

## 13.6: Colligative Properties: Vapor Pressure Lowering, Freezing Point Depression, Boiling Point Elevation, and Osmotic Pressure

### Key Concept Video Colligative Properties

Have you ever wondered why salt is added to the ice in an ice-cream maker? Or why salt is scattered on icy roads in cold climates? Salt lowers the temperature at which a saltwater solution freezes. A salt and water solution is liquid even below 0 °C. When salt is added to ice in the ice-cream maker, an ice/water/salt mixture forms that can reach a temperature of about -10 °C, at which point the cream freezes. On the winter road, the salt allows the ice to melt when the ambient temperature is below freezing.

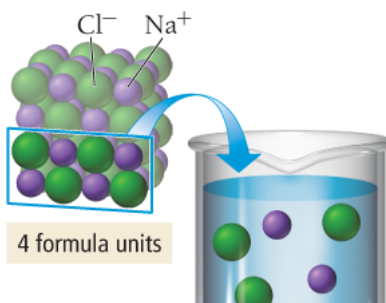


In winter, salt is applied to roads so that the ice will melt at lower temperatures.

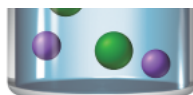
The depression of the freezing point of water by salt is an example of a **colligative property**, a property that depends on the number of particles dissolved in solution, not on the type of particle. In this section, we examine four colligative properties: vapor pressure lowering, freezing point depression, boiling point elevation, and osmotic pressure. Because these properties depend on the *number* of dissolved particles, we treat nonelectrolytes slightly differently than electrolytes when determining colligative properties. (See [Section 8.4](#) for a review of the difference between electrolytes and nonelectrolytes.) When 1 mol of a nonelectrolyte dissolves in water, it forms 1 mol of dissolved particles. When 1 mol of an electrolyte dissolves in water, however, it normally forms more than 1 mol of dissolved particles (as illustrated in [Figure 13.13](#)). For example, when 1 mol of NaCl dissolves in water, it forms 1 mol of dissolved  $\text{Na}^+$  ions and 1 mol of dissolved  $\text{Cl}^-$ . Therefore, the resulting solution has 2 mol of dissolved particles. The colligative properties of electrolyte solutions reflect this higher concentration of dissolved particles. In this section, we examine colligative properties of nonelectrolyte solutions; we then expand the concept to include electrolyte solutions in [Section 13.7](#).

**Figure 13.13** Electrolyte Dissociation

When sodium chloride dissolves in water, each mole of NaCl produces 2 mol of particles: 1 mol of  $\text{Na}^+$  cations and 1 mol of  $\text{Cl}^-$  anions.



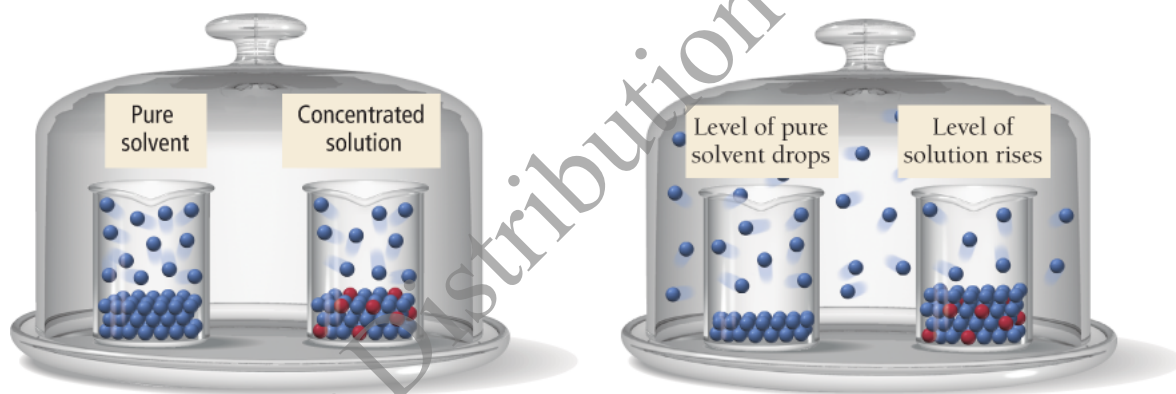
8 ions



## Vapor Pressure Lowering

Recall from [Section 11.5](#) that the vapor pressure of a liquid is the pressure of the gas above the liquid when the two are in dynamic equilibrium (that is, when the rate of vaporization equals the rate of condensation). What is the effect of a nonvolatile nonelectrolyte solute on the vapor pressure of the liquid into which it dissolves? The basic answer to this question is that *the vapor pressure of the solution is lower than the vapor pressure of the pure solvent*.

The vapor pressure of a solution is lower than that of the pure solvent because of nature's tendency toward mixing (toward greater entropy) that we discussed in [Section 13.2](#). We can see a dramatic demonstration of this tendency by placing a beaker of concentrated solution of a nonvolatile solute and a beaker of the pure solvent in a sealed container. Over time, the level of the pure solvent will drop, and the level of the solution will rise as molecules vaporize out of the pure solvent and condense into the solution. Why? The reason is nature's tendency to mix. If a pure solvent and concentrated solution are combined in a beaker, they naturally form a mixture in which the concentrated solution becomes less concentrated than it was initially. Similarly, if a pure solvent and concentrated solution are placed in a sealed container—even though they are in separate beakers—the two mix so that the concentrated solution becomes less concentrated.



The net transfer of solvent from the beaker containing pure solvent to the beaker containing the solution indicates that the vapor pressure of the solution is lower than that of the pure solvent. As solvent molecules vaporize, the vapor pressure in the sealed container rises. Before dynamic equilibrium can be attained, however, the pressure in the sealed container exceeds the vapor pressure of the solution, causing molecules to condense into the solution (the beaker on the right). Molecules constantly vaporize from the pure solvent (the beaker on the left), but the solvent's vapor pressure is never reached because molecules are constantly condensing into the solution. The result is a continuous transfer of solvent molecules from the pure solvent to the solution.

We can quantify the vapor pressure of a solution with **Raoult's law**:

$$P_{\text{solution}} = \chi_{\text{solvent}} P_{\text{solvent}}^{\circ}$$

In this equation,  $P_{\text{solution}}$  is the vapor pressure of the solution,  $\chi_{\text{solvent}}$  is the mole fraction of the solvent, and  $P_{\text{solvent}}^{\circ}$  is the vapor pressure of the pure solvent at the same temperature. For example, suppose a water sample at 25 °C contains 0.90 mol of water and 0.10 mol of a nonvolatile solute such as sucrose. The pure water has a vapor pressure of 23.8 torr. We calculate the vapor pressure of the solution by substituting into Raoult's law:

$$\begin{aligned} P_{\text{solution}} &= \chi_{\text{H}_2\text{O}} P_{\text{H}_2\text{O}}^{\circ} \\ &= 0.90(23.8 \text{ torr}) \\ &= 21.4 \text{ torr} \end{aligned}$$

Notice that the vapor pressure of the solution is directly proportional to the amount of the solvent in the solution. Because the solvent particles compose 90% of all of the particles in the solution, the vapor pressure of the solution is 90% of the vapor pressure of the pure solvent.

### Example 13.6 Calculating the Vapor Pressure of a Solution Containing a Nonvolatile Nonelectrolyte Solute

Calculate the vapor pressure at 25 °C of a solution containing 99.5 g sucrose ( $C_{12}H_{22}O_{11}$ ) and 300.0 mL water. The vapor pressure of pure water at 25 °C is 23.8 torr. Assume the density of water to be 1.00 g/mL.

**SORT** You are given the mass of sucrose and volume of water in a solution. You are also given the vapor pressure and density of pure water and asked to find the vapor pressure of the solution.

**GIVEN:**

99.5 g  $C_{12}H_{22}O_{11}$

300.0 mL  $H_2O$

$P_{H_2O}^\circ = 23.8$  torr at 25°C

$d_{H_2O} = 1.00$  g/mL

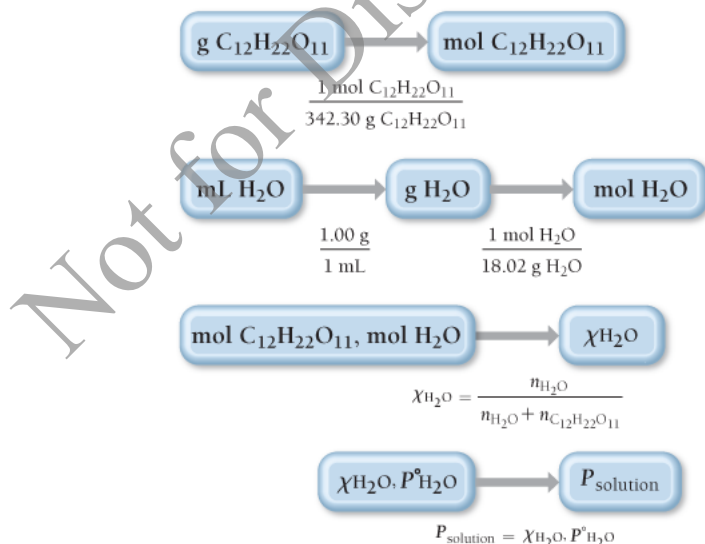
**FIND:**  $P_{\text{solution}}$

**STRATEGIZE** Raoult's law relates the vapor pressure of a solution to the mole fraction of the solvent and the vapor pressure of the pure solvent. Begin by calculating the amount in moles of sucrose and water.

Calculate the mole fraction of the solvent from the calculated amounts of solute and solvent.

Then use Raoult's law to calculate the vapor pressure of the solution.

**CONCEPTUAL PLAN**



**SOLVE** Calculate the number of moles of each solution component.

Use the number of moles of each component to calculate the mole fraction of the solvent ( $H_2O$ ).

Use the mole fraction of water and the vapor pressure of pure water to calculate the vapor pressure of the solution.

**SOLUTION**

**SOLUTION**

$$99.5 \text{ g C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}}{342.30 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}} = 0.2907 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}$$

$$300.0 \text{ mL H}_2\text{O} \times \frac{1.00 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 16.65 \text{ mol H}_2\text{O}$$

$$\begin{aligned} \chi_{\text{H}_2\text{O}} &= \frac{n_{\text{H}_2\text{O}}}{n_{\text{C}_{12}\text{H}_{22}\text{O}_{11}} + n_{\text{H}_2\text{O}}} \\ &= \frac{16.65 \text{ mol}}{0.2907 \text{ mol} + 16.65 \text{ mol}} \\ &= 0.9828 \end{aligned}$$

$$\begin{aligned} P_{\text{solution}} &= \chi_{\text{H}_2\text{O}} P_{\text{H}_2\text{O}}^\circ \\ &= 0.9828 (23.8 \text{ torr}) \\ &= 23.4 \text{ torr} \end{aligned}$$

**CHECK** The units of the answer are correct. The magnitude of the answer seems right because the calculated vapor pressure of the solution is just below that of the pure liquid, as you would expect for a solution with a large mole fraction of solvent.

**FOR PRACTICE 13.6** Calculate the vapor pressure at 25 °C of a solution containing 55.3 g ethylene glycol ( $\text{HOCH}_2\text{CH}_2\text{OH}$ ) and 285.2 g water. The vapor pressure of pure water at 25 °C is 23.8 torr.

**FOR MORE PRACTICE 13.6** A solution containing ethylene glycol and water has a vapor pressure of 7.88 torr at 10 °C. Pure water has a vapor pressure of 9.21 torr at 10 °C. What is the mole fraction of ethylene glycol in the solution?

**Interactive Worked Example 13.6 Calculating the Vapor Pressure of a Solution Containing a Nonvolatile Nonelectrolyte Solute**

## Vapor Pressures of Solutions Containing a Volatile (Nonelectrolyte) Solute

Some solutions contain not only a volatile solvent, but also a volatile *solute*. In this case, *both* the solvent and the solute contribute to the overall vapor pressure of the solution. A solution like this may be an **ideal solution** (in which case its behavior follows Raoult's law at all concentrations for both the solvent and the solute), or it may be nonideal (in which case it does not follow Raoult's law). An ideal solution is similar in concept to an ideal gas. Just as an ideal gas follows the ideal gas law exactly, an ideal solution follows Raoult's law exactly.

In an ideal solution, the solute–solvent interactions are similar in magnitude to the solute–solute and solvent–solvent interactions. In this type of solution, the solute simply dilutes the solvent and ideal behavior is observed. The vapor pressure of each solution component is described by Raoult's law throughout the entire composition range of the solution. For a two-component solution containing liquids A and B, we can write:

$$\begin{aligned} P_A &= \chi_A P_A^\circ \\ P_B &= \chi_B P_B^\circ \end{aligned}$$

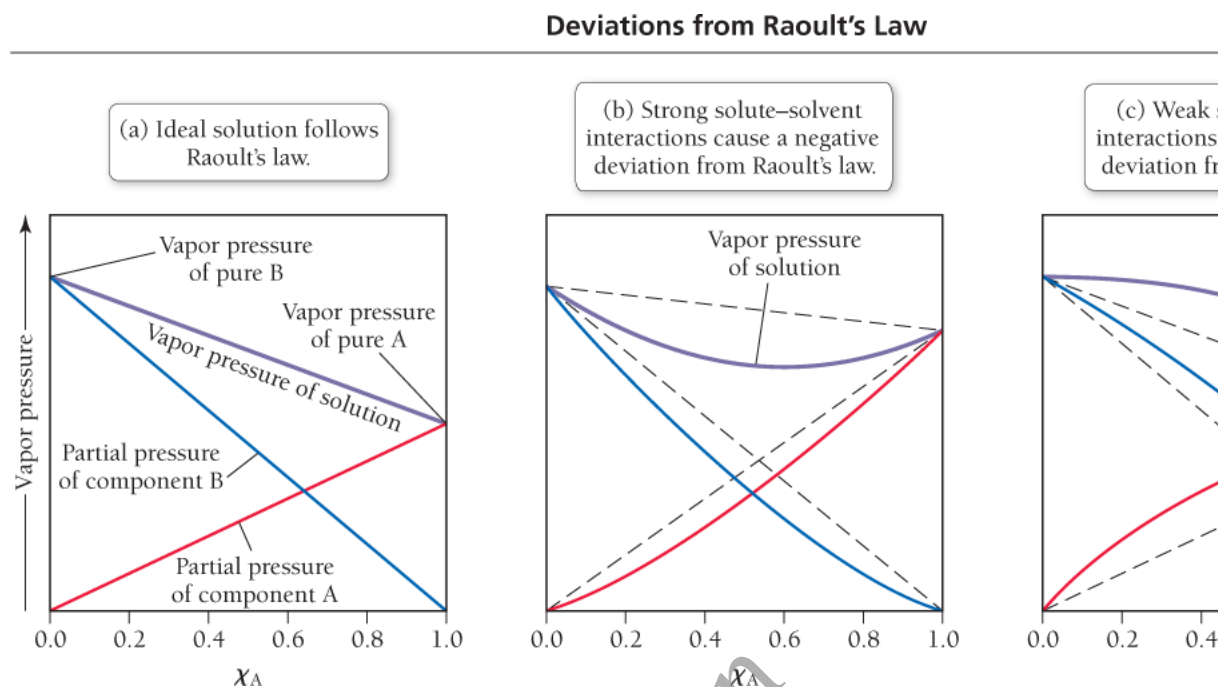
Over a complete range of composition of a solution, it no longer makes sense to designate a solvent and solute, so we simply label the two components A and B.

The total pressure above such a solution is the sum of the partial pressures of the components:

$$P_{\text{tot}} = P_A + P_B$$

Figure 13.14(a) is a plot of vapor pressure versus solution composition for an ideal two-component solution.

Figure 13.14 Behavior of Ideal and Nonideal Solutions



In a nonideal solution, the solute-solvent interactions are either stronger or weaker than the solvent-solvent interactions. If the solute-solvent interactions are stronger, then the solute tends to prevent the solvent from vaporizing as readily as it would otherwise. If the solution is sufficiently dilute, then the effect is small and Raoult's law works as an approximation. However, if the solution is not dilute, the effect is significant and the vapor pressure of the solution is *less than* that predicted by Raoult's law, as shown in Figure 13.14(b).

If, on the other hand, the solute-solvent interactions are weaker than the solvent-solvent interactions, then the solute tends to allow more vaporization than would occur with just the solvent. If the solution is not dilute, the effect is significant and the vapor pressure of the solution is *greater than* predicted by Raoult's law, as shown in Figure 13.14(c).

### Example 13.7 Calculating the Vapor Pressure of a Two-Component Solution

A solution contains 3.95 g of carbon disulfide ( $\text{CS}_2$ ) and 2.43 g of acetone ( $\text{CH}_3\text{COCH}_3$ ). The vapor pressures at  $35^\circ\text{C}$  of pure carbon disulfide and pure acetone are 515 torr and 332 torr, respectively. Assuming ideal behavior, calculate the vapor pressures of each component and the total vapor pressure above the solution. The experimentally measured total vapor pressure of the solution at  $35^\circ\text{C}$  is 645 torr. Is the solution ideal? If not, what can you say about the relative strength of carbon disulfide-acetone interactions compared to the acetone-acetone and carbon disulfide-carbon disulfide interactions?

**SORT** You are given the masses and vapor pressures of carbon disulfide and acetone and are asked to find the vapor pressures of each component in the mixture and the total pressure assuming ideal behavior.

**GIVEN:**

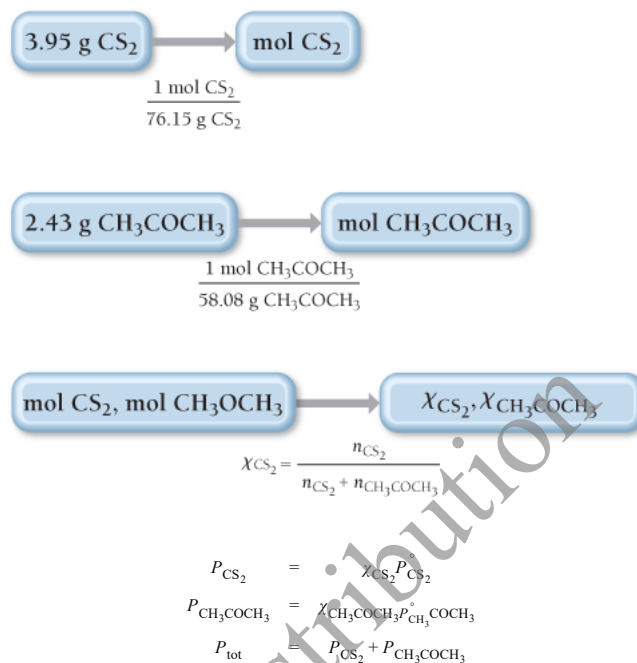
3.95 g  $\text{CS}_2$   
 2.34 g  $\text{CH}_3\text{COCH}_3$   
 $P_{\text{CS}_2}^\circ = 515 \text{ torr (at } 35^\circ\text{C)}$   
 $P_{\text{CH}_3\text{COCH}_3}^\circ = 332 \text{ torr (at } 35^\circ\text{C)}$   
 $P_{\text{tot}}(\text{exp}) = 645 \text{ torr (at } 35^\circ\text{C)}$

**FIND:**  $P_{\text{CS}_2}$ ,  $P_{\text{CH}_3\text{COCH}_3}$ ,  $P_{\text{tot}}(\text{ideal})$

**STRATEGIZE** This problem requires the use of Raoult's law to calculate the partial pressures of each component. In order to use Raoult's law, you must first calculate the mole fractions of the two components. Convert the masses of each component to moles and use the definition of mole fraction to calculate the mole fraction of carbon disulfide. You can then find the mole fraction of acetone because the mole fractions of the two components add up to 1.

Use the mole fraction of each component along with Raoult's law to calculate the partial pressure of each component. The total pressure is the sum of the partial pressures.

**CONCEPTUAL PLAN**



**RELATIONSHIPS USED**

$$\begin{aligned}
 \chi_A &= \frac{n_A}{n_A + n_B} \text{ (mole fraction definition)} \\
 P_A &= \chi_A P_A^\circ \text{ (Raoult's law)}
 \end{aligned}$$

**SOLVE** Begin by converting the mass of each component to the amounts in moles.

**SOLUTION**

$$\begin{aligned}
 3.95 \text{ g CS}_2 \times \frac{1 \text{ mol CS}_2}{76.15 \text{ g CS}_2} &= 0.05187 \text{ mol CS}_2 \\
 2.43 \text{ g CH}_3\text{COCH}_3 \times \frac{1 \text{ mol CH}_3\text{COCH}_3}{58.08 \text{ g CH}_3\text{COCH}_3} &= 0.04184 \text{ mol CH}_3\text{COCH}_3
 \end{aligned}$$

Calculate the mole fraction of carbon disulfide.

$$\begin{aligned}
 \chi_{\text{CS}_2} &= \frac{n_{\text{CS}_2}}{n_{\text{CS}_2} + n_{\text{CH}_3\text{COCH}_3}} \\
 &= \frac{0.05187 \text{ mol}}{0.05187 \text{ mol} + 0.04184 \text{ mol}} \\
 &= 0.5535
 \end{aligned}$$

Calculate the mole fraction of acetone by subtracting the mole fraction of carbon disulfide from 1.

$$\begin{aligned}
 \chi_{\text{CH}_3\text{COCH}_3} &= 1 - 0.5535 \\
 &= 0.4465
 \end{aligned}$$

Calculate the partial pressures of carbon disulfide and acetone by using Raoult's law and the given values of the vapor pressures of the pure substances.

$$\begin{aligned}
 P_{\text{CS}_2} &= \chi_{\text{CS}_2} P_{\text{CS}_2}^\circ \\
 &= 0.553 \cdot 5 \text{ (515 torr)} \\
 &= 285 \text{ torr} \\
 P_{\text{CH}_3\text{COCH}_3} &= \chi_{\text{CH}_3\text{COCH}_3} P_{\text{CH}_3\text{COCH}_3}^\circ \\
 &= 0.446 \cdot 5 \text{ (332 torr)} \\
 &= 148 \text{ torr}
 \end{aligned}$$

Calculate the total pressure by summing the partial pressures.

$$\begin{aligned}
 P_{\text{tot}}(\text{ideal}) &= 285 \text{ torr} + 148 \text{ torr} \\
 &= 433 \text{ torr}
 \end{aligned}$$

Lastly, compare the calculated total pressure for the ideal case to the experimentally measured total pressure. Since the experimentally measured pressure is greater than the calculated pressure, you can conclude that the interactions between the two components are weaker than the interactions between the components themselves.

$$\begin{aligned}
 P_{\text{tot}}(\text{exp}) &= 645 \text{ torr} \\
 P_{\text{tot}}(\text{exp}) &> P_{\text{tot}}(\text{ideal})
 \end{aligned}$$

The solution is not ideal and shows positive deviations from Raoult's law. Therefore, carbon disulfide–acetone interactions must be weaker than acetone–acetone and carbon disulfide–carbon disulfide interactions.

**CHECK** The units of the answer (torr) are correct. The magnitude seems reasonable given the partial pressures of the pure substances.

**FOR PRACTICE 13.7** A solution of benzene ( $\text{C}_6\text{H}_6$ ) and toluene ( $\text{C}_7\text{H}_8$ ) is 25.0% benzene by mass. The vapor pressures of pure benzene and pure toluene at 25 °C are 94.2 torr and 28.4 torr, respectively. Assuming ideal behavior, calculate the following:

- the vapor pressure of each solution component in the mixture
- the total pressure above the solution
- the composition of the vapor in mass percent

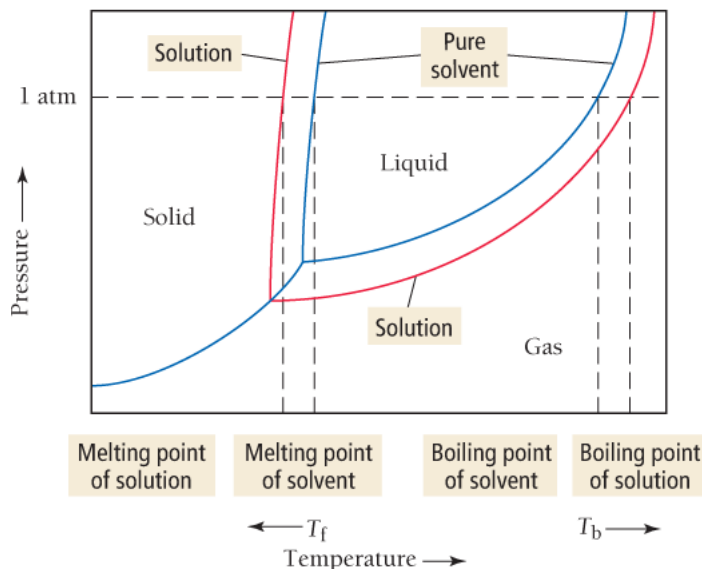
Why is the composition of the vapor different from the composition of the solution?

### Conceptual Connection 13.6 Raoult's Law

## Freezing Point Depression and Boiling Point Elevation

Vapor pressure lowering occurs at all temperatures. We can see the effect of vapor pressure lowering over a range of temperatures by comparing the phase diagrams for a pure solvent and for a solution containing a nonvolatile solute:

A nonvolatile solute lowers the vapor pressure of a solution, resulting in a lower freezing point and a higher boiling point.



Notice that the vapor pressure for the solution is shifted downward compared to that of the pure solvent. Consequently, the vapor pressure curve intersects the solid–gas curve at a lower temperature. The net effect is that the solution has a *lower melting point* and a *higher boiling point* than the pure solvent. These effects are called **freezing point depression** and **boiling point elevation**, both of which are colligative properties (like vapor pressure lowering).

The freezing point of a solution containing a nonvolatile solute is lower than the freezing point of the pure solvent. For example, antifreeze, used to prevent the freezing of engine blocks in cold climates, is an aqueous solution of ethylene glycol ( $\text{C}_2\text{H}_6\text{O}_2$ ). The more concentrated the solution, the lower the freezing point becomes.



Antifreeze is an aqueous solution of ethylene glycol or propylene glycol. The solution has a lower freezing point and a higher boiling point than pure water.

The amount that the freezing point decreases is given by the following equation:

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

where

- $\Delta T_f$  is the change in temperature of the freezing point in Celsius degrees (relative to the freezing point of the pure solvent), usually reported as a positive number;
- $m$  is the molality of the solution in moles solute per kilogram solvent; and



- $K_f$  is the freezing point depression constant for the solvent.

For water,

$$K_f = 1.86^\circ\text{C}/m$$

When an aqueous solution containing a dissolved solid solute freezes slowly, the ice that forms does not normally contain much of the solute. For example, when ice forms in ocean water, the ice is not saltwater but freshwater. As the ice forms, the crystal structure of the ice tends to exclude the solute particles. You can verify this yourself by partially freezing a saltwater solution in the freezer. Take out the newly formed ice, rinse it several times, and taste it. Compare its taste to the taste of the original solution. The ice is much less salty.

Table 13.6 lists freezing point depression and boiling point elevation constants for several liquids. Calculating the freezing point of a solution involves substituting into the freezing point depression equation, as Example 13.8 demonstrates.

Table 13.6 Freezing Point Depression and Boiling Point Elevation Constants for Several Liquid Solvents

Solvent	Normal Freezing Point ( $^\circ\text{C}$ )	$K_f$ ( $^\circ\text{C}/m$ )	Normal Boiling Point ( $^\circ\text{C}$ )	$K_b$ ( $^\circ\text{C}/m$ )
Benzene ( $\text{C}_6\text{H}_6$ )	5.5	5.12	80.1	2.53
Carbon tetrachloride ( $\text{CCl}_4$ )	-22.9	29.9	76.7	5.03
Chloroform ( $\text{CHCl}_3$ )	-63.5	4.70	61.2	3.63
Ethanol ( $\text{C}_2\text{H}_5\text{OH}$ )	-114.1	1.99	78.3	1.22
Diethyl ether ( $\text{C}_4\text{H}_{10}\text{O}$ )	-116.3	1.79	34.6	2.02
Water ( $\text{H}_2\text{O}$ )	0.00	1.86	100.0	0.512

### Example 13.8 Freezing Point Depression

What is the freezing point of a 1.7- $m$  aqueous ethylene glycol solution?

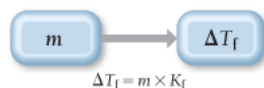
**SORT** You are given the molality of a solution and asked to find its freezing point.

**GIVEN:** 1.7- $m$  solution

**FIND:** freezing point (from  $\Delta T_f$ )

**STRATEGIZE** To solve this problem, use the freezing point depression equation.

**CONCEPTUAL PLAN**



**SOLVE** Substitute into the equation to calculate  $\Delta T_f$ . The actual freezing point is the freezing point of pure water ( $0.00^\circ\text{C}$ ) -  $\Delta T_f$ .

**SOLUTION**

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

$$= 1.7\,m \times 1.86^{\circ}\text{C}/m$$

$$= 3.2^{\circ}\text{C}$$

$$\text{freezing point} = 0.00^{\circ}\text{C} - 3.2^{\circ}\text{C}$$

$$= -3.2^{\circ}\text{C}$$

**CHECK** The units of the answer are correct. The magnitude seems about right. The expected range for freezing points of an aqueous solution is anywhere from  $-10^{\circ}\text{C}$  to just below  $0^{\circ}\text{C}$ . Any answers out of this range would be suspect.

### FOR PRACTICE 13.8

Calculate the freezing point of a  $2.6\text{-}m$  aqueous sucrose solution.

The boiling point of a solution containing a nonvolatile solute is higher than the boiling point of the pure solvent. In automobiles, antifreeze not only prevents the freezing of water within engine blocks in cold climates, it also prevents the boiling of water within engine blocks in hot climates. The amount that the boiling point rises in solutions is given by the equation:

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

where

- $\Delta T_b$  is the change in temperature of the boiling point in Celsius degrees (relative to the boiling point of the pure solvent);
- $m$  is the molality of the solution in moles solute per kilogram solvent; and
- $K_b$  is the boiling point elevation constant for the solvent.

For water,

$$K_b = 0.512^{\circ}\text{C}/m$$

Calculating the boiling point of a solution involves substituting into the boiling point elevation equation, as

**Example 13.9** demonstrates.

### Conceptual Connection 13.7 Boiling Point Elevation

### Example 13.9 Boiling Point Elevation

What mass of ethylene glycol ( $\text{C}_2\text{H}_6\text{O}_2$ ), in grams, must you add to  $1.0\text{ kg}$  of water to produce a solution that boils at  $105.0^{\circ}\text{C}$ ? (Since pure water boils at  $100.0^{\circ}\text{C}$ , the change in boiling point is  $5.0^{\circ}\text{C}$ .)

**SORT** You are given the desired boiling point of an ethylene glycol solution containing  $1.0\text{ kg}$  of water and asked to determine the mass of ethylene glycol needed to achieve the boiling point.

**GIVEN:**  $\Delta T_b = 5.0^{\circ}\text{C}$ ,  $1.0\text{ kg H}_2\text{O}$

**FIND:**  $\text{g C}_2\text{H}_6\text{O}_2$

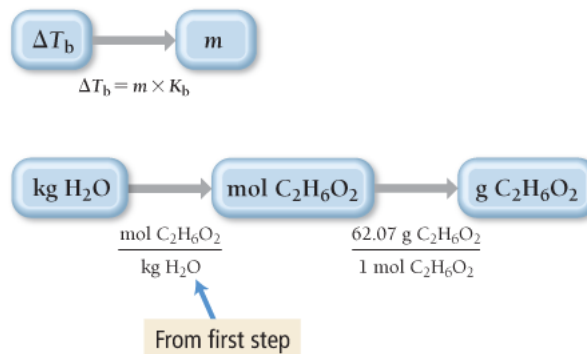
**STRATEGIZE** To solve this problem, use the boiling point elevation equation to calculate the desired molality of the solution from  $\Delta T_b$ .

Then use that molality to determine how many moles of ethylene glycol are needed per kilogram of water.

Finally, calculate the molar mass of ethylene glycol and use it to convert from moles of ethylene glycol to

Finally, calculate the molar mass of ethylene glycol and use it to convert from moles of ethylene glycol to mass of ethylene glycol.

### CONCEPTUAL PLAN



### RELATIONSHIPS USED

$$\text{C}_2\text{H}_6\text{O}_2 \text{ molar mass} = 62.07 \text{ g/mol}$$

$$\Delta T_b = m \times K_b \text{ (boiling point elevation)}$$

**SOLVE** Begin by solving the boiling point elevation equation for molality and substituting the required quantities to calculate  $m$ .

### SOLUTION

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

$$m = \frac{\Delta T_b}{K_b} = \frac{5.0^\circ\text{C}}{0.512^\circ\text{C}/m} = 9.77m$$

$$1.0 \text{ kg H}_2\text{O} \times \frac{9.77 \text{ mol C}_2\text{H}_6\text{O}_2}{\text{kg H}_2\text{O}} \times \frac{62.07 \text{ g C}_2\text{H}_6\text{O}_2}{1 \text{ mol C}_2\text{H}_6\text{O}_2} = 6.1 \times 10^2 \text{ g C}_2\text{H}_6\text{O}_2$$

**CHECK** The units of the answer are correct. The magnitude might seem a little high initially, but the boiling point elevation constant is so small that a lot of solute is required to raise the boiling point by a small amount.

**FOR PRACTICE 13.9** Calculate the boiling point of a 3.60- $m$  aqueous sucrose solution.

### Interactive Worked Example 13.9 Boiling Point Elevation

## Osmotic Pressure

**Osmosis**  $\Phi$  is the flow of solvent from a solution of lower solute concentration to one of higher solute concentration. Concentrated solutions draw solvent from more dilute solutions because of nature's tendency to mix. **Figure 13.15**  $\square$  illustrates an osmosis cell. The left side of the cell contains a concentrated saltwater solution, and the right side of the cell contains pure water. A **semipermeable membrane**  $\Phi$ —a membrane that selectively allows some substances to pass through but not others—separates the two halves of the cell.

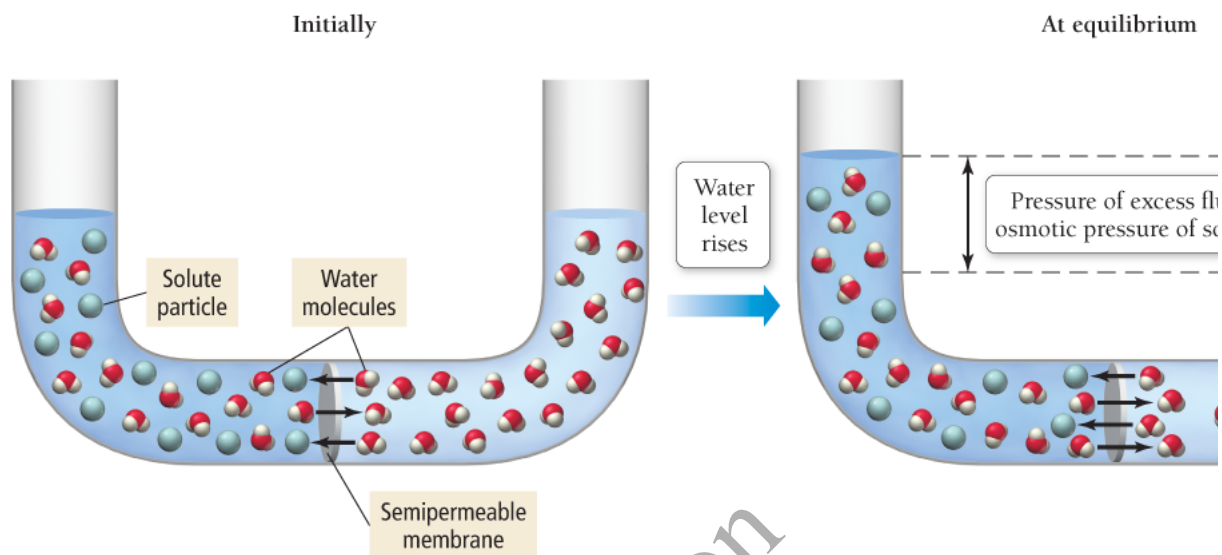
Water flows by osmosis from the pure-water side of the cell through the semipermeable membrane and into the saltwater side. Over time, the water level on the left side of the cell rises, while the water level on the right side of the cell falls. Notice the similarity between osmosis and vapor pressure lowering. In both cases, as a solution becomes concentrated, it develops a tendency to draw pure solvent to itself. In the case of vapor pressure lowering, the pure solvent is drawn from the gas state. In the case of osmosis, the pure solvent is drawn from the liquid state. In both cases, the solution becomes more dilute as it draws the pure solvent to itself—nature's

tendency to mix is powerful.

**Figure 13.15 An Osmosis Cell**

In an osmosis cell, water flows from the pure-water side of the cell through the semipermeable membrane to the saltwater side.

## Osmosis and Osmotic Pressure



If external pressure is applied to the solution in the left side of the cell, this process can be opposed and even stopped. The pressure required to stop the osmotic flow, called the **osmotic pressure**  $\pi$ , is given by the following equation:

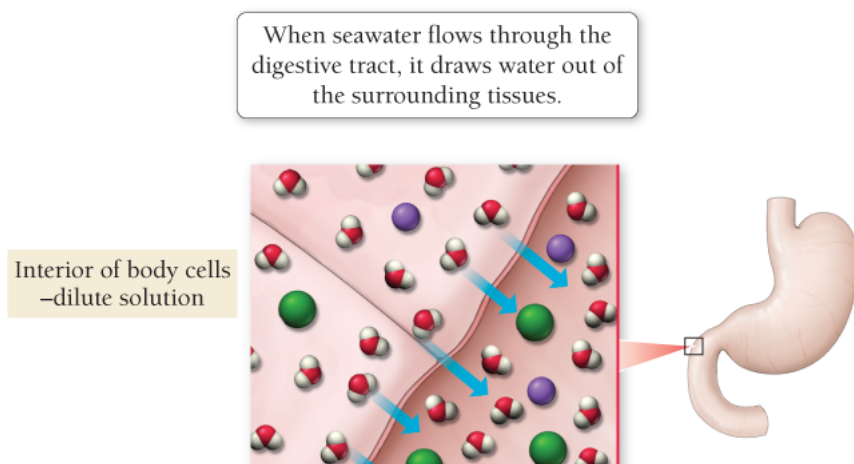
$$\pi = MRT$$

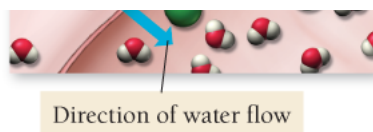
where  $M$  is the molarity of the solution,  $T$  is the temperature (in kelvins), and  $R$  is the ideal gas constant ( $0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K}$ ).

Osmosis is the reason that you should never drink seawater—or any concentrated solution. Seawater draws water *out of the body* as it passes through the stomach and intestines, resulting in diarrhea and further dehydration (Figure 13.16). We can think of seawater as a *thirsty solution*—one that draws more water to itself. Consequently, seawater should never be consumed as drinking water.

**Figure 13.16 The Effect of Drinking Seawater**

Seawater is a more concentrated solution than the fluids in body cells. As a result, when seawater flows through the digestive tract, it draws water out of the surrounding tissues.





Seawater in intestinal tract  
—concentrated solution

Direction of water flow

### Example 13.10 Osmotic Pressure

The osmotic pressure of a solution containing 5.87 mg of an unknown protein per 10.0 mL of solution is 2.45 torr at 25 °C. Find the molar mass of the unknown protein.

**SORT** The problem states that a solution of an unknown protein contains 5.87 mg of the protein per 10.0 mL of solution. You are also given the osmotic pressure of the solution at a particular temperature and asked to find the molar mass of the unknown protein.

**GIVEN:**

5.87 mg of protein  
10.0-mL solution  
 $\Pi = 2.45$  torr  
 $T = 25^\circ\text{C}$

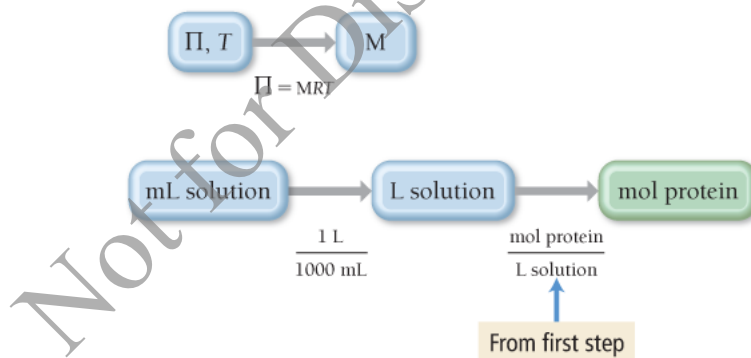
**FIND:** molar mass of protein (g/mol)

**STRATEGIZE** Step 1: Use the given osmotic pressure and temperature to find the molarity of the protein solution.

Step 2: Use the molarity calculated in Step 1 to find the number of moles of protein in 10 mL of solution.

Step 3: Finally, use the number of moles of the protein calculated in Step 2 and the given mass of the protein in 10.0 mL of solution to find the molar mass.

**CONCEPTUAL PLAN**



$$\text{molar mass} = \frac{\text{mass protein}}{\text{moles protein}}$$

**RELATIONSHIPS USED**  $\Pi = MRT$  (osmotic pressure equation)

**SOLVE**

Step 1: Begin by solving the osmotic pressure equation for molarity and substituting in the required quantities in the correct units to calculate  $M$ .

Step 2: Begin with the given volume, convert to liters, then use the molarity to find the number of moles of protein.

Step 3: Use the given mass and the number of moles from Step 2 to calculate the molar mass of the

protein.

### SOLUTION

$$\begin{aligned}\Pi &= MRT \\ M &= \frac{\Pi}{RT} = \frac{2.45 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}}}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} (298 \text{ K})} \\ &= 1.31 \times 10^{-4} \text{ M}\end{aligned}$$

$$10.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.31 \times 10^{-4} \text{ mol}}{\text{L}} = 1.31 \times 10^{-6} \text{ mol}$$

$$\begin{aligned}\text{molar mass} &= \frac{\text{mass protein}}{\text{moles protein}} \\ &= \frac{5.87 \times 10^{-3} \text{ g}}{1.31 \times 10^{-6} \text{ mol}} \\ &= 4.45 \times 10^3 \text{ g/mol}\end{aligned}$$

**CHECK** The units of the answer are correct. The magnitude might seem a little high initially, but proteins are large molecules and therefore have high molar masses.

**FOR PRACTICE 13.10** Calculate the osmotic pressure (in atm) of a solution containing 1.50 g of ethylene glycol ( $\text{C}_2\text{H}_6\text{O}_2$ ) in 50.0 mL of solution at 25 °C.

Not for Distribution

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