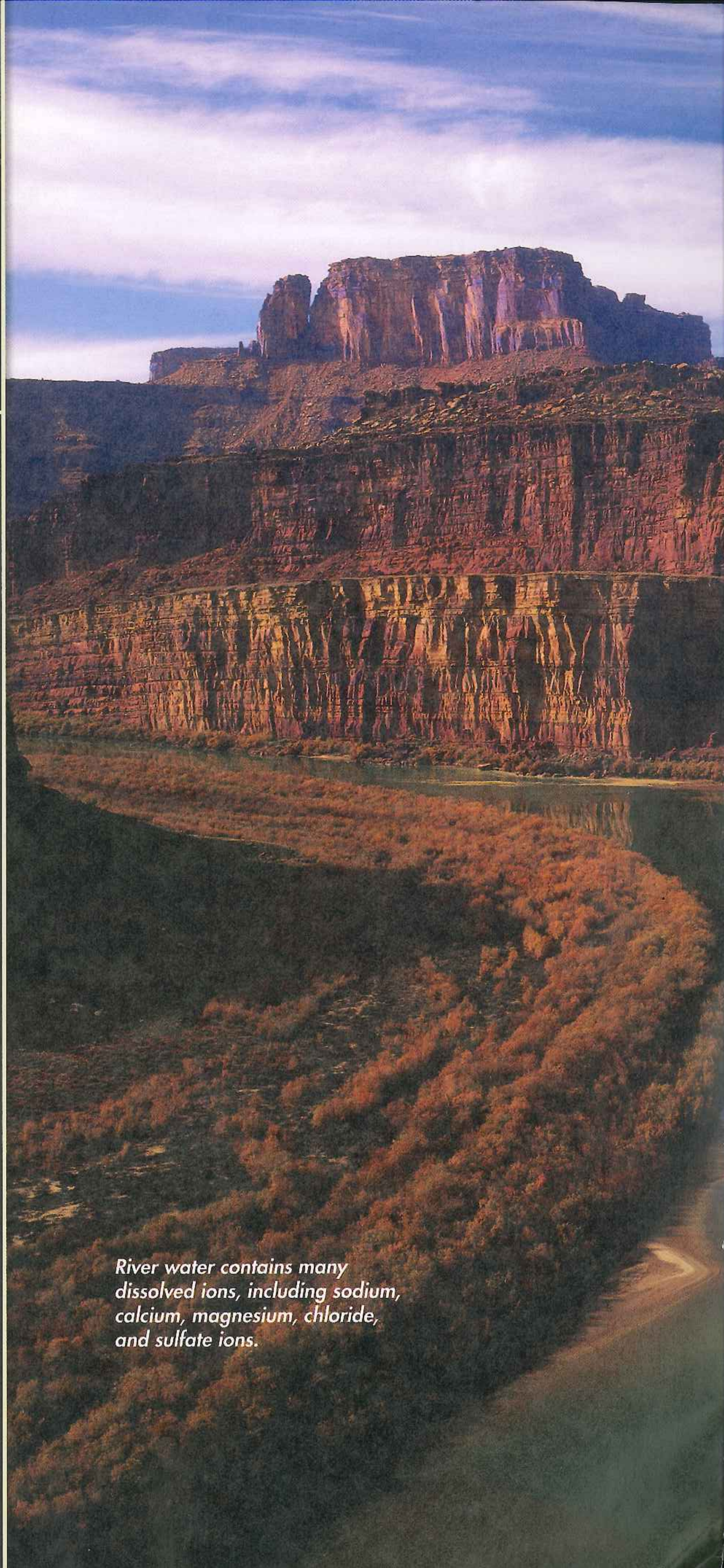
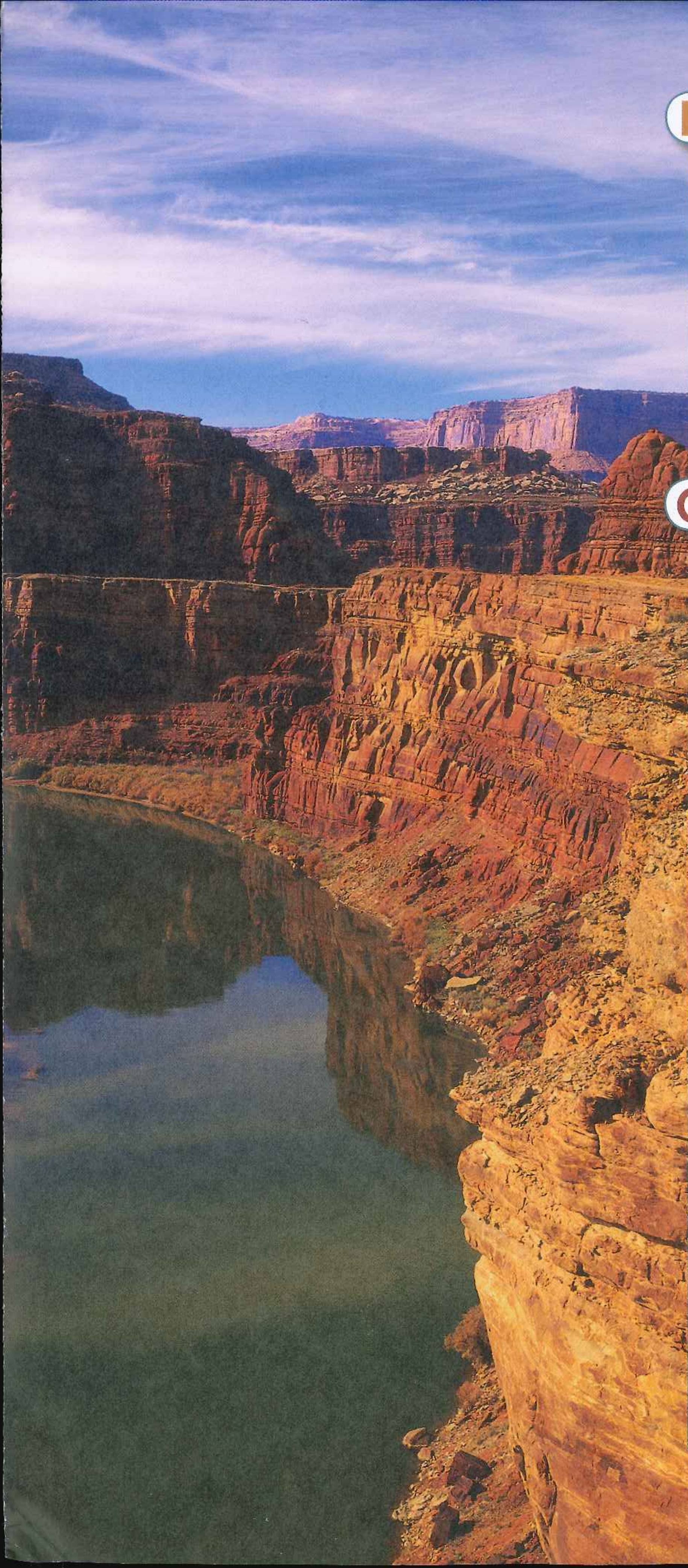


River water contains many dissolved ions, including sodium, calcium, magnesium, chloride, and sulfate ions.



BIGIDEA

THE MOLE AND QUANTIFYING MATTER

Essential Questions:

1. *What properties are used to describe the nature of solutions?*
2. *In what ways can you quantify the concentration of a solution?*

CHEMYSTERY

That Sinking Feeling

It was there one minute and gone the next—a stretch of road that looked like it had been swallowed up by the ground beneath it. The pavement had fallen into a sinkhole caused by moving groundwater. Groundwater consists of rain and melted snow that has soaked into the ground. Over time, groundwater can dissolve huge amounts of rock, gradually forming beautiful limestone caves, as well as destructive sinkholes. One recorded sinkhole swallowed a house, several other buildings, five cars, and a swimming pool!

What factors do you think contribute to the formation of sinkholes? How do you think sinkhole accidents can be prevented?

▶ Connect to the **BIG**IDEA As you read about solutions, try to identify the factors that cause sinkholes to form.

NATIONAL SCIENCE EDUCATION STANDARDS

A-1, A-2, B-5, E-2, F-4, F-6

16.1 Properties of Solutions

CHEMISTRY & YOU

Q: *How can you grow a tree made out of crystals?* You're already familiar with the concept of liquids freezing. But what about crystals growing from a solution? The crystallization of a solute from solution is a physical change that is different from freezing. The crystal tree shown here began its "life" as an ordinary aqueous solution. The tree trunk, made of absorbent paper, soaks up the liquid. As water evaporates from the solution, the solutes crystallize onto the paper, forming delicate "leaves." Not all solutions will crystallize as this one did. The rate of crystallization depends on the nature of the solute and solvent, as well as on the temperature and humidity of the surroundings.

Key Questions

What factors affect how fast a substance dissolves?

How can you describe the equilibrium in a saturated solution?

What factors affect the solubility of a substance?

Vocabulary

- saturated solution
- solubility
- unsaturated solution
- miscible
- immiscible
- supersaturated solution
- Henry's law

Solution Formation

What factors affect how fast a substance dissolves?

Have you noticed, when making tea, that granulated sugar dissolves faster than sugar cubes, and that both granulated sugar and sugar cubes dissolve faster in hot tea or when you stir the tea? Figure 16.1 illustrates these observations. You will be able to explain these observations once you have gained an understanding of the properties of solutions.

Recall that solutions are homogeneous mixtures that may be solid, liquid, or gaseous. The compositions of the solvent and the solute determine whether or not a substance will dissolve. **Factors that affect how fast a substance dissolves include agitation, temperature, and the particle size of the solute.** Each of these factors involves the contact of the solute with the solvent.

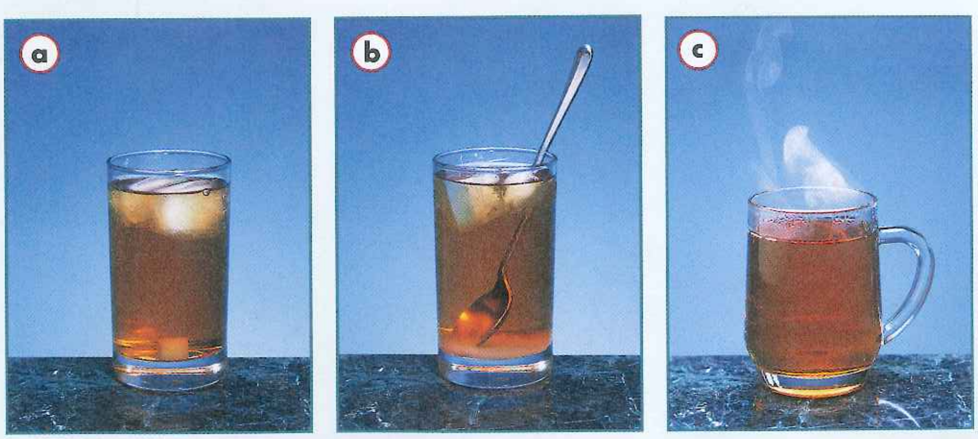


Figure 16.1 Dissolving Sugar
Stirring and heating increase the rate at which a solute dissolves. **a.** A cube of sugar in cold tea dissolves slowly. **b.** Granulated sugar dissolves in cold water more quickly than a sugar cube, especially with stirring. **c.** Granulated sugar dissolves very quickly in hot tea.

Agitation If a teaspoon of granulated sugar (sucrose) is placed in a glass of tea, the crystals dissolve slowly. If the contents of the glass are stirred, however, the crystals dissolve more quickly. The dissolving process occurs at the surface of the sugar crystals. Stirring speeds up the process because fresh solvent (the water in tea) is continually brought into contact with the surface of the solute (sugar). It's important to realize, however, that agitation (stirring or shaking) affects only the rate at which a solid solute dissolves. It does not influence the amount of solute that will dissolve. An insoluble substance remains undissolved regardless of how vigorously or for how long the solvent/solute system is agitated.

Temperature Temperature also influences the rate at which a solute dissolves. Sugar dissolves much more rapidly in hot tea than in iced tea. At higher temperatures, the kinetic energy of water molecules is greater than at lower temperatures, so the molecules move faster. The more rapid motion of the solvent molecules leads to an increase in the frequency and the force of the collisions between water molecules and the surfaces of the sugar crystals.

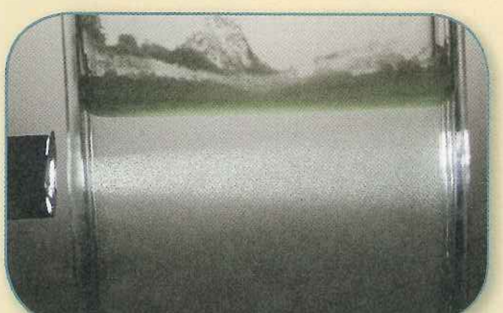
Particle Size of the Solute The rate at which a solute dissolves also depends upon the size of the solute particles. A spoonful of granulated sugar dissolves more quickly than a sugar cube because the smaller particles in granulated sugar expose a much greater surface area to the colliding water molecules. Remember, the dissolving process is a surface phenomenon. The more surface area of the solute that is exposed, the faster the rate of dissolving.

Quick Lab

Purpose To classify mixtures as solutions or colloids using the Tyndall effect

Materials

- sodium hydrogen carbonate
- cornstarch
- stirring rod
- distilled water (or tap water)
- flashlight
- masking tape
- 3 jars with parallel sides
- teaspoon
- cup



Solutions and Colloids

Procedure

1. In a cup, make a paste: Mix one-half teaspoon of cornstarch with 4 teaspoons of water.

2. Fill one jar with water. Add one-half teaspoon of sodium hydrogen carbonate to a second jar and fill with water. Stir to mix. Add the cornstarch paste to the third jar and fill with water. Stir to mix.

3. Turn out the lights in the room. Shine the beam of light from the flashlight at each of the jars and record your observations.

Analyze and Conclude

1. Observe In which of the jars in the experiment was it possible to see the path of the beam of light?

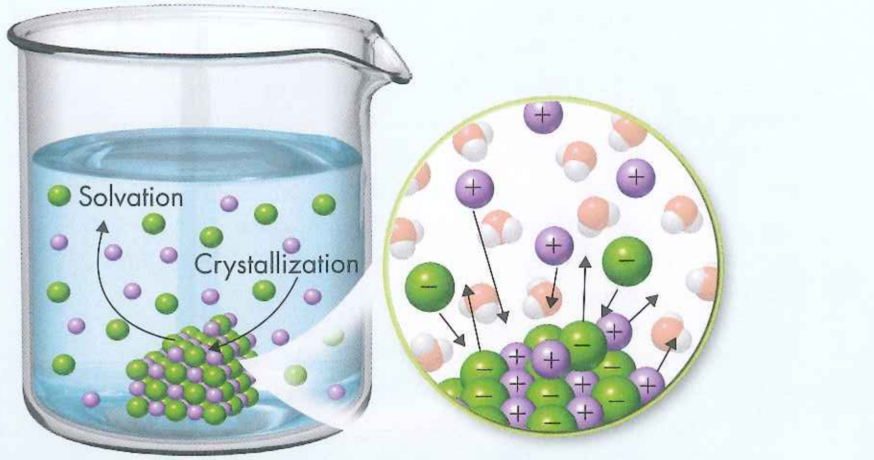
2. Infer What made the light beam visible?

3. Explain If a system that made the light beam visible were filtered, would the light beam be visible in the filtrate? Explain your answer.

4. Predict What would you observe if you were to replace the sodium hydrogen carbonate with sucrose (cane sugar) or sodium chloride (table salt)? If you were to replace the cornstarch with flour or diluted milk?

Figure 16.2 Saturated Solution
In a saturated solution, a state of dynamic equilibrium exists between the solution and the excess solute. The rate of solvation (dissolving) equals the rate of crystallization, so the total amount of dissolved solute remains constant.
Predict ***What would happen if you added more solute to this saturated solution?***

KINETIC ART ***See a saturated solution animated online.***



Solubility

How can you describe the equilibrium in a saturated solution?

If you add 36.0 g of sodium chloride to 100 g of water at 25°C, all of the 36.0 g of salt dissolves. But if you add one more gram of salt and stir, no matter how vigorously or for how long, only 0.2 g of the last portion will dissolve. Why does the remaining 0.8 g of salt remain undissolved? According to the kinetic theory, water molecules are in continuous motion. Therefore, they should continue to bombard the excess solid, solvating and removing the ions. As ions are solvated, they dissolve in the water. Based on this information, you might expect all of the sodium chloride to dissolve eventually. That does not happen, however, because an exchange process is occurring. New particles from the solid are solvated and enter into solution, as shown in Figure 16.2. At the same time, an equal number of already-dissolved particles crystallize. These particles come out of solution and are deposited as a solid. The mass of undissolved crystals remains constant.

What is happening in Figure 16.2? Particles move from the solid into the solution. Some dissolved particles move from the solution back to the solid. Because these two processes occur at the same rate, no net change occurs in the overall system. Such a solution is said to be saturated. A **saturated solution** contains the maximum amount of solute for a given quantity of solvent at a constant temperature and pressure. **In a saturated solution, a state of dynamic equilibrium exists between the solution and any undissolved solute, provided that the temperature remains constant.** At 25°C, 36.2 g of sodium chloride dissolved in 100 g of water forms a saturated solution. If additional solute is added to this solution, it will not dissolve.

The **solubility** of a substance is the amount of solute that dissolves in a given quantity of a solvent at a specified temperature and pressure to produce a saturated solution. Solubility is often expressed in grams of solute per 100 g of solvent (g/100 g H_2O). Sometimes the solubility of a gas is expressed in grams per liter of solution (g/L). A solution that contains less solute than a saturated solution at a given temperature and pressure is an **unsaturated solution.** If additional solute is added to an unsaturated solution, the solute will dissolve until the solution is saturated.

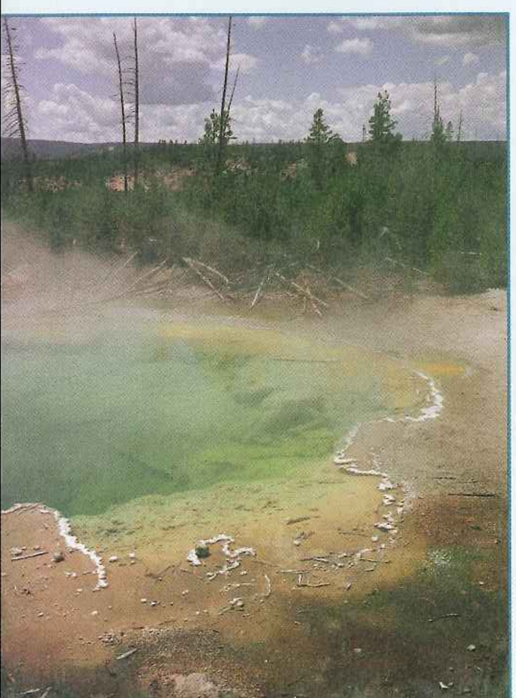


Figure 16.3 Hot Spring
The water in this hot spring in Yellowstone National Park is saturated with minerals. As the water cools near the edges of the spring, some of the minerals crystallize because they are less soluble at the lower temperature.

Some liquids—for example, water and ethanol—are infinitely soluble in each other. Any amount of ethanol will dissolve in a given volume of water, and vice versa. Similarly, ethylene glycol and water mix in all proportions. Pairs of liquids such as these are said to be completely miscible. Two liquids are **miscible** if they dissolve in each other in all proportions. In such a solution, the liquid that is present in the larger amount is usually considered the solvent. Liquids that are slightly soluble in each other—for example, water and diethyl ether—are partially miscible. Liquids that are insoluble in one another are **immiscible.** As you can see in Figure 16.4, oil and water are examples of immiscible liquids.

Figure 16.4 Oil and Water Vegetable oil is not soluble in water. Liquids that are insoluble in one another are immiscible.

Factors Affecting Solubility

What factors affect the solubility of a substance?

You have read that solubility is defined as the mass of solute that dissolves in a given mass of a solvent at a specified temperature. **Temperature affects the solubility of solid, liquid, and gaseous solutes in a solvent; both temperature and pressure affect the solubility of gaseous solutes.**

Temperature The solubility of most solid substances increases as the temperature of the solvent increases. For sodium chloride (NaCl), the increase in solubility is small—from 36.2 g per 100 g of water at 25°C to 39.2 g per 100 g of water at 100°C. Figure 16.5 shows how the solubility of several substances varies with temperature.

For a few substances, solubility decreases with temperature. For example, the solubility of ytterbium sulfate ($Yb_2(SO_4)_3$) in water drops from 44.2 g per 100 g of water at 0°C to 5.8 g per 100 g of water at 90°C. Table 16.1 on the next page lists the solubilities of some common substances in water at various temperatures.

InterpretGraphs

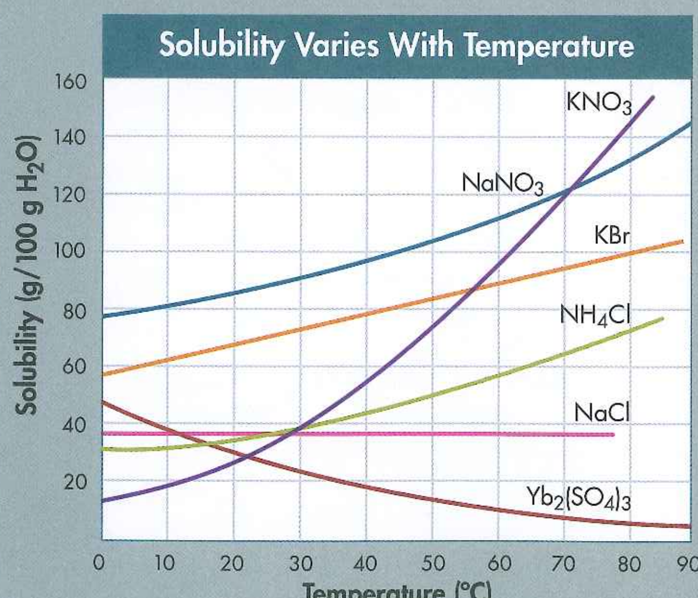


Figure 16.5 Changing the temperature usually affects the solubility of a substance.

a. Read Graphs What happens to the solubility of KNO_3 as the temperature increases?

b. Identify Which substance exhibits the least change in solubility as temperature increases?

c. Predict Suppose you added some solid NaCl to a saturated solution of NaCl at 20°C and warmed the mixture to 40°C. What would happen to the added NaCl?

d. Infer The mineral deposits around the hot spring in Figure 16.3 include NaCl and KCl. How do you think the solubility of KCl changes as the temperature decreases? Explain your answer.

Table 16.1

Solubilities of Substances in Water at Various Temperatures

Substance		Solubility (g/100 g H_2O)			
Name	Formula	0°C	20°C	50°C	100°C
Barium hydroxide	$Ba(OH)_2$	1.67	31.89	—	—
Barium sulfate	$BaSO_4$	0.00019	0.00025	0.00034	—
Calcium hydroxide	$Ca(OH)_2$	0.189	0.173	—	0.07
Lead(II) chloride	$PbCl_2$	0.60	0.99	1.70	—
Lithium carbonate	Li_2CO_3	1.5	1.3	1.1	0.70
Potassium chlorate	$KClO_3$	4.0	7.4	19.3	56.0
Potassium chloride	KCl	27.6	34.0	42.6	57.6
Sodium chloride	$NaCl$	35.7	36.0	37.0	39.2
Sodium nitrate	$NaNO_3$	74	88.0	114.0	182
Aluminum chloride	$AlCl_3$	30.84	31.03	31.60	33.32
Silver nitrate	$AgNO_3$	122	222.0	455.0	733
Lithium bromide	$LiBr$	143.0	166	203	266.0
Sucrose (table sugar)	$C_{12}H_{22}O_{11}$	179	230.9	260.4	487
Hydrogen*	H_2	0.00019	0.00016	0.00013	0.0
Oxygen*	O_2	0.0070	0.0043	0.0026	0.0
Carbon dioxide*	CO_2	0.335	0.169	0.076	0.0

*Gas at 101 kPa (1 atm) total pressure

Suppose you make a saturated solution of sodium ethanoate (sodium acetate) at 30°C and let the solution stand undisturbed as it cools to 25°C. Because the solubility of this compound is greater at 30°C than at 25°C, you expect that solid sodium ethanoate will crystallize from the solution as the temperature drops. But no crystals form. You have made a supersaturated solution. A **supersaturated solution** contains more solute than it can theoretically hold at a given temperature. The crystallization of a supersaturated solution can be initiated if a very small crystal, called a seed crystal, of the solute is added. The rate at which excess solute deposits upon the surface of a seed crystal can be very rapid, as shown in Figure 16.6. Crystallization can also occur if the inside of the container is scratched.

Another example of crystallization in a supersaturated solution is the production of rock candy. A solution is supersaturated with sugar. Seed crystals cause the sugar to crystallize out of solution onto a string for you to enjoy.

CHEMISTRY & YOU

Q: *How do you think crystal growing kits work? Use what you know about solubility and saturated solutions to explain your answer.*

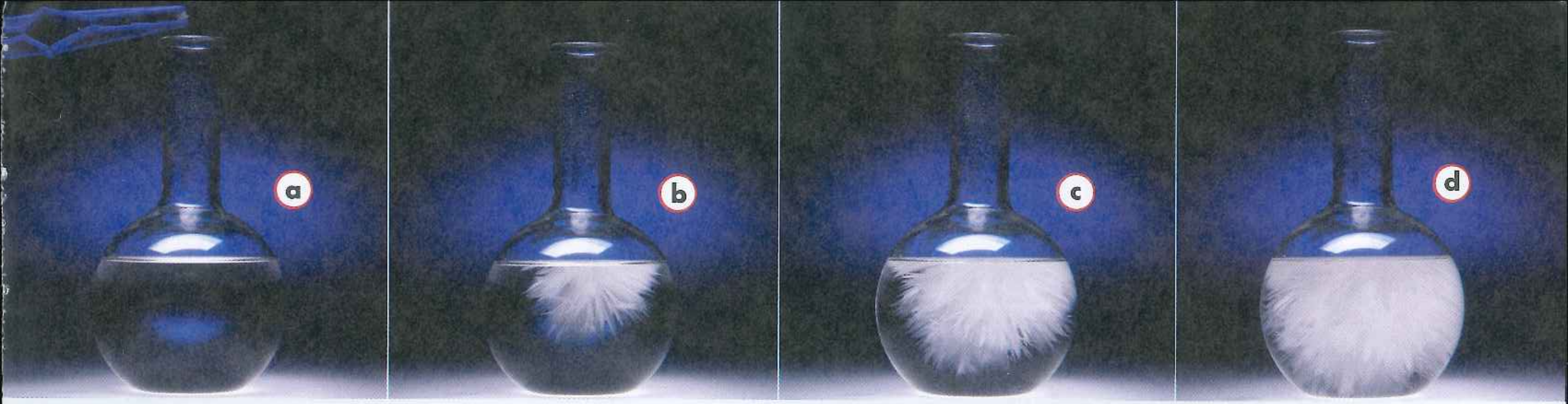


Figure 16.6 Supersaturated Solution
A supersaturated solution of sodium ethanoate ($NaC_2H_3O_2(aq)$) crystallizes rapidly when disturbed.
a. The solution is clear before a seed crystal is added. **b.** Crystals begin to form in the solution immediately after the addition of a seed crystal. **c–d.** The excess solute crystallizes rapidly.
Infer ***When the crystallization has ceased, will the solution be saturated or unsaturated?***

The effect of temperature on the solubility of gases in liquid solvents is opposite that of solids. The solubilities of most gases are greater in cold water than in hot. For example, Table 16.1 shows that the most important component of air for living beings—oxygen—becomes less soluble in water as the temperature of the solution rises. This fact has some important consequences. When an industrial plant takes water from a lake to use for cooling and then dumps the resulting heated water back into the lake, the temperature of the entire lake increases. Such a change in temperature is known as thermal pollution. Aquatic animal and plant life can be severely affected because the increase in temperature lowers the concentration of dissolved oxygen in the lake water.

Pressure Changes in pressure have little effect on the solubility of solids and liquids, but pressure strongly influences the solubility of gases. Gas solubility increases as the partial pressure of the gas above the solution increases. Carbonated beverages are a good example. These drinks contain large amounts of carbon dioxide (CO_2) dissolved in water. Dissolved CO_2 makes the liquid fizz and your mouth tingle. The drinks are bottled under a high pressure of CO_2 gas, which forces large amounts of the gas into solution. When a carbonated-beverage container is opened, the partial pressure of CO_2 above the liquid decreases. Immediately, bubbles of CO_2 form in the liquid and escape from the open bottle, as shown in Figure 16.7. As a result, the concentration of dissolved CO_2 decreases. If the bottle is left open, the drink becomes "flat" as the solution loses most of its CO_2.

How is the partial pressure of carbon dioxide gas related to the solubility of CO_2 in a carbonated beverage? The relationship is described by **Henry's law,** which 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. In other words, as the pressure of the gas above the liquid increases, the solubility of the gas increases. Similarly, as the pressure of the gas decreases, the solubility of the gas decreases. You can write the relationship in the form of an equation.

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

S_1 is the solubility of a gas at one pressure, P_1; S_2 is the solubility at another pressure, P_2.

Figure 16.7 CO_2 in Solution
When a carbonated-beverage bottle is sealed, the pressure of CO_2 above the liquid is high, and the concentration of CO_2 in the liquid is also high. When the cap is removed, the pressure of CO_2 gas above the liquid decreases, and carbon dioxide bubbles out of the liquid.

Sample Problem 16.1

CHEM TUTOR

Using Henry's Law

If the solubility of a gas in water is 0.77 g/L at 3.5 atm of pressure, what is its solubility (in g/L) at 1.0 atm of pressure? (The temperature is held constant at 25°C.)

1 Analyze **List the knowns and the unknown.** Use Henry's law to solve for the unknown solubility.

KNOWNS	UNKNOWN
$P_1 = 3.5\text{ atm}$	$S_2 = ?\text{ g/L}$
$S_1 = 0.77\text{ g/L}$	
$P_2 = 1.0\text{ atm}$	

2 Calculate **Solve for the unknown.**

State the equation for Henry's law.

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

Isolate S_2 by multiplying both sides by P_2:

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

Solve Henry's law for S_2. Substitute the known values and calculate.

$$S_2 = \frac{S_1 \times P_2}{P_1} = \frac{0.77\text{ g/L} \times 1.0\,\cancel{\text{atm}}}{3.5\,\cancel{\text{atm}}} = \boxed{0.22\text{ g/L}}$$

3 Evaluate **Does the result make sense?** The new pressure is approximately one third of the original pressure, so the new solubility should be approximately one third of the original. The answer is correctly expressed to two significant figures.

In Problem 1, you're solving Henry's law for an unknown solubility. In Problem 2, you're solving for an unknown pressure.

1. The solubility of a gas in water is 0.16 g/L at 104 kPa. What is the solubility when the pressure of the gas is increased to 288 kPa? Assume the temperature remains constant.

2. A gas has a solubility in water at 0°C of 3.6 g/L at a pressure of 1.0 atm. What pressure is needed to produce an aqueous solution containing 9.5 g/L of the same gas at 0°C?

16.1 LessonCheck

ONLINE PROBLEMS

3. **Review** What determines how fast a substance will dissolve?

4. **Describe** How can you describe the state of equilibrium in a saturated solution that contains undissolved solute?

5. **Describe** What condition(s) determine the solubilities of solid, liquid, and gaseous solutes in a solvent?

6. **Identify** Name a unit used to express solubility.

7. **Describe** What determines whether or not a substance will dissolve?

8. **Explain** What would you do to change

- **a.** a saturated solid/liquid solution to an unsaturated solution?
- **b.** a saturated gas/liquid solution to an unsaturated solution?

9. **Calculate** The solubility of a gas is 0.58 g/L at a pressure of 104 kPa. What is its solubility if the pressure increases to 250 kPa at the same temperature?

16.2 Concentrations of Solutions

CHEMISTRY & YOU

Q: *How can you describe the concentration of a solution?* Clean drinking water is important for all communities. What constitutes clean water? Your federal and state governments set standards limiting the amount of contaminants allowed in drinking water. These contaminants include metals, pesticides, and bacteria. Water must be tested continually to ensure that the concentrations of these contaminants do not exceed established limits.

Key Questions

How do you calculate the molarity of a solution?

What effect does dilution have on the amount of solute?

How do percent by volume and percent by mass differ?

Vocabulary

- concentration
- dilute solution
- concentrated solution
- molarity (M)

Molarity

How do you calculate the molarity of a solution?

You have learned that a substance can dissolve to some extent in a particular solvent to form a solution. This lesson focuses on ways to express the actual extent of dissolving. The **concentration** of a solution is a measure of the amount of solute that is dissolved in a given quantity of solvent. A solution that contains a relatively small amount of solute is a **dilute solution.** By contrast, a **concentrated solution** contains a large amount of solute. An aqueous solution of sodium chloride containing 1 g NaCl per 100 g H_2O might be described as dilute when compared with a sodium chloride solution containing 30 g NaCl per 100 g H_2O. But the same solution might be described as concentrated when compared with a solution containing only 0.01 g NaCl per 100 g H_2O. You can see that the terms *concentrated* and *dilute* are only qualitative descriptions of the amount of a solute in solution.

How can concentration be expressed quantitatively? In chemistry, the most important unit of concentration is molarity. **Molarity (M)** is the number of moles of solute dissolved in one liter of solution. Molarity is also known as molar concentration. When the symbol M is accompanied by a numerical value, it is read as "molar." Figure 16.8 illustrates the procedure for making a 0.5M, or 0.5-molar, solution. Note that the volume involved is the total volume of the resulting solution, not the volume of the solvent alone.

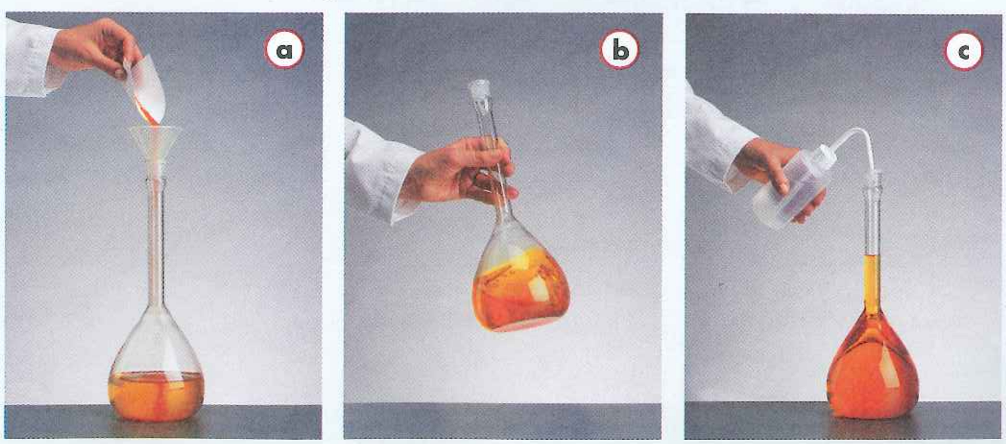


Figure 16.8 How to Make a 0.5M Solution
a. Add 0.5 mol of solute to a 1-L volumetric flask that is half filled with distilled water.
b. Swirl the flask carefully to dissolve the solute.
c. Fill the flask with water exactly to the 1-L mark.

To calculate the molarity of a solution, divide the number of moles of solute by the volume of the solution in liters.

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

For example, suppose 2 mol of glucose are dissolved in 5 L of solution. You would calculate the molarity of the solution as follows:

$$\frac{2 \text{ mol glucose}}{5 \text{ L solution}} = 0.4 \text{ mol/L} = 0.4M$$

If the amount of solute in a solution is expressed in mass units instead of moles, you can calculate molarity by using the appropriate conversion factors, as shown in Sample Problem 16.2.

Sample Problem 16.2

Calculating Molarity

Intravenous (IV) saline solutions are often administered to patients in the hospital. One saline solution contains 0.90 g NaCl in exactly 100 mL of solution. What is the molarity of the solution?

1 Analyze List the knowns and the unknown. Convert the concentration from g/100 mL to mol/L. The sequence is g/100 mL ⟶ mol/100 mL ⟶ mol/L.

KNOWNS

solution concentration = 0.90 g NaCl/100 mL

molar mass of NaCl = 58.5 g/mol

UNKNOWN

solution concentration = ?M

2 Calculate Solve for the unknown.

Use the molar mass to convert g NaCl/100 mL to mol NaCl/100 mL. Then convert the volume units so that your answer is expressed in mol/L.

$$\text{Solution Concentration} = \frac{0.90 \cancel{\text{g NaCl}}}{100 \cancel{\text{mL}}} \times \frac{1 \text{ mol NaCl}}{58.5 \cancel{\text{g NaCl}}} \times \frac{1000 \cancel{\text{mL}}}{1 \text{ L}}$$

$$= 0.15 \text{ mol/L}$$

$$= 0.15M$$

The relationship 1 L = 1000 mL gives you the conversion factor 1000 mL/1 L.

3 Evaluate Does the result make sense? The answer should be less than 1M because a concentration of 0.90 g/100 mL is the same as 9.0 g/1000 mL (9.0 g/1 L), and 9.0 g is less than 1 mol of NaCl. The answer is correctly expressed to two significant figures.

10. A solution has a volume of 2.0 L and contains 36.0 g of glucose ($C_6H_{12}O_6$). If the molar mass of glucose is 180 g/mol, what is the molarity of the solution?

11. A solution has a volume of 250 mL and contains 0.70 mol NaCl. What is its molarity?

In some cases, you may need to determine the number of moles of solute dissolved in a given volume of solution. You can do this if the molarity of the solution is known. For example, how many moles are in 2.00 L of 2.5*M* lithium chloride (LiCl)? First, rearrange the formula for molarity to solve for the number of moles. Then, substitute the known values for molarity and volume.

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

$$\begin{aligned}\text{Moles of solute} &= \text{molarity } (M) \times \text{liters of solution } (V)\\ &= 2.5M \times 2.00\ \text{L} = \left(\frac{2.5\ \text{mol}}{1\ \cancel{\text{L}}}\right) \times 2.00\ \cancel{\text{L}}\\ &= 5.0\ \text{mol}\end{aligned}$$

Thus, 2.00 L of 2.5*M* lithium chloride solution contains 5.0 mol of LiCl.

SampleProblem 16.3

Calculating the Moles of Solute in a Solution

Household laundry bleach is a dilute aqueous solution of sodium hypochlorite (NaClO). How many moles of solute are present in 1.5 L of 0.70*M* NaClO?

1 Analyze List the knowns and the unknown. The conversion is volume of solution ⟶ moles of solute. Molarity has the units mol/L, so you can use it as a conversion factor between moles of solute and volume of solution.

KNOWNS
volume of solution = 1.5 L
solution concentration = 0.70*M* NaClO

UNKNOWN
moles solute = ? mol

2 Calculate Solve for the unknown.

Multiply the given volume by the molarity expressed in mol/L.

$$1.5\ \cancel{\text{L}} \times \frac{0.70\ \text{mol NaClO}}{1\ \cancel{\text{L}}} = 1.1\ \text{mol NaClO}$$

3 Evaluate Does the result make sense? The answer should be greater than 1 mol but less than 1.5 mol, because the solution concentration is less than 0.75 mol/L and the volume is less than 2 L. The answer is correctly expressed to two significant figures.

12. How many moles of ammonium nitrate are in 335 mL of 0.425*M* NH_4NO_3?

13. How many moles of solute are in 250 mL of 2.0*M* $CaCl_2$? How many grams of $CaCl_2$ is this?

Make sure that your volume units cancel when you do these problems. If they don't, then you're probably missing a conversion factor in your calculations.

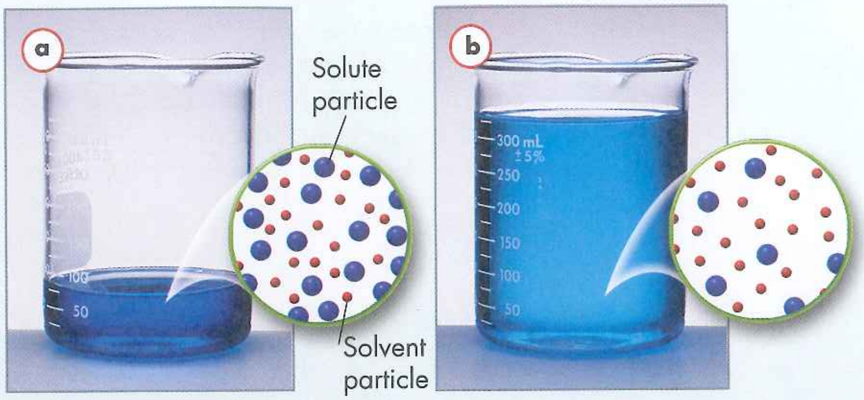


Figure 16.9 Dilution
Adding solvent to a concentrated solution lowers the concentration, but the total number of moles of solute present remains the same.

Making Dilutions

What effect does dilution have on the amount of solute?

Both of the solutions in Figure 16.9 contain the same amount of solute. You can tell by the color of solution (a) that it is more concentrated than the solution (b); that is, solution (a) has the greater molarity. The more dilute solution (b) was made from solution (a) by adding more solvent. **Diluting a solution reduces the number of moles of solute per unit volume, but the total number of moles of solute in solution does not change.** You can also express this concept by writing an equation.

$$\text{Moles of solute before dilution} = \text{moles of solute after dilution}$$

Now recall the definition of molarity and how it can be rearranged to solve for moles of solute.

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

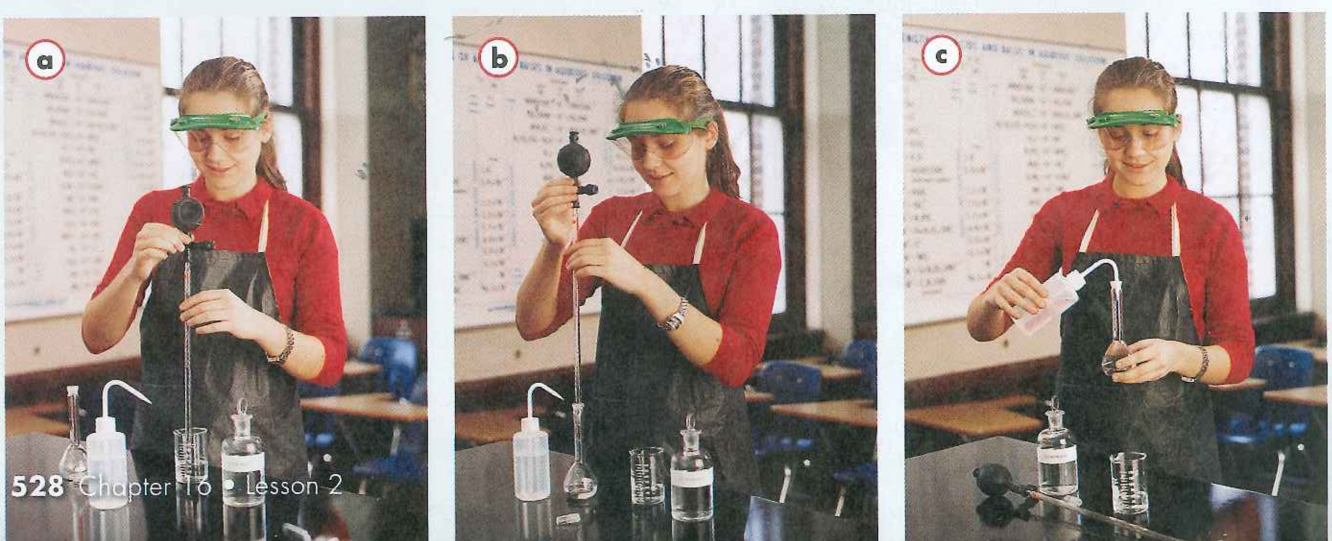
$$\text{Moles of solute} = \text{molarity } (M) \times \text{liters of solution } (V)$$

The total number of moles of solute remains unchanged upon dilution, so you can now write this equation:

$$\text{Moles of solute} = M_1 \times V_1 = M_2 \times V_2$$

M_1 and V_1 are the molarity and volume of the initial solution, and M_2 and V_2 are the molarity and volume of the diluted solution. Volumes can be in liters or milliliters, as long as the same units are used for both V_1 and V_2. Figure 16.10 illustrates the procedure used for making a dilution in the lab.

Figure 16.10 Making a Dilution
The student is preparing 100 mL of 0.40M $MgSO_4$ from a stock solution of 2.0M $MgSO_4$. **a.** She measures 20 mL of the stock solution with a 20-mL pipette. **b.** She transfers the 20 mL to a 100-mL volumetric flask. **c.** She carefully adds water to the mark to make 100 mL of solution.
Compare ***How many moles of $MgSO_4$ are in 20 mL of the stock solution? In 100 mL of the dilute solution?***



CHEM TUTOR

Sample Problem 16.4

Preparing a Dilute Solution

How many milliliters of aqueous $2.00M$ $MgSO_4$ solution must be diluted with water to prepare 100.0 mL of aqueous $0.400M$ $MgSO_4$?

KNOWNS

$M_1 = 2.00M\ MgSO_4$
$M_2 = 0.400M\ MgSO_4$
$V_2 = 100.0\text{mL of } 0.400M\ MgSO_4$

UNKNOWN

$V_1 = ?\ \text{mL of } 2.00M\ MgSO_4$

1 Analyze List the knowns and the unknown. Use the equation $M_1 \times V_1 = M_2 \times V_2$ to solve for the unknown initial volume of solution (V_1) that is diluted with water.

2 Calculate Solve for the unknown.

Solve for V_1 and substitute the known values into the equation.

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.400\cancel{M} \times 100.0\text{ mL}}{2.00\cancel{M}} = \boxed{20.0\text{ mL}}$$

Thus 20.0 mL of the initial solution must be diluted by adding enough water to increase the volume to 100.0 mL.

3 Evaluate Does the result make sense? The initial concentration is five times larger than the dilute concentration. Because the number of moles of solute does not change, the initial volume of solution should be one fifth the final volume of the diluted solution.

14. How many milliliters of a solution of $4.00M$ KI are needed to prepare 250.0 mL of $0.760M$ KI?

15. How could you prepare 250 mL of $0.20M$ NaCl using only a solution of $1.0M$ NaCl and water?

What kind of volume-measuring device would you use to make the dilution in Sample Problem 16.4? The dilution requires a molarity with three significant figures, so you would need to measure 20.0 mL of the $2.00M$ $MgSO_4$ solution with a 20-mL volumetric pipette or a burette. (A graduated cylinder would not provide enough precision.) You would transfer the solution to a 100-mL volumetric flask and add distilled water to the flask exactly up to the etched line. The contents would then be 100.0 mL of $0.400M$ $MgSO_4$.

Percent Solutions

How do percent by volume and percent by mass differ?

If both the solute and the solvent are liquids, a convenient way to make a solution is to measure the volumes of the solute and the solution. The concentration of the solute is then expressed as a percent of the solution by volume. **Percent by volume of a solution is the ratio of the volume of solute to the volume of solution.** For example, isopropyl alcohol (2-propanol) is sold as a 91 percent solution by volume. You could prepare such a solution by diluting 91 mL of pure isopropyl alcohol with enough water to make 100 mL of solution. The concentration is written as 91 percent by volume, 91 percent (volume/volume), or 91% (v/v).

$$\text{Percent by volume } (\%(v/v)) = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

Sample Problem 16.5

Calculating Percent by Volume

What is the percent by volume of ethanol (C_2H_6O, or ethyl alcohol) in the final solution when 85 mL of ethanol is diluted to a volume of 250 mL with water?

KNOWNS

volume of solute = 85 mL ethanol

volume of solution = 250 mL

UNKNOWN

Percent by volume = ? % ethanol (v/v)

1 Analyze List the knowns and the unknown. Use the known values for the volume of solute and volume of solution to calculate percent by volume.

2 Calculate Solve for the unknown.

State the equation for percent by volume.

$$\text{Percent by volume (\% (v/v))} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$$

Substitute the known values into the equation and solve.

$$\text{\% (v/v)} = \frac{85 \cancel{\text{mL}} \text{ ethanol}}{250 \cancel{\text{mL}}} \times 100\%$$

$$= \boxed{34\% \text{ ethanol (v/v)}}$$

3 Evaluate Does the result make sense? The volume of the solute is about one third the volume of the solution, so the answer is reasonable. The answer is correctly expressed to two significant figures.

16. If 10 mL of propanone (C_3H_6O, or acetone) is diluted with water to a total solution volume of 200 mL, what is the percent by volume of propanone in the solution?

17. A bottle of the antiseptic hydrogen peroxide (H_2O_2) is labeled 3.0% (v/v). How many mL H_2O_2 are in a 400.0-mL bottle of this solution?

Another way to express the concentration of a solution is as a percent by mass, or percent (mass/mass). **Percent by mass of a solution is the ratio of the mass of the solute to the mass of the solution.**

$$\text{Percent by mass (\%(m/m))} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

You can also define percent by mass as the number of grams of solute per 100 g of solution. Percent by mass is sometimes a convenient measure of concentration when the solute is a solid. For example, a solution containing 7 g of sodium chloride in 100 grams of solution has a concentration of 7 percent by mass—also written as 7 percent (mass/mass) or 7% (m/m).

You have probably seen information on food labels expressed as a percent composition. For example, the label on a fruit-flavored drink often indicates the "percent juice" contained in the product. Such information can be misleading unless the units are given. When you describe percent solutions, be sure to specify whether the concentration is % (v/v) or % (m/m).

CHEMISTRY & YOU

Q: *What are three ways to calculate the concentration of a solution?*

Sample Problem 16.6

Using Percent by Mass as a Conversion Factor

How many grams of glucose ($C_6H_{12}O_6$) are needed to make 2000 g of a 2.8% glucose (m/m) solution?

1 Analyze List the knowns and the unknown. The conversion is mass of solution $\longrightarrow$ mass of solute. In a 2.8% $C_6H_{12}O_6$ (m/m) solution, each 100 g of solution contains 2.8 g of glucose. Used as a conversion factor, the concentration allows you to convert g of solution to g of $C_6H_{12}O_6$.

KNOWNS

mass of solution = 2000 g

percent by mass = 2.8% $C_6H_{12}O_6$ (m/m)

UNKNOWN

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

2 Calculate Solve for the unknown.

Write the percent by mass as a conversion factor with g $C_6H_{12}O_6$ in the numerator.

$$\frac{2.8\text{ g } C_6H_{12}O_6}{100\text{ g solution}}$$

Multiply the mass of the solution by the conversion factor.

$$2000\ \cancel{\text{g solution}} \times \frac{2.8\text{ g } C_6H_{12}O_6}{100\ \cancel{\text{g solution}}} = \boxed{56\text{ g } C_6H_{12}O_6}$$

3 Evaluate Does the result make sense? The prepared mass of the solution is 20×100 g. Since a 100-g sample of 2.8% (m/m) solution contains 2.8 g of solute, you need $20 \times 2.8\text{ g} = 56$ g of solute. To make the solution, mix 56 g of $C_6H_{12}O_6$ with 1944 g of solvent. (56 g of solute + 1944 g of solvent = 2000 g of solution)

18. Calculate the grams of solute required to make 250 g of 0.10% $MgSO_4$ (m/m).

You can solve this problem by using either dimensional analysis or algebra.

16.2 LessonCheck

19. **Review** How do you calculate the molarity of a solution?

20. **Compare** How does the number of moles of solute before a dilution compare with the number of moles of solute after the dilution?

21. **Identify** What are two ways of expressing the concentration of a solution as a percent?

22. **Calculate** What is the molarity of a solution containing 400 g $CuSO_4$ in 4.00 L of solution?

23. **Calculate** How many milliliters of a stock solution of 2.00M KNO_3 would you need to prepare 100.0 mL of 0.150M KNO_3?

24. **Calculate** How many moles of solute are present in 50.0 mL of 0.20M KNO_3?

25. **Calculate** What is the concentration, in percent (v/v), of a solution containing 50 mL of diethyl ether ($C_4H_{10}O$) in 2.5 L of solution?

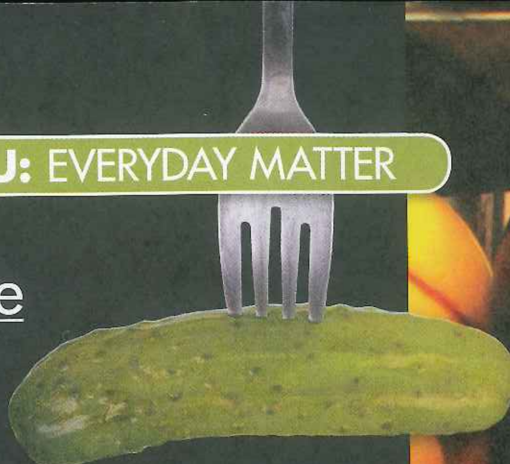
26. **Calculate** What mass of K_2SO_4 would you need to prepare 1500 g of 5.0% K_2SO_4 (m/m) solution?

BIGIDEA
THE MOLE AND QUANTIFYING MATTER

27. What information would you need in order to convert molarity to percent by volume?

CHEMISTRY & YOU: EVERYDAY MATTER

Art of the Pickle



Every culture has its own version of the pickle, and the art of making pickles dates back to ancient history. The earliest known pickles were produced more than 4000 years ago using cucumbers native to India. In Korea, kimchi (pickled cabbage) has been produced for more than 3000 years. Ancient Egyptians and Greeks wrote about the nutritive value and healing power of pickles.

Nowadays, you can often count on the savory taste of a pickle when you order food at a restaurant or from a street vendor. Pickles come in many varieties. You might already know the taste of the dill pickles that accompany sandwiches, or the relish found at a hot dog stand. Other kinds of pickles include Japanese pickled ginger (often served with sushi), European pickled herring, and the hot pickled peppers found in some Mexican salsas.

Pickling is a way of preserving food using a solution of salt, acid (usually vinegar), spices, and/or sugar. Soaking vegetables and meats in the pickling solution prevents the growth of harmful bacteria and imparts a tangy, savory flavor to the food.

Main Types of Pickling

Pickling in Acid Solution

Process This type of pickling typically involves immersing and/or cooking vegetables for a few hours or days in a solution containing vinegar and spices. Vinegar is a 5% (v/v) aqueous solution of acetic acid ($C_2H_4O_2$).

Examples Acid solutions are used to prepare pickled ginger (above), pickled beets, bread-and-butter pickles, pickled herring, and hot dog relish.

Pickling in Brine Solution

Process In brine-based pickling, food is soaked in a salt solution called brine for 4 to 6 weeks. The brine encourages the growth of acid-producing bacteria. The concentration of the brine is usually 5–10% NaCl (m/m).

Examples Common brine pickles include dill cucumber pickles, sauerkraut (pickled fermented cabbage), Middle Eastern pickled turnips, and Indian mango pickles and chutneys.

SALTY SOLUTIONS As pickles soak, the salt in the solution encourages the growth of acid-producing bacteria, which prevent the growth of bacteria that cause spoilage. The salt concentration needs to be as exact as possible. If it's too high, the acid producers can't flourish; if it's too low, then not enough acid producers will grow, and your pickles may spoil.

Take It Further

1. Calculate A pickle recipe calls for 1.5 cups of pickling salt (NaCl) to be mixed with 1 gallon of water. If 1 cup of pickling salt has a mass of 220 grams, what is the percent by mass of NaCl in the pickling solution? (1 gal = 3785 cm^3)

2. Evaluate Is enough information given in the previous question to calculate the percent by volume of NaCl in the pickling solution? Explain your answer.

3. Calculate How many grams of pickling salt would you need to prepare 3500 g of 5.0% NaCl (m/m) solution for a batch of cucumber pickles?

16.3 Colligative Properties of Solutions

CHEMISTRY & YOU

Q: *Why do you need salt to make ice cream?* Here's a hint—it's *not* because ice cream is supposed to taste salty. Temperatures below 0°C are needed to make ice cream. Ice-cream makers know that if you add rock salt to ice, the mixture freezes at a few degrees below 0°C. In this lesson, you will discover how a solute can change the freezing point of a solution.

Key Question

What are three colligative properties of solutions?

Vocabulary

- colligative property
- freezing-point depression
- boiling-point elevation

Describing Colligative Properties

What are three colligative properties of solutions?

You already know that the physical properties of a solution differ from those of the pure solvent used to make the solution. After all, tea is not the same as pure water. But it might surprise you to learn that some of these differences in properties have little to do with the specific identity of the solute. Instead, they depend upon the mere presence of solute particles in the solution.

A **colligative property** is a property of solutions that depends only upon the number of solute particles, not upon their identity. **Three important colligative properties of solutions are vapor-pressure lowering, freezing-point depression, and boiling-point elevation.**

Vapor-Pressure Lowering Recall that vapor pressure is the pressure exerted by a vapor that is in dynamic equilibrium with its liquid in a closed system. A solution that contains a solute that is nonvolatile (not easily vaporized) always has a lower vapor pressure than the pure solvent, as shown in Figure 16.11. Glucose, a molecular compound, and sodium chloride, an ionic compound, are examples of nonvolatile solutes. When glucose or sodium chloride is dissolved in a solvent, the vapor pressure of the solution is lower than the vapor pressure of the pure solvent. Why is this true?

KINETIC ART See vapor pressure animated online.

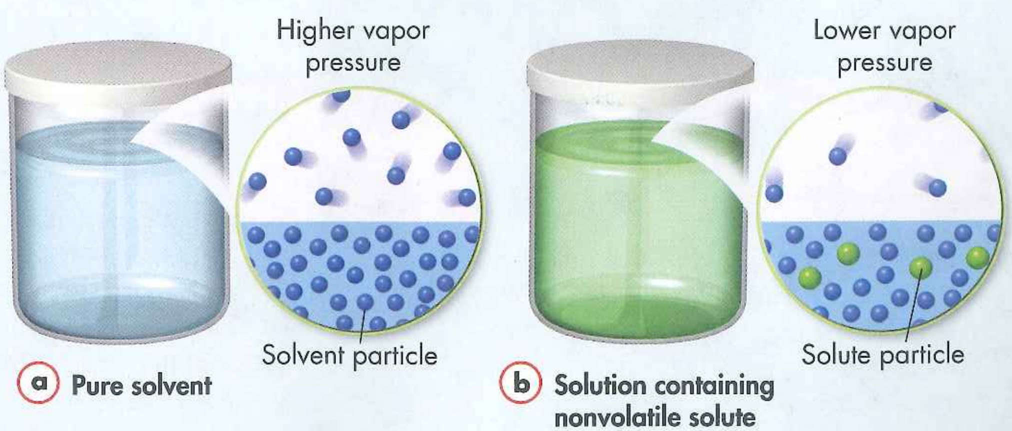
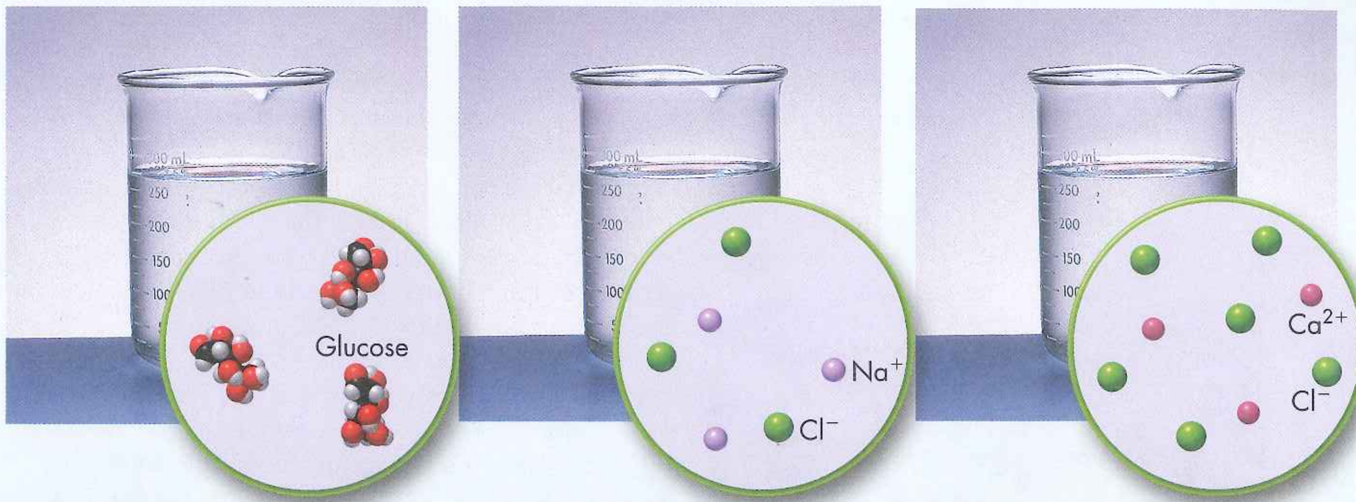


Figure 16.11 Vapor Pressure The vapor pressure of a solution of a nonvolatile solute is less than the vapor pressure of a pure solvent. **a.** Equilibrium is established between the liquid and vapor in a pure solvent. **b.** In a solution, solute particles reduce the number of solvent particles able to escape the liquid. Equilibrium is established at a lower vapor pressure.

Figure 16.12 Molecular vs. Ionic Solutes
Particle concentrations differ for dissolved molecular and ionic compounds in water.
Compare ***Which solution has the lowest vapor pressure? The highest?***



Glucose in Solution
Three moles of glucose dissolved in water produce 3 mol of particles because glucose does not dissociate.

Sodium Chloride in Solution
Three moles of sodium chloride dissolved in water produce 6 mol of particles because each formula unit of NaCl dissociates into two ions.

Calcium Chloride in Solution
Three moles of calcium chloride dissolved in water produce 9 mol of particles because each formula unit of $CaCl_2$ dissociates into three ions.

In an aqueous solution of sodium chloride, sodium ions and chloride ions are dispersed throughout the liquid water. Both within the liquid and at the surface, the ions are surrounded by layers of associated water molecules, or shells of water of solvation. The formation of these shells of water of solvation reduces the number of solvent molecules that have enough kinetic energy to escape as vapor. Thus, the solution has a lower vapor pressure than the pure solvent (water) would have at the same temperature.

Ionic solutes that dissociate, such as sodium chloride and calcium chloride, have greater effects on the vapor pressure than does a nondissociating solute such as glucose. Recall that each formula unit of the sodium chloride (NaCl) produces two particles in solution, a sodium ion and a chloride ion.

$$\mathrm{NaCl}(s) \xrightarrow{\mathrm{H_2O}} \mathrm{Na}^+(aq) + \mathrm{Cl}^-(aq)$$

Each formula unit of calcium chloride ($CaCl_2$) produces three particles, a calcium ion and two chloride ions.

$$\mathrm{CaCl_2}(s) \xrightarrow{\mathrm{H_2O}} \mathrm{Ca}^{2+}(aq) + 2\mathrm{Cl}^-(aq)$$

When glucose dissolves, the molecules do not dissociate.

$$\mathrm{C_6H_{12}O_6}(s) \xrightarrow{\mathrm{H_2O}} \mathrm{C_6H_{12}O_6}(aq)$$

Figure 16.12 compares the number of particles in three solutions of the same concentration. The decrease in a solution's vapor pressure is proportional to the number of particles the solute makes in solution. For example, the vapor-pressure lowering caused by 0.1 mol of sodium chloride in 1000 g of water is twice that caused by 0.1 mol of glucose in the same quantity of water. In the same way, 0.1 mol of $CaCl_2$ in 1000 g of water produces three times the vapor-pressure lowering as 0.1 mol of glucose in the same quantity of water.

READING SUPPORT

Build Study Skills: ***Concept Map***
As you read, construct a concept map that organizes the major ideas of this lesson. ***What factor determines the magnitude of colligative properties of solutions?***

Learn more about **freezing-point depression** *online.*

Freezing-Point Depression When a substance freezes, the particles of the solid take on an orderly pattern. The presence of a solute in water disrupts the formation of this pattern because of the shells of water of solvation. As a result, more kinetic energy must be withdrawn from a solution than from the pure solvent to cause the solution to solidify. The freezing point of a solution is lower than the freezing point of the pure solvent. The difference in temperature between the freezing point of a solution and the freezing point of the pure solvent is called the **freezing-point depression.**

Freezing-point depression is another colligative property. The magnitude of the freezing-point depression is proportional to the number of solute particles dissolved in the solvent and does not depend upon their identity. The addition of 1 mol of solute particles to 1000 g of water lowers the freezing point by 1.86°C. For example, if you add 1 mol (180 g) of glucose to 1000 g of water, the solution freezes at −1.86°C. However, if you add 1 mol (58.5 g) of sodium chloride to 1000 g of water, the solution freezes at −3.72°C, double the change for glucose. This difference occurs because 1 mol NaCl produces 2 mol of particles and, thus, doubles the freezing-point depression.

The freezing-point depression of aqueous solutions plays an important role in helping to keep travelers safe in cold, icy weather. The truck in Figure 16.13 spreads a layer of salt on the icy road to make the ice melt. The melted ice forms a solution with a lower freezing point than that of pure water. Similarly, ethylene glycol ($C_2H_6O_2$, antifreeze) is added to the water in automobile cooling systems to depress the freezing point of the water below 0°C. Automobiles can thus withstand subfreezing temperatures without freezing up.

Figure 16.13 De-icing Measures
Roads can be free of ice even at temperatures below 0°C if salt is applied. A common de-icer used on aircraft is a mixture of water and propylene glycol.
Infer *Why do you think $CaCl_2$ is a more effective road de-icer than NaCl?*

Figure 16.14 Antifreeze
The fluid circulating through a car's cooling system is a solution of water and ethylene glycol, or antifreeze. The resulting mixture freezes below 0°C and boils above 100°C.

Boiling-Point Elevation The boiling point of a substance is the temperature at which the vapor pressure of the liquid phase equals atmospheric pressure. As you just learned, adding a nonvolatile solute to a liquid solvent decreases the vapor pressure of the solvent. Because of the decrease in vapor pressure, additional kinetic energy must be added to raise the vapor pressure of the liquid phase of the solution to atmospheric pressure and initiate boiling. Thus, the boiling point of a solution is higher than the boiling point of the pure solvent. The difference in temperature between the boiling point of a solution and the boiling point of the pure solvent is the **boiling-point elevation.**

Figure 16.14 shows antifreeze being poured into a car's coolant tank. The antifreeze doesn't just lower the freezing point of the water in the cooling system. It also elevates the boiling point, which helps protect the engine from overheating in the summer.

Boiling-point elevation is a colligative property; it depends on the concentration of particles, not on their identity. Therefore, you can think about boiling-point elevation in terms of particles. It takes additional kinetic energy for the solvent particles to overcome the attractive forces that keep them in the liquid. Thus, the presence of a solute elevates the boiling point of the solvent. The magnitude of the boiling-point elevation is proportional to the number of solute particles dissolved in the solvent. The boiling point of water increases by 0.512°C for every mole of particles that the solute forms when dissolved in 1000 g of water.

CHEMISTRY & YOU

Q: *Solutes other than NaCl could be used to produce the same freezing-point depression in an ice-cream machine. What factors do you think make NaCl a good choice?*

ONLINE PROBLEMS

16.3 LessonCheck

28. Identify Name three colligative properties of solutions.

29. Explain Why does a solution have a lower vapor pressure than the pure solvent of that solution?

30. Explain Why does a solution have a depressed freezing point and an elevated boiling point compared with the pure solvent?

31. Compare Would a dilute or a concentrated sodium fluoride solution have a higher boiling point? Explain.

32. Compare An equal number of moles of KI and MgI_2 are dissolved in equal volumes of water. Which solution has the higher

a. boiling point?

b. vapor pressure?

c. freezing point?

33. Apply Concepts Review what you learned in Lesson 13.2 about the relationship between the vapor pressure of liquids and their boiling points. Explain why only nonvolatile solutes cause the elevation of the solvent's boiling point.

16.4 Calculations Involving Colligative Properties

CHEMISTRY & YOU

Q: *How hot is a pot of boiling pasta?* Cooking instructions for a wide variety of foods, from dried pasta to fresh vegetables, often call for the addition of a small amount of salt to the cooking water. Most people like the flavor of food cooked with salt.

But adding salt can have another effect on the cooking process. Recall that dissolved salt elevates the boiling point of water. Suppose you added a teaspoon of salt to two liters of water. A teaspoon of salt has a mass of about 20 g. Would the resulting boiling-point increase be enough to shorten the time required for cooking? In this lesson, you will learn how to calculate the amount the boiling point of the cooking water would rise.

Key Questions

What are two ways of expressing the ratio of solute to solvent in a solution?

How are freezing-point depression and boiling-point elevation related to molality?

Vocabulary

- molality (m)
- mole fraction
- molal freezing-point depression constant (K_f)
- molal boiling-point elevation constant (K_b)

Molality and Mole Fraction

What are two ways of expressing the ratio of solute to solvent in a solution?

Recall that colligative properties of solutions depend only on the number of solute particles dissolved in a given amount of solvent. **Chemists use two ways to express the ratio of solute particles to solvent particles: in molality and in mole fractions.**

Molality (m) is the number of moles of solute dissolved in 1 kilogram (1000 grams) of solvent. Molality is also known as molal concentration.

$$\text{Molality } (m) = \frac{\text{moles of solute}}{\text{kilogram of solvent}}$$

Note that molality is not the same as molarity. Molality refers to moles of solute per kilogram of solvent rather than moles of solute per liter of solution. In the case of water as the solvent, 1 kg or 1000 g equals a volume of 1000 mL, or 1 L.

You can prepare a solution that is 1.00 molal ($1m$) in glucose, for example, by adding 1.00 mol (180 g) of glucose to 1000 g of water. A 0.500 molal ($0.500m$) sodium chloride solution is prepared by dissolving 0.500 mol (29.3 g) of NaCl in 1.000 kg (1000 g) of water.

SampleProblem 16.7

Using Molality

How many grams of potassium iodide must be dissolved in 500.0 g of water to produce a 0.060 molal KI solution?

1 Analyze List the knowns and the unknown. According to the definition of molality, the final solution must contain 0.060 mol KI per 1000 g H_2O. Use the molality as a conversion factor to convert from mass of the solvent (H_2O) to moles of the solute (KI). Then use the molar mass of KI to convert from mol KI to g KI. The steps are as follows: mass of $H_2O \longrightarrow$ mol KI $\longrightarrow$ g KI.

KNOWNS

mass of water = 500.0 g = 0.5000 kg

solution concentration = 0.060*m*

molar mass of KI = 166.0 g/mol

UNKNOWN

mass of solute = ? g KI

2 Calculate Solve for the unknown.

Identify the conversion factor based on 0.060*m* that allows you to convert from g H_2O to mol KI.

$$\frac{0.060 \text{ mol KI}}{1.000 \text{ kg } H_2O}$$

Identify the conversion factor based on the molar mass of KI that allows you to convert from mol KI to g KI.

$$\frac{166.0 \text{ g KI}}{1 \text{ mol KI}}$$

Multiply the known solvent volume by the conversion factors.

$$0.5000 \text{ } \cancel{\text{kg } H_2O} \times \frac{0.060 \text{ } \cancel{\text{mol KI}}}{1.000 \text{ } \cancel{\text{kg } H_2O}} \times \frac{166.0 \text{ g KI}}{1 \text{ } \cancel{\text{mol KI}}} = \boxed{5.0 \text{ g KI}}$$

To make the 0.060-molal KI solution, you would dissolve 5.0 g of KI in 500.0 g of water.

3 Evaluate Does this result make sense? A 1-molal KI solution is one molar mass of KI (166.0 g) dissolved in 1000 g of water. The desired molal concentration (0.060*m*) is about $\frac{1}{20}$ of that value, so the mass of KI should be much less than the molar mass. The answer is correctly expressed to two significant figures.

34. How many grams of sodium fluoride are needed to prepare a 0.400*m* NaF solution that contains 750 g of water?

35. Calculate the molality of a solution prepared by dissolving 10.0 g NaCl in 600 g of water.

Remember: Molality equals moles of solute dissolved per kilogram of solvent.

The concentration of a solution also can be expressed as a mole fraction. The **mole fraction** of a solute in a solution is the ratio of the moles of that solute to the total number of moles of solvent and solute. In a solution containing n_A mol of solute A and n_B mol of solvent B, the mole fraction of solute A (X_A) and the mole fraction of solvent B (X_B) can be expressed as follows:

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

Note that mole fraction is a dimensionless quantity. The sum of the mole fractions of all the components in a solution equals unity, or one.

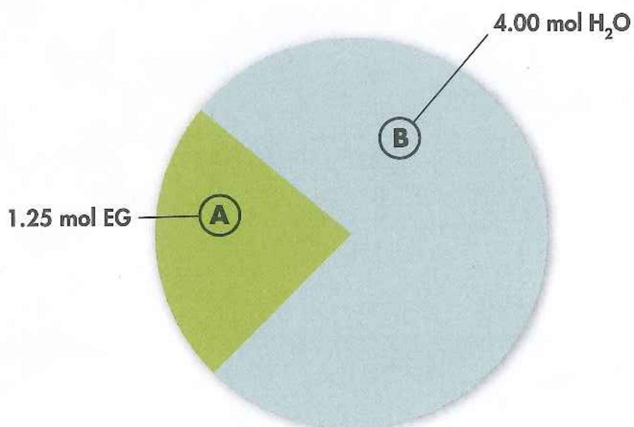
Figure 16.15 below and Sample Problem 16.8 on the next page illustrate how to calculate the mole fractions of the solute and solvent for a solution of ethylene glycol (EG) in water.

Figure 16.15 Mole Fraction
Ethylene glycol (EG) is added to water as antifreeze in the proportions shown. A mole fraction is the ratio of the number of moles of one substance to the total number of moles of all substances in the solution.
Infer ***What is the sum of all mole fractions in a solution?***

$$\text{Total moles} = A + B = 5.25 \text{ mol}$$

$$\text{Mole fraction EG} = \frac{A}{A+B} = \frac{1.25}{5.25}$$

$$\text{Mole fraction } H_2O = \frac{B}{A+B} = \frac{4.00}{5.25}$$



Freezing-Point Depression and Boiling-Point Elevation

How are freezing-point depression and boiling-point elevation related to molality?

Depressions of freezing points and elevations of boiling points are usually quite small. For example, if you add a teaspoon of salt to a pot of water and boil the resulting solution, you will have a hard time detecting any change in the boiling point using a cooking thermometer. It turns out that the elevation is just a small fraction of a degree Celsius. To measure colligative properties accurately, you would need a thermometer that can measure temperatures to the nearest 0.001°C.

Another way to determine the magnitudes of colligative properties is by calculating them. You can do this if you know the molality of the solution and some reference data about the solvent.

SampleProblem 16.8

Calculating Mole Fractions

Ethylene glycol (EG, or $C_2H_6O_2$) is added to automobile cooling systems to protect against cold weather. What is the mole fraction of each component in a solution containing 1.25 mol of ethylene glycol and 4.00 mol of water?

1 Analyze **List the knowns and the unknowns.** The given quantities of solute (EG) and solvent (water) are expressed in moles. Use the equations for mole fraction of a solute and mole fraction of a solvent to solve this problem. (The pie graph in Figure 16.15 gives you a visual representation of the mole fraction of each component.)

KNOWNS

moles of ethylene glycol (n_{EG}) = 1.25 mol EG

moles of water (n_{H_2O}) = 4.00 mol H_2O

UNKNOWNS

mole fraction EG (X_{EG}) = ?

mole fraction H_2O (X_{H_2O}) = ?

2 Calculate **Solve for the unknowns.**

Write the equation for the mole fraction of ethylene glycol (X_{EG}) in the solution.

$$X_{EG} = \frac{n_{EG}}{n_{EG} + n_{H_2O}}$$

Note that the denominator for each mole fraction is the same: the total number of moles of solvent and solute in the solution.

Write the equation for the mole fraction of water (X_{H_2O}) in the solution.

$$X_{H_2O} = \frac{n_{H_2O}}{n_{EG} + n_{H_2O}}$$

Substitute the known values into each equation.

$$X_{EG} = \frac{n_{EG}}{n_{EG} + n_{H_2O}} = \frac{1.25\ \cancel{mol}}{1.25\ \cancel{mol} + 4.00\ \cancel{mol}} = 0.238$$

$$X_{H_2O} = \frac{n_{H_2O}}{n_{EG} + n_{H_2O}} = \frac{4.00\ \cancel{mol}}{1.25\ \cancel{mol} + 4.00\ \cancel{mol}} = 0.762$$

3 Evaluate **Does the result make sense?** The sum of the mole fractions of all the components in the solution equals 1 ($X_{EG} + X_{H_2O} = 1.000$). Each answer is correctly expressed to three significant figures.

36. What is the mole fraction of each component in a solution made by mixing 300 g of ethanol (C_2H_6O) and 500 g of water?

37. A solution contains 50.0 g of carbon tetrachloride (CCl_4) and 50.0 g of chloroform ($CHCl_3$). Calculate the mole fraction of each component in the solution.

Remember: Mole fraction is a mole ratio, not a mass ratio. If the given quantities are masses, you must first convert each mass to moles using the molar mass of the substance.

Interpret Graphs

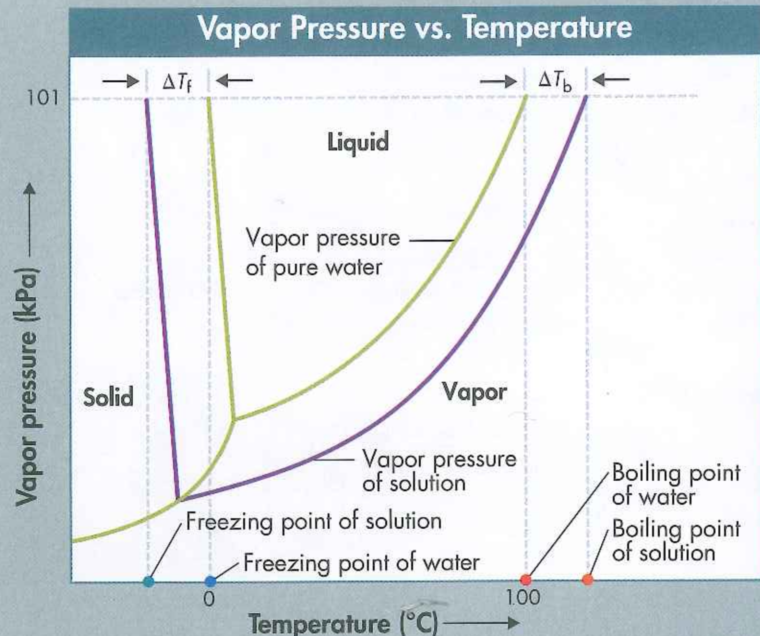


Figure 16.16 The graph shows the relationship between vapor pressure and temperature for pure water and aqueous solutions.

a. Read Graphs What is the freezing point of water? What is the boiling point?

b. Compare How do the freezing and boiling points of the solution compare to those of pure water?

c. Draw Conclusions Does adding a solute to water allow it to remain as a liquid over a longer or shorter temperature range? Explain.

The graph in Figure 16.16 shows that the freezing point of a solvent is lowered, and its boiling point is raised, by the addition of a nonvolatile solute. **The magnitudes of the freezing-point depression (ΔT_f) and the boiling-point elevation (ΔT_b) of a solution are directly proportional to the molal concentration (m), assuming the solute is molecular, not ionic.**

$$\Delta T_f \propto m$$
$$\Delta T_b \propto m$$

The change in the freezing temperature (T_f) is the difference between the freezing point of the solution and the freezing point of the pure solvent. Similarly, the change in the boiling temperature (T_b) is the difference between the boiling point of the solution and the boiling point of the pure solvent. The term m is the molal concentration of the solution.

With the addition of a constant, the proportionality between the freezing point depression (ΔT_f) and the molality m can be expressed as an equation.

$$\Delta T_f = K_f \times m$$

The constant, K_f, is the **molal freezing-point depression constant,** which is equal to the change in freezing point for a 1-molal solution of a nonvolatile molecular solute. The value of K_f depends upon the solvent. Its units are °C/m. Table 16.2 lists the K_f values for water and some other solvents.

Table 16.2
K_f and K_b Values for Some Common Solvents

Solvent	K_f (°C/m)	K_b (°C/m)
Acetic acid	3.90	3.07
Benzene	5.12	2.53
Camphor	37.7	5.95
Cyclohexane	20.2	2.79
Ethanol	1.99	1.19
Nitrobenzene	7.00	5.24
Phenol	7.40	3.56
Water	1.86	0.512

Sample Problem 16.9

Calculating the Freezing-Point Depression of a Solution

Antifreeze protects a car from freezing. It also protects it from overheating. Calculate the freezing-point depression of a solution containing exactly 100 g of ethylene glycol ($C_2H_6O_2$) antifreeze in 0.500 kg of water.

1 Analyze **List the knowns and the unknown.** Calculate the number of moles of $C_2H_6O_2$ and the molality of the solution. Then calculate the freezing-point depression using $\Delta T_f = K_f \times m$.

KNOWNS
mass of $C_2H_6O_2$ = 100 g
mass of water = 0.500 kg
K_f for H_2O = 1.86°C/m
molar mass of $C_2H_6O_2$ = 62.0 g/mol

UNKNOWN
ΔT_f = ?°C

2 Calculate **Solve for the unknown.**

Use the molar mass of $C_2H_6O_2$ to convert the mass of solute to moles.

$$100\,\cancel{g\,C_2H_6O_2} \times \frac{1\text{ mol } C_2H_6O_2}{62.0\,\cancel{g\,C_2H_6O_2}} = 1.61\text{ mol } C_2H_6O_2$$

Calculate the molality of the solution.

$$m = \frac{\text{mol solute}}{\text{kg solvent}} = \frac{1.61\text{ mol}}{0.500\text{ kg}} = 3.22m$$

Calculate the freezing-point depression.

$$\Delta T_f = K_f \times m = 1.86°\text{C}/\cancel{m} \times 3.22\cancel{m} = \boxed{5.99°\text{C}}$$

The freezing point of the solution is 0.00°C − 5.99°C = −5.99°C.

3 Evaluate **Does the result make sense?** A 1-molal solution reduces the freezing temperature by 1.86°C, so a decrease of 5.99°C for an approximately 3-molal solution is reasonable.

38. What is the freezing-point depression of an aqueous solution of 10.0 g of glucose ($C_6H_{12}O_6$) in 50.0 g H_2O?

39. Calculate the freezing-point depression of a benzene solution containing 400 g of benzene and 200 g of the molecular compound acetone (C_3H_6O). K_f for benzene is 5.12°C/m.

As you might expect, the boiling-point elevation of a solution can also be expressed as an equation. In this case, the proportionality constant is K_b, the **molal boiling-point elevation constant,** which is equal to the change in boiling point for a 1-molal solution of a nonvolatile molecular solute.

$$\Delta T_b = K_b \times m$$

Table 16.2 lists the K_b values for some solvents. Like K_f, K_b has units of °C/m.

Sample Problem 16.9 above described how to determine ΔT_f if the solute is a molecular compound. But for ionic compounds, both ΔT_f and ΔT_b depend upon the number of ions produced by each formula unit. This number is used to calculate an effective molality, as you'll see in Sample Problem 16.10.

CHEMISTRY & YOU

Q: *Does pasta cook at 100°C? After you read Sample Problem 16.10, calculate the boiling-point elevation for the solution described at the beginning of the lesson on page 538.*

Sample Problem 16.10

Calculating the Boiling Point of a Solution

What is the boiling point of a 1.50*m* NaCl solution?

1 Analyze **List the knowns and the unknown.** Each formula unit of NaCl dissociates into two particles, according to the equation $NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$. Based on the total number of dissociated particles, the effective molality is $2 \times 1.50m = 3.00m$. Calculate the boiling-point elevation (using the equation $\Delta T_b = K_b \times m$), and then add it to 100°C.

KNOWNS

solution concentration = 1.50*m* NaCl

K_b for H_2O = 0.512°C/*m*

UNKNOWN

boiling point = ?°C

2 Calculate **Solve for the unknown.**

Calculate the boiling-point elevation, making sure to use the molality of total dissociated particles in solution.

$$\Delta T_b = K_b \times m = 0.512°C/\cancel{m} \times 3.00\cancel{m} = 1.54°C$$

Calculate the boiling point of the solution.

$$T_b = 100°C + 1.54°C = \boxed{101.54°C}$$

3 Evaluate **Does the result make sense?** The boiling point increases about 0.5°C for each mole of solute particles, so the total change is reasonable. Because the boiling point of water is defined as exactly 100°C, this value does not limit the number of significant figures in the solution of the problem.

40. What is the boiling point of a solution that contains 1.25 mol $CaCl_2$ in 1400 g of water?

41. What mass of NaCl would have to be dissolved in 1.000 kg of water to raise the boiling point by 2.00°C?

ONLINE PROBLEMS 16.4 LessonCheck

42. **List** What are two ways of expressing the ratio of solute particles to solvent particles?

43. **Explain** How are freezing-point depression and boiling-point elevation related to molality?

44. **Calculate** How many grams of sodium bromide must be dissolved in 400.0 g of water to produce a 0.500 molal solution?

45. **Calculate** Calculate the mole fraction of each component in a solution of 2.50 mol ethanoic acid ($C_2H_4O_2$) in 10.00 mol of water.

46. **Predict** What is the freezing point of a solution of 12.0 g of CCl_4 dissolved in 750.0 g of benzene? The freezing point of benzene is 5.48°C; K_f is 5.12°C/*m*.

47. **Make Generalizations** Look at the table on page R1 of the Elements Handbook showing the distribution of elements in the oceans. What generalization can you make about the temperature at which ocean water will freeze? What effect does the presence of dissolved elements in the ocean have on the rate of evaporation of ocean water?

Small-Scale Lab

Making a Solution

Purpose

To make a solution and use carefully measured data to calculate the solution's concentration

Materials

- solid NaCl
- water
- 50-mL volumetric flask
- balance

Procedure

Measure the mass of a clean, dry, volumetric flask. Add enough solid NaCl to fill approximately one tenth of the volume of the flask. Measure the mass of the flask again. Half fill the flask with water and shake it gently until all the NaCl dissolves. Fill the flask with water to the 50-mL mark and measure the mass again.

Analyze and Conclude

Answer the following questions based on your data.

1. Percent by mass tells how many grams of solute are present in 100 g of solution.

$$\% \text{ by mass} = \frac{\text{mass of solute}}{\text{mass of solute} + \text{solvent}} \times 100\%$$

a. Calculate the mass of the solute (NaCl).

b. Calculate the mass of the solvent (water).

c. Calculate the percent by mass of NaCl in the solution.

2. Mole fraction tells how many moles of solute are present for every 1 mol of total solution.

$$\text{Mole fraction} = \frac{\text{mol NaCl}}{\text{mol NaCl} + \text{mol } H_2O}$$

a. Calculate the moles of NaCl solute.
Molar mass of NaCl = 58.5 g/mol

b. Calculate the moles of water.
Molar mass of H_2O = 18.0 g/mol

c. Calculate the mole fraction of your solution.

3. Molality (m) tells how many moles of solute are present in 1 kg of solvent.

$$m = \frac{\text{mol NaCl}}{\text{kg } H_2O}$$

Calculate the molality of your solution.

4. Molarity (M) tells how many moles of solute are dissolved in 1 L of solution.

$$M = \frac{\text{mol NaCl}}{\text{L solution}}$$

a. Calculate the liters of solution.

b. Calculate the molarity of the NaCl solution.

5. Density tells how many grams of solution are present in 1 mL of solution.

$$\text{Density} = \frac{\text{g solution}}{\text{mL solution}}$$

Calculate the density of the solution.

You're the Chemist

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. Analyze Data Measure the mass of an empty volumetric flask. Use a small-scale pipette to extract a sample of your NaCl solution and deliver it to the flask. Measure the mass of the flask again and fill it with water to the 50-mL line. Measure the mass of the flask again. Calculate the concentration of this dilute solution using the same units you used to calculate the concentration of the NaCl solution. Are the results you obtained reasonable?

2. Design an Experiment Design and carry out an experiment to make a solution of table sugar quantitatively. Calculate the concentration of the table sugar solution using the same units you used to calculate the concentration of the NaCl solution. Is the effective molality of the table sugar solution the same as the effective molality of a sodium chloride solution of the same concentration? Recall that effective molality is the concentration value used to calculate boiling-point elevation and freezing-point depression.

16 Study Guide

BIG IDEA

THE MOLE AND QUANTIFYING MATTER

Solubility, miscibility, concentration, and colligative properties are used to describe and characterize solutions. Solution concentration can be quantified in terms of molarity (moles of solute per liter of solution), molality (moles of solute per kilogram of solvent), percent by volume, and percent by mass.

16.1 Properties of Solutions

Factors that determine how fast a substance dissolves are stirring, temperature, and surface area.

In a saturated solution, a state of dynamic equilibrium exists between the solution and any undissolved solute, provided that the temperature remains constant.

Temperature affects the solubility of solid, liquid, and gaseous solutes in a solvent; both temperature and pressure affect the solubility of gaseous solutes.

- saturated solution (520)
- solubility (520)
- unsaturated solution (520)
- miscible (521)
- immiscible (521)
- supersaturated solution (522)
- Henry's law (523)

Key Equation

$$\text{Henry's law: } \frac{S_1}{P_1} = \frac{S_2}{P_2}$$

16.2 Concentrations of Solutions

To calculate the molarity of a solution, divide the moles of solute by the volume of the solution in liters.

Diluting a solution reduces the number of moles of solute per unit volume, but the total number of moles of solute in solution does not change.

Percent by volume is the ratio of the volume of solute to the volume of solution. Percent by mass is the ratio of the mass of the solute to the mass of the solution.

- concentration (525)
- dilute solution (525)
- concentrated solution (525)
- molarity (M) (525)

Key Equations

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

$$M_1 \times V_1 = M_2 \times V_2$$

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

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

16.3 Colligative Properties of Solutions

Colligative properties of solutions include vapor-pressure lowering, freezing-point depression, and boiling-point elevation.

- colligative property (534)
- freezing-point depression (536)
- boiling-point elevation (537)

16.4 Calculations Involving Colligative Properties

Chemists use two ways to express the ratio of solute to solvent: in molality and in mole fractions.

The magnitudes of freezing-point depression and boiling-point elevation are proportional to molality.

- molality (m) (538)
- mole fraction (540)
- molal freezing-point depression constant (K_f) (542)
- molal boiling-point elevation constant (K_b) (543)

Key Equations

$$\text{Molality } (m) = \frac{\text{moles of solute}}{\text{kilogram of solvent}}$$

$$\text{Mole fractions: } X_A = \frac{n_A}{n_A + n_B} \qquad X_B = \frac{n_B}{n_A + n_B}$$

$$\Delta T_f = K_f \times m$$

$$\Delta T_b = K_b \times m$$

Math Tune-Up: Solution Concentration Problems

Problem	❶ Analyze	❷ Calculate	❸ Evaluate
What volume of 12.00*M* sulfuric acid is required to prepare 1.00 L of 0.400*M* sulfuric acid?	Knowns: $M_1 = 12.00M\ H_2SO_4$ $M_2 = 0.400M\ H_2SO_4$ $V_2 = 1.00$ L of $0.400M\ H_2SO_4$ Unknown: $V_1 = ?$ L of $12.00M\ H_2SO_4$ Use the following equation to solve for the unknown initial volume of solution that is diluted: $M_1 \times V_1 = M_2 \times V_2$	Solve the equation for V_1 and substitute. $V_1 = \frac{M_2 \times V_2}{M_1}$ $V_1 = \frac{0.400\cancel{M} \times 1.00\text{ L}}{12.00\cancel{M}}$ $V_1 = 0.0333$ L	The concentration of the initial solution (12.00*M*) is 30 times larger than the concentration of the diluted solution (0.400*M*). So, the volume of the solution to be diluted should be one thirtieth the final volume of the diluted solution.
Ethanol is mixed with gasoline to make a solution called gasohol. What is the percent by volume of ethanol in gasohol when 95 mL of ethanol is added to sufficient gasoline to make 1.0 L of gasohol?	Knowns: volume of ethanol = 95 mL volume of solution = 1.0 L Unknown: solution concentration = ? % (v/v) Use the equation for percent by volume: $\%\ (v/v) = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$	Make sure the known volumes are expressed in the same units. Then calculate percent by volume of ethanol. $\%\ (v/v) = \frac{0.095\ \cancel{L}}{1.00\ \cancel{L}} \times 100\%$ $= 9.5\%\ (v/v)$	The volume of the solute is about one tenth the volume of the solution, so the answer is reasonable. The answer is correctly expressed to two significant figures.
Calculate the molality of a solution prepared by mixing 5.40 g LiBr with 444 g of water.	Knowns: mass of solute = 5.40 g mass of water = 444 g molar mass of LiBr = 86.8 g Unknown: solution concentration = ?*m* Use the equation for molal concentration: $\text{Molality} = \frac{\text{mol of solute}}{\text{kg of solvent}}$	Convert the mass of the solute to moles of solute. $5.40\ \cancel{\text{g LiBr}} \times \frac{1\text{ mol LiBr}}{86.8\ \cancel{\text{g LiBr}}} =$ 0.0622 mol LiBr Calculate molality. $\text{Molality} = \frac{0.0622\text{ mol LiBr}}{0.444\text{ kg } H_2O}$ $= 0.140m$	The answer has the correct units (mol of solute per kg of solvent) and is correctly expressed to three significant figures.

Remember: Molality is mol of solute per kg of solvent. Make sure you have the correct mass units in the denominator.

16 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

16.1 Properties of Solutions

48. Name and distinguish between the two components of a solution.

49. Explain why the dissolved component does not settle out of a solution.

50. Define the following terms: *solubility, saturated solution, unsaturated solution, miscible,* and *immiscible.*

51. If a saturated solution of sodium nitrate is cooled, what change might you observe?

52. Can a solution with undissolved solute be supersaturated? Explain.

53. What mass of $AgNO_3$ can be dissolved in 250 g of water at 20°C? Use Table 16.1.

54. What is the effect of pressure on the solubility of gases in liquids?

★**55.** The solubility of methane, the major component of natural gas, in water at 20°C and 1.00 atm pressure is 0.026 g/L. If the temperature remains constant, what will be the solubility of this gas at the following pressures?

a. 0.60 atm
b. 1.80 atm

16.2 Concentrations of Solutions

56. Knowing the molarity of a solution is more meaningful than knowing whether a solution is dilute or concentrated. Explain.

57. Define *molarity,* and then calculate the molarity of each solution.

a. 1.0 mol KCl in 750 mL of solution
b. 0.50 mol $MgCl_2$ in 1.5 L of solution

★**58.** How many milliliters of 0.500*M* KCl solution would you need to dilute to make 100.0 mL of 0.100*M* KCl?

★**59.** Calculate the molarity of a solution that contains 0.50 g of NaCl dissolved in 100 mL of solution.

60. Calculate the moles and grams of solute in each solution.

a. 1.0 L of 0.50*M* NaCl
b. 5.0×10^2 mL of 2.0*M* KNO_3
c. 250 mL of 0.10*M* $CaCl_2$
d. 2.0 L of 0.30*M* Na_2SO_4

★**61.** Calculate the grams of solute required to make the following solutions:

a. 2500 g of saline solution (0.90% NaCl (m/m))
b. 0.050 kg of 4.0% (m/m) $MgCl_2$

62. What is the percent by mass of sodium chloride in each of the following solutions?

a. 44 g NaCl dissolved in 756 g H_2O
b. 15 g NaCl dissolved in 485 g H_2O
c. 135 g NaCl dissolved in 765 g H_2O

★**63.** What is the concentration (in % (v/v)) of the following solutions?

a. 25 mL of ethanol (C_2H_6O) is diluted to a volume of 150 mL with water.
b. 175 mL of isopropyl alcohol (C_3H_8O) is diluted with water to a total volume of 275 mL.

16.3 Colligative Properties of Solutions

64. What are colligative properties? Identify three colligative properties and explain why each occurs.

65. Which has the higher boiling point:

a. seawater or distilled water?
b. 1.0*M* KNO_3 or 1.5*M* KNO_3?
c. 0.100*M* KCl or 0.100*M* $MgCl_2$?

66. Why does a 1*m* solution of calcium nitrate have a lower freezing point than a 1*m* solution of sodium nitrate?

67. Explain how a decrease in the vapor pressure of a solution results in an increase in its boiling point.

68. When the water inside a living cell freezes, the ice crystals damage the cell. The wood frog is a unique creature that can survive being frozen. In extremely cold conditions, the frog's liver produces large amounts of glucose ($C_6H_{12}O_6$), which becomes concentrated in the frog's cells. How does the glucose help prevent ice from forming in the frog's cells?

16.4 Calculations Involving Colligative Properties

69. Distinguish between a 1*M* solution and a 1*m* solution.

70. Describe how you would make an aqueous solution of methanol (CH_4O) in which the mole fraction of methanol is 0.40.

71. What is the boiling point of each solution?

a. 0.50 mol glucose in 1000 g H_2O
b. 1.50 mol NaCl in 1000 g H_2O

***72.** What is the freezing point of each solution?

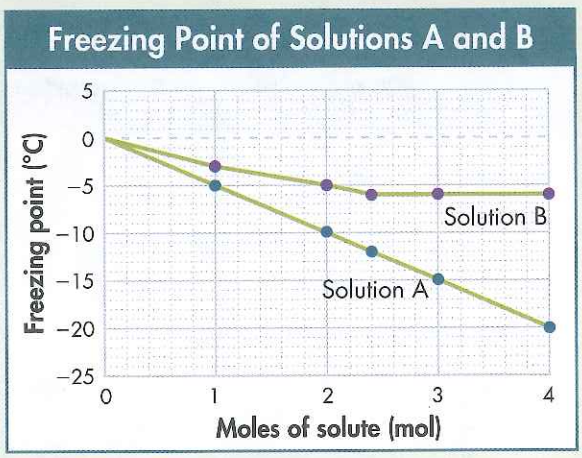
a. 1.40 mol Na_2SO_4 in 1750 g H_2O
b. 0.060 mol $MgSO_4$ in 100 g H_2O

73. Determine the freezing points of each 0.20*m* aqueous solution.

a. K_2SO_4
b. $CsNO_3$
c. $Al(NO_3)_3$

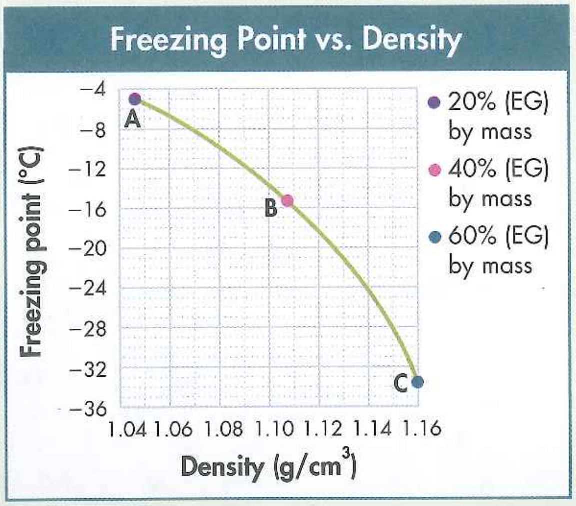
Understand Concepts

74. Different numbers of moles of two different solutes, A and B, were added to identical quantities of water. The graph shows the freezing point of each of the solutions formed.



a. Explain the relative slopes of the two lines between 0 and 2 mol of solute added.
b. Why does the freezing point for solution B not continue to drop as amounts of solute B are added beyond 2.4 mol?

75. A mixture of ethylene glycol (EG) and water is used as antifreeze in automobile engines. The freezing point and density of the mixture vary with the percent by mass of (EG) in the mixture. On the following graph, point A represents 20% (EG) by mass; point B, 40%; and point C, 60%.



a. What is the density of the antifreeze mixture that freezes at −25°C?
b. What is the freezing point of a mixture that has a density of 1.06?
c. Estimate the freezing point of a mixture that is 30% by mass (EG).

***76.** Calculate the freezing- and boiling-point changes for a solution containing 12.0 g of naphthalene ($C_{10}H_8$) in 50.0 g of benzene.

77. Describe how you would prepare an aqueous solution of acetone (C_3H_6O) in which the mole fraction of acetone is 0.25.

78. The solubility of sodium hydrogen carbonate ($NaHCO_3$) in water at 20°C is 9.6 g/100 g H_2O. What is the mole fraction of $NaHCO_3$ in a saturated solution? What is the molality of the solution?

79. A solution is labeled 0.150*m* NaCl. What are the mole fractions of the solute and solvent in this solution?

80. You are given a clear aqueous solution containing KNO_3. How would you determine experimentally if the solution is unsaturated, saturated, or supersaturated?

81. Plot a graph of solubility versus temperature for the three gases listed in Table 16.1.

82. Calculate the freezing point and the boiling point of a solution that contains 15.0 g of urea (CH_4N_2O) in 250 g of water. Urea is a covalently bonded compound.

83. Calculate the mole fractions in a solution that is 25.0 g of ethanol (C_2H_6O) and 40.0 g of water.

84. Estimate the freezing point of an aqueous solution of 20.0 g of glucose ($C_6H_{12}O_6$) dissolved in 500.0 g of water.

*85. The solubility of KCl in water at 20°C is 34.0 g KCl/100 g H_2O. A warm solution containing 50.0 g KCl in 130 g H_2O is cooled to 20°C.
   a. How many grams of KCl remain dissolved?
   b. How many grams came out of solution?

86. How many moles of ions are present when 0.10 mol of each compound is dissolved in water?
   a. K_2SO_4
   b. $Fe(NO_3)_3$
   c. $Al_2(SO_4)_3$
   d. $NiSO_4$

Think Critically

* 87. **Analyze Data** A solution contains 26.5 g NaCl in 75.0 g H_2O at 20°C. Determine if the solution is unsaturated, saturated, or supersaturated. (The solubility of NaCl at 20°C is 36.0 g/100 g H_2O.)

88. **Infer** An aqueous solution freezes at −2.47°C. What is its boiling point?

89. **Calculate** Percent (mass/volume), or % (m/v), is the number of grams of solute per 100 mL of solution. Hydrogen peroxide is often sold commercially as a 3.0% (m/v) aqueous solution.
   a. If you buy a 250-mL bottle of 3.0% H_2O_2 (m/v), how many grams of hydrogen peroxide have you purchased?
   b. What is the molarity of this solution?

90. **Calculate** How many grams of $NaNO_3$ will precipitate if a saturated solution of $NaNO_3$ in 200 g H_2O at 50°C is cooled to 20°C?

*91. **Calculate** What is the molar mass of a nondissociating compound if 5.76 g of the compound in 750 g of benzene gives a freezing-point depression of 0.460°C?

*92. **Calculate** The molality of an aqueous solution of sugar ($C_{12}H_{22}O_{11}$) is 1.62*m*. Calculate the mole fractions of sugar and water.

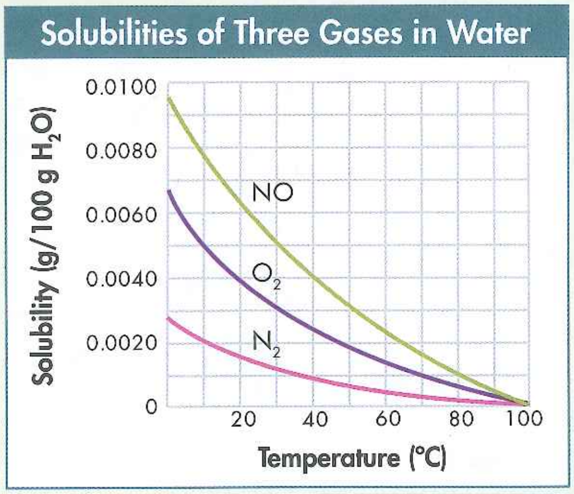
93. **Apply Concepts** Why might calcium chloride spread on icy roads be more effective at melting ice than an equal amount of sodium chloride?

94. **Calculate** The following table lists the molar concentrations of the most abundant monatomic ions in seawater. Calculate the mass in grams of each ion contained in 5.00 L of seawater. The density of seawater is 1.024 g/mL.

Ion	Molarity (*M*)
Chloride	0.546
Sodium	0.470
Magnesium	0.053
Calcium	0.0103
Potassium	0.0102

95. **Compare and Contrast** Which will have a greater boiling point elevation: 3.00 g $Ca(NO_3)_2$ in 60.0 g of water or 6.00 g $Ca(NO_3)_2$ in 30.0 g of water?

96. **Interpret Graphs** The graph shows the effect of temperature on the solubilities of oxygen gas (O_2), nitrogen gas (N_2), and nitrogen monoxide (NO) in water.



   a. How does an increase in temperature affect the solubility in water of each gas?
   b. At what temperature do the gases become virtually insoluble?
   c. Use kinetic theory to explain the solubility behavior shown in the graph.

Enrichment

97. **Calculate** When an excess of zinc is added to 800 mL of a hydrochloric acid solution, the solution evolves 1.21 L of hydrogen gas measured over water at 21°C and 747.5 mm Hg. What was the molarity of the acid? The vapor pressure of water at 21°C is 18.6 mm Hg.

98. **Calculate** How many milliliters of 1.50*M* HNO_3 contain enough nitric acid to dissolve an old copper penny with a mass of 3.94 g?

$$3Cu + 8HNO_3 \longrightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$$

99. **Graph** One way to express the solubility of a compound is in terms of moles of compound that will dissolve in 1 kg of water. Solubility depends on temperature. Plot a graph of the solubility of potassium nitrate (KNO_3) from the following data:

Temperature (°C)	Solubility (mol/kg)
0	1.61
20	2.80
40	5.78
60	11.20
80	16.76
100	24.50

Using your graph, estimate

a. the solubility of KNO_3 at 76°C and at 33°C.
b. the temperature at which its solubility is 17.6 mol/kg of water.
c. the temperature at which the solubility is 4.24 mol/kg of water.

*100. **Calculate** A 250-mL sample of Na_2SO_4 is reacted with an excess of $BaCl_2$. If 5.28 g $BaSO_4$ is precipitated, what is the molarity of the Na_2SO_4 solution?

101. **Design an Experiment** Suppose you have an unknown compound and want to identify it by means of its molar mass. Design an experiment that uses the concept of freezing-point depression to obtain the molar mass. What laboratory measurements would you need to make? What calculations would be needed?

Write About Science

102. **Describe** Find a recipe for rock candy online or in a cookbook. Write a short paragraph describing how the recipe applies key concepts that you have learned about solutions. Use the terms *solute, solvent, solubility, crystallization*, and *supersaturated solution* in your paragraph.

*103. **Sequence** Write a stepwise procedure for preparing 100 mL of 0.50*M* KCl, starting with a stock solution that is 2.0*M* KCl.

CHEMYSTERY

That Sinking Feeling

Although you can't see it happening, the groundwater beneath your feet is very slowly dissolving away rocks and minerals below ground. Eventually, enough of these mineral solutes dissolve to hollow out underground cavities or caverns. A sinkhole occurs when the roof of an underground cavern depresses or collapses.

104. **Explain** Why do you think areas underlain by salt beds are prone to sinkholes?

105. **Infer** As you read earlier in this chapter, agitation can speed up the rate at which a solid dissolves in liquid. What forces might contribute to agitating groundwater as it dissolves minerals underground?

106. **Connect to the BIGIDEA** Limestone, which contains mostly calcium carbonate ($CaCO_3$), is insoluble in water. Yet areas underlain by limestone are prone to sinkholes. Read the article on limestone caves on page R9. How does the "dissolving" of limestone differ from solvation?

Cumulative Review

107. Convert each of the following mass measurements to its equivalent in kilograms.

a. 347 g
b. 73 mg
c. 9.43 mg
d. 877 mg

★108. Rubidium has two naturally occurring isotopes. Rubidium-85 (72.165%) has a mass of 84.912 amu. Rubidium-87 (27.835%) has a mass of 86.909 amu. Calculate the average atomic mass of rubidium.

109. What is the most significant difference between the Thomson model of the atom and the Rutherford model?

110. Name and give the symbol for the element in the following positions in the periodic table:

a. Group 7B, Period 4
b. Group 3A, Period 5
c. Group 1A, Period 7
d. Group 6A, Period 6

111. How many atoms of each element are present in four formula units of calcium permanganate?

112. Draw electron dot structures for the following atoms:

a. I b. Te c. Sb d. Sr

113. Terephthalic acid is an organic compound used in the synthesis of polyesters. Terephthalic acid contains 57.8 percent C, 3.64 percent H, and 38.5 percent O. The molar mass is approximately 166 g/mol. What is the molecular formula of terephthalic acid?

★114. The photograph shows one mole each of iron, copper, mercury, and sulfur.

a. What is the mass of each element?
b. How many atoms are in each sample?
c. How many moles is 25.0 g of each element?

★115. What is the volume occupied by 1500 g of hydrogen gas (H_2) at STP?

116. Identify the type of chemical reaction.

a. $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$
b. $2H_2O(l) \longrightarrow O_2(g) + 2H_2(g)$
c. $2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$
d. $C_2H_6O(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$
e. $Cl_2(aq) + 2KBr(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$
f. $Pb(NO_3)_2(aq) + 2NaCl(aq) \longrightarrow PbCl_2(s) + 2NaNO_3(aq)$

117. Write the net ionic equation for the following reaction:

$$2HI(aq) + Na_2S(aq) \longrightarrow H_2S(g) + 2NaI(aq)$$

118. Indicate by simple equations how the following substances ionize or dissociate in water:

a. NH_4Cl
b. $Cu(NO_3)_2$
c. HNO_3
d. $HC_2H_3O_2$
e. Na_2SO_4
f. $HgCl_2$

119. The equation for the combustion of methanol (CH_4O) is the following:

$$2CH_4O(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l)$$

What volume of oxygen, measured at STP, is required to completely burn 35.0 g of methanol?

★120. A cylinder of nitrogen gas at 25°C and 101.3 kPa is heated to 45°C. What is the new pressure of the gas?

121. Why does an ideal gas not exist?

122. What relationship exists between surface tension and intermolecular attractions in a liquid?

123. The solubility of hydrogen chloride gas in the polar solvent water is much greater than its solubility in the nonpolar solvent benzene. Why?

124. When soap is shaken with water, which is formed: a solution, a suspension, or a colloid? Explain.

If You Have Trouble With . . .

Question	107	108	109	110	111	112	113	114	115	116	117	118	119	120	121	122	123	124
See Chapter	3	4	5	6	7	7	10	10	10	11	11	11	12	14	14	15	15	15

Standardized Test Prep

Select the choice that best answers each question or completes each statement.

1. An aqueous solution is 65% (v/v) rubbing alcohol. How many milliliters of water are in a 95-mL sample of this solution?
(A) 62 mL (C) 33 mL
(B) 1.5 mL (D) 30 mL

2. When 2.0 mol of methanol is dissolved in 45 g of water, the mole fraction of methanol is
(A) 0.44. (C) 2.25.
(B) 0.043. (D) 0.55.

The lettered choices below refer to Questions 3–6. A lettered choice may be used once, more than once, or not at all.

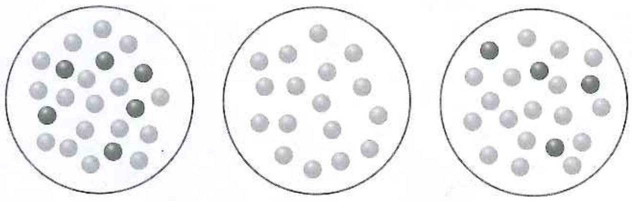
(A) moles/liter of solution
(B) grams/mole
(C) moles/kilogram of solvent
(D) °C/molal
(E) no units

Which of the above units is appropriate for each measurement?

3. molality
4. mole fraction
5. molar mass
6. molarity

Use the atomic windows to answer Questions 7–9. The windows show water and two aqueous solutions with different concentrations. Black spheres represent solute particles; gray spheres represent water.

(A) (B) (C)



10. Which of these actions will cause more sugar to dissolve in a saturated sugar water solution?
I. Add more sugar while stirring.
II. Add more sugar and heat the solution.
III. Grind the sugar to a powder; then add while stirring.

(A) I only (D) I and II only
(B) II only (E) II and III only
(C) III only

Tips for Success

Reading Data Tables Data tables are used to summarize data. When reading a table, try to figure out the relationships between the different columns and rows of information.

Use the description and the data table to answer Questions 11–14.

A student measured the freezing points of three different aqueous solutions at five different concentrations. The data are shown below.

Molarity (*M*)	Freezing Point Depression (°C)		
	NaCl	$CaCl_2$	C_2H_6O
0.5	1.7	2.6	0.95
1.0	3.5	5.6	2.0
1.5	5.3	8.3	3.0
2.0	7.2	11.2	4.1
2.5	9.4	14.0	5.3

11. Graph the data for all three solutes on the same graph, using molarity as the independent variable.

12. Summarize the relationship between molarity and freezing-point depression.

13. Compare the slopes of the three lines and explain any difference.

14. If you collected similar data for KOH and added a fourth line to your graph, which existing line would the new line approximate?

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12	13	14
See Lesson	16.2	16.4	16.4	16.4	16.2	16.2	16.3	16.3	16.3	16.1	16.3	16.4	16.3	16.3