Sample Exercise 13.1 Predicting Solubility Patterns

Predict whether each of the following substances is more likely to dissolve in the nonpolar solvent carbon tetrachloride (CCl_4) or in water: C_7H_{16}, Na_2SO_4, HCl, and I_2.

Solution

Analyze We are given two solvents, one that is nonpolar (CCl_4) and the other that is polar (H_2O), and asked to determine which will be the better solvent for each solute listed.

Plan By examining the formulas of the solutes, we can predict whether they are ionic or molecular. For those that are molecular, we can predict whether they are polar or nonpolar. We can then apply the idea that the nonpolar solvent will be better for the nonpolar solutes, whereas the polar solvent will be better for the ionic and polar solutes.

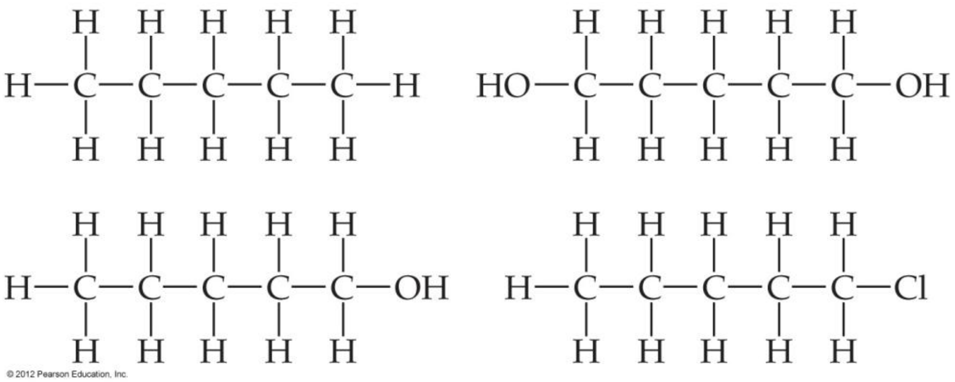
Solve C_7H_{16} is a hydrocarbon, so it is molecular and nonpolar. Na_2SO_4, a compound containing a metal and nonmetals, is ionic. HCl, a diatomic molecule containing two nonmetals that differ in electronegativity, is polar. I_2, a diatomic molecule with atoms of equal electronegativity, is nonpolar. We would therefore predict that C_7H_{16} and I_2 (the nonpolar solutes) would be more soluble in the nonpolar CCl_4 than in polar H_2O, whereas water would be the better solvent for Na_2SO_4 and HCl (the ionic and polar covalent solutes).

Sample Exercise 13.1 Predicting Solubility Patterns

Continued

Practice Exercise

Arrange the following substances in order of increasing solubility in water:



$$\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array} \qquad \begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ HO & - & C & - & C & - & C & - & C & - & C & - & OH \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array}$$

$$\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & OH \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array} \qquad \begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & C & - & Cl \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array}$$

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Answer: $C_5H_{12} < C_5H_{11}Cl < C_5H_{11}OH < C_5H_{10}(OH)_2$ (in order of increasing polarity and hydrogen-bonding ability)

Sample Exercise 13.2 A Henry’s Law Calculation

Calculate the concentration of CO_2 in a soft drink that is bottled with a partial pressure of CO_2 of 4.0 atm over the liquid at 25 °C. The Henry’s law constant for CO_2 in water at this temperature is 3.4×10^{-2} mol/L-atm.

Solution

Analyze We are given the partial pressure of CO_2, P_{CO_2}, and the Henry’s law constant, k, and asked to calculate the concentration of CO_2 in the solution.

Plan With the information given, we can use Henry’s law, Equation 13.4, to calculate the solubility, S_{CO_2}.

$$S_g = kP_g \qquad [13.4]$$

Solve $S_{CO_2} = kP_{CO_2} = (3.4 \times 10^{-2}\ \text{mol/L-atm})(4.0\ \text{atm}) = 0.14\ \text{mol/L} = 0.14\ M$

Check The units are correct for solubility, and the answer has two significant figures consistent with both the partial pressure of CO_2 and the value of Henry’s constant.

Practice Exercise

Calculate the concentration of CO_2 in a soft drink after the bottle is opened and equilibrates at 25 °C under a CO_2 partial pressure of 3.0×10^{-4} atm.

Answer: $1.0 \times 10^{-5}\ M$

Sample Exercise 13.3 Converting Units of Temperature

(a) A solution is made by dissolving 13.5 g of glucose ($C_6H_{12}O_6$) in 0.100 kg of water. What is the mass percentage of solute in this solution? **(b)** A 2.5-g sample of groundwater was found to contain 5.4 μg of Zn^{2+}. What is the concentration of Zn^{2+} in parts per million?

Solution

(a) Analyze We are given the number of grams of solute (13.5 g) and the number of grams of solvent (0.100 kg = 100 g). From this we must calculate the mass percentage of solute.

Plan We can calculate the mass percentage by using Equation 13.5. The mass of the solution is the sum of the mass of solute (glucose) and the mass of solvent (water).

$$\text{Mass \% of component} = \frac{\text{mass of component in soln}}{\text{total mass of soln}} \times 100 \qquad [13.5]$$

Solve

$$\text{Mass \% of glucose} = \frac{\text{mass glucose}}{\text{mass soln}} \times 100 = \frac{13.5\text{ g}}{13.5\text{ g} + 100\text{ g}} \times 100 = 11.9\%$$

Comment The mass percentage of water in this solution is (100 – 11.9)% = 88.1%.

Sample Exercise 13.3 Converting Units of Temperature

Continued

(b) Analyze In this case we are given the number of micrograms of solute. Because 1 μg is 1×10^{-6} g, 5.4 μg = 5.4×10^{-6} g

Plan We calculate the parts per million using Equation 13.6.

$$\text{ppm of component} = \frac{\text{mass of component in soln}}{\text{total mass of soln}} \times 10^6 \qquad [13.6]$$

Solve

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of soln}} \times 10^6 = \frac{5.4 \times 10^{-6}\ \text{g}}{2.5\ \text{g}} \times 10^6 = 2.2\ \text{ppm}$$

Practice Exercise

(a) Calculate the mass percentage of NaCl in a solution containing 1.50 g of NaCl in 50.0 g of water.
(b) A commercial bleaching solution contains 3.62 mass % sodium hypochlorite, NaOCl. What is the mass of NaOCl in a bottle containing 2.50 kg of bleaching solution?

Answer: **(a)** 2.91%, **(b)** 90.5 g of NaOCl

Sample Exercise 13.4 Calculation of Molarity

A solution is made by dissolving 4.35 g glucose ($C_6H_{12}O_6$) in 25.0 mL of water at 25 °C. Calculate the molality of glucose in the solution. Water has a density of 1.00 g/mL.

Solution

Analyze We are asked to calculate a molality. To do this, we must determine the number of moles of solute (glucose) and the number of kilograms of solvent (water).

Plan We use the molar mass of $C_6H_{12}O_6$ to convert grams to moles. We use the density of water to convert milliliters to kilograms. The molality equals the number of moles of solute divided by the number of kilograms of solvent (Equation 13.9).

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}} \qquad [13.9]$$

Solve Use the molar mass of glucose, 180.2 g/mol, to convert grams to moles:

$$\text{Mol } C_6H_{12}O_6 = (4.35 \text{ g } C_6H_{12}O_6)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.2 \text{ g } C_6H_{12}O_6}\right) = 0.0241 \text{ mol } C_6H_{12}O_6$$

Because water has a density of 1.00 g/mL, the mass of the solvent is

$$(25.0 \text{ mL})(1.00 \text{ g/mL}) = 25.0 \text{ g} = 0.0250 \text{ kg}$$

Finally, use Equation 13.9 to obtain the molality:

$$\text{Molality of } C_6H_{12}O_6 = \frac{0.0241 \text{ mol } C_6H_{12}O_6}{0.0250 \text{ kg } H_2O} = 0.964\ m$$

Practice Exercise

What is the molality of a solution made by dissolving 36.5 g of naphthalene ($C_{10}H_8$) in 425 g of toluene (C_7H_8)?

Answers: 0.670 *m*

Sample Exercise 13.5 Calculation of Mole Fraction and Molarity

An aqueous solution of hydrochloric acid contains 36% HCl by mass. **(a)** Calculate the mole fraction of HCl in the solution. **(b)** Calculate the molality of HCl in the solution.

Solution

Analyze We are asked to calculate the concentration of the solute, HCl, in two related concentration units, given only the percentage by mass of the solute in the solution.

Plan In converting concentration units based on the mass or moles of solute and solvent (mass percentage, mole fraction, and molality), it is useful to assume a certain total mass of solution. Let's assume that there is exactly 100 g of solution. Because the solution is 36% HCl, it contains 36 g of HCl and (100 – 36) g = 64 g of H_20. We must convert grams of solute (HCl) to moles to calculate either mole fraction or molality. We must convert grams of solvent (H_2O) to moles to calculate mole fractions and to kilograms to calculate molality.

Solve (a) To calculate the mole fraction of HCl, we convert the masses of HCl and H_2O to moles and then use Equation 13.7:

$$\text{Moles HCl} = (36\text{ g HCl})\left(\frac{1\text{ mol HCl}}{36.5\text{ g HCl}}\right) = 0.99\text{ mol HCl}$$

$$\text{Moles H}_2\text{O} = (64\text{ g H}_2\text{O})\left(\frac{1\text{ mol H}_2\text{O}}{18\text{ g H}_2\text{O}}\right) = 3.6\text{ mol H}_2\text{O}$$

$$X_{\text{HCl}} = \frac{\text{moles HCl}}{\text{moles H}_2\text{O} + \text{moles HCl}} = \frac{0.99}{3.6 + 0.99} = \frac{0.99}{4.6} = 0.22$$

$$\text{Mole fraction of component} = \frac{\text{moles of component}}{\text{total moles of all components}} \qquad [13.7]$$

Sample Exercise 13.5 Calculation of Mole Fraction and Molarity

Continued

(b) To calculate the molality of HCl in the solution, we use Equation 13.9. We calculated the number of moles of HCl in part (a), and the mass of solvent is 64 g = 0.064 kg:

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}} \qquad [13.9]$$

$$\text{Molality of HCl} = \frac{0.99 \text{ mol HCl}}{0.064 \text{ kg } H_2O} = 15\ m$$

Practice Exercise

A commercial bleach solution contains 3.62 mass % NaOCl in water.
Calculate **(a)** the mole fraction and **(b)** the molality of NaOCl in the solution.

Answer: **(a)** 9.00×10^{-3}, **(b)** 0.505 *m*

Sample Exercise 13.6 Calculation of Molarity Using the Density of the Solution

A solution with a density of 0.876 g/mL contains 5.0 g of toluene (C_7H_8) and 225 g of benzene. Calculate the molarity of the solution.

Solution

Analyze Our goal is to calculate the molarity of a solution, given the masses of solute (5.0 g) and solvent (225 g) and the density of the solution (0.876 g/mL).

Plan The molarity of a solution is the number of moles of solute divided by the number of liters of solution (Equation 13.8). The number of moles of solute (C_7H_8) is calculated from the number of grams of solute and its molar mass. The volume of the solution is obtained from the mass of the solution (mass of solute + mass of solvent = 5.0 g + 225 g = 230 g) and its density.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of soln}} \qquad [13.8]$$

Solve The number of moles of solute is

$$\text{Moles } C_7H_8 = (5.0\text{ g } C_7H_8)\left(\frac{1\text{ mol } C_7H_8}{92\text{ g } C_7H_8}\right) = 0.054\text{ mol}$$

The density of the solution is used to convert the mass of the solution to its volume:

$$\text{Milliliters soln} = (230\text{ g})\left(\frac{1\text{ mL}}{0.876\text{ g}}\right) = 263\text{ mL}$$

Molarity is moles of solute per liter of solution:

$$\text{Molarity} = \left(\frac{\text{moles } C_7H_8}{\text{liter soln}}\right) = \left(\frac{0.054\text{ mol } C_7H_8}{263\text{ mL soln}}\right)\left(\frac{1000\text{ mL soln}}{1\text{ L soln}}\right) = 0.21\ M$$

Sample Exercise 13.6 Calculation of Molarity Using the Density of the Solution

Continued

Check The magnitude of our answer is reasonable. Rounding moles to 0.05 and liters to 0.25 gives a molarity of

$$(0.05 \text{ mol})/(0.25 \text{ L}) = 0.2\ M$$

The units for our answer (mol/L) are correct, and the answer, 0.21, has two significant figures, corresponding to the number of significant figures in the mass of solute (2).

Comment Because the mass of the solvent (0.225 kg) and the volume of the solution (0.263) are similar in magnitude, the molarity and molality are also similar in magnitude:

$$(0.054 \text{ mol } C_7H_8)/(0.225 \text{ kg solvent}) = 0.24\ m$$

Practice Exercise

A solution containing equal masses of glycerol ($C_3H_8O_3$) and water has a density of 1.10 g/mL. Calculate **(a)** the molality of glycerol, **(b)** the mole fraction of glycerol, **(c)** the molarity of glycerol in the solution.

Answer: **(a)** 10.9 *m*, **(b)** $XC_3H_8O_3 = 0.163$, **(c)** 5.97 *M*

Sample Exercise 13.7 Calculation of Vapor Pressure of a Solution

Glycerin ($C_3H_8O_3$) is a nonvolatile nonelectrolyte with a density of 1.26 g/mL at 25 °C. Calculate the vapor pressure at 25 °C of a solution made by adding 50.0 mL of glycerin to 500.0 mL of water. The vapor pressure of pure water at 25 °C is 23.8 torr (Appendix B), and its density is 1.00 g/mL.

Solution

Analyze Our goal is to calculate the vapor pressure of a solution, given the volumes of solute and solvent and the density of the solute.

Plan We can use Raoult's law (Equation 13.10) to calculate the vapor pressure of a solution. The mole fraction of the solvent in the solution, X_{solvent}, is the ratio of the number of moles of solvent (H_2O) to total moles of solution (moles $C_3H_8O_3$ + moles H_2O).

$$P_{\text{solution}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}} \qquad [13.10]$$

Solve To calculate the mole fraction of water in the solution, we must determine the number of moles of $C_3H_8O_3$ and H_2O:

$$\text{Moles } C_3H_8O_3 = (50.0 \text{ mL } C_3H_8O_3)\left(\frac{1.26 \text{ g } C_3H_8O_3}{1 \text{ mL } C_3H_8O_3}\right)\left(\frac{1 \text{ mol } C_3H_8O_3}{92.1 \text{ g } C_3H_8O_3}\right) = 0.684 \text{ mol}$$

$$\text{Moles } H_2O = (500.0 \text{ mL } H_2O)\left(\frac{1.00 \text{ g } H_2O}{1 \text{ mL } H_2O}\right)\left(\frac{1 \text{ mol } H_2O}{18.0 \text{ g } H_2O}\right) = 27.8 \text{ mol}$$

$$X_{H_2O} = \frac{\text{mol } H_2O}{\text{mol } H_2O + \text{mol } C_3H_8O_3} = \frac{27.8}{27.8 + 0.684} = 0.976$$

We now use Raoult's law to calculate the vapor pressure of water for the solution:

$$P_{H_2O} = X_{H_2O} P^{\circ}_{H_2O} = (0.976)(23.8 \text{ torr}) = 23.2 \text{ torr}$$

Sample Exercise 13.7 Calculation of Vapor Pressure of a Solution

Continued

Comment The vapor pressure of the solution has been lowered by 23.8 torr – 23.2 torr = 0.6 torr relative to that of pure water. The vapor-pressure lowering can be calculated directly using Equation 13.11 together with the mole fraction of the solute, $C_3H_8O_3$: $\Delta P = X_{C_3H_8O_3}P^{\circ}_{H_2O}$ = (0.024)(23.8 torr) = 0.57 torr. Notice that the use of Equation 13.11 gives one more significant figure than the number obtained by subtracting the vapor pressure of the solution from that of the pure solvent.

$$\Delta P = X_{\text{solute}}P^{\circ}_{\text{solvent}} \qquad [13.11]$$

Practice Exercise

The vapor pressure of pure water at 110 °C is 1070 torr. A solution of ethylene glycol and water has a vapor Pressure of 1.00 atm at 110 °C. Assuming that Raoult’s law is obeyed, what is the mole fraction of ethylene glycol in the solution?

Answer: 0.290

Sample Exercise 13.8 Calculating Of Boiling-Point Elevation and Freezing-Point Depression

Automotive antifreeze consists of ethylene glycol, $CH_2(OH)CH_2(OH)$, a nonvolatile nonelectrolyte. Calculate the boiling point and freezing point of a 25.0 mass % solution of ethylene glycol in water.

Solution

Analyze We are given that a solution contains 25.0 mass % of a nonvolatile, nonelectrolyte solute and asked to calculate the boiling and freezing points of the solution. To do this, we need to calculate the boiling-point elevation and freezing-point depression.

Plan To calculate the boiling-point elevation and the freezing-point depression using Equations 13.12 and 13.13, we must express the concentration of the solution as molality. Let's assume for convenience that we have 1000 g of solution. Because the solution is 25.0 mass % ethylene glycol, the masses of ethylene glycol and water in the solution are 250 and 750 g, respectively. Using these quantities, we can calculate the molality of the solution, which we use with the molal boiling-point-elevation and freezing-point-depression constants (Table 13.3) to calculate ΔT_b and ΔT_f .We add ΔT_b to the boiling point and subtract ΔT_f from the freezing point of the solvent to obtain the boiling point and freezing point of the solution.

$$\Delta T_b = K_b m \qquad [13.12]$$

$$\Delta T_f = K_f m \qquad [13.13]$$

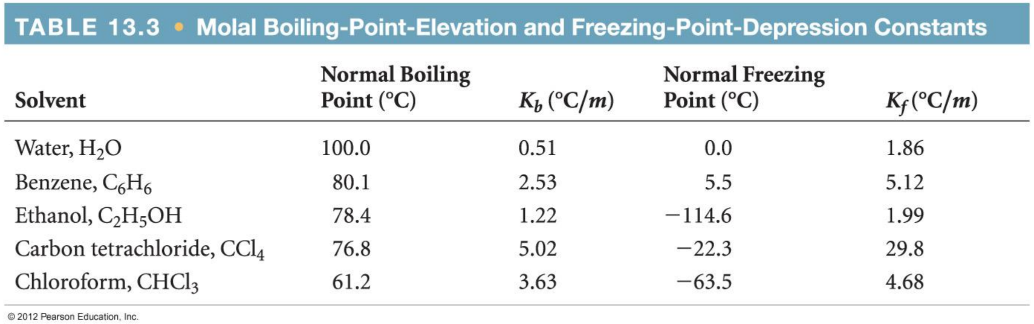

TABLE 13.3 • Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H_2O	100.0	0.51	0.0	1.86
Benzene, C_6H_6	80.1	2.53	5.5	5.12
Ethanol, C_2H_5OH	78.4	1.22	−114.6	1.99
Carbon tetrachloride, CCl_4	76.8	5.02	−22.3	29.8
Chloroform, $CHCl_3$	61.2	3.63	−63.5	4.68

Sample Exercise 13.8 Calculating Of Boiling-Point Elevation and Freezing-Point Depression

Continued

Solve The molality of the solution is calculated as follows:

$$\text{Molality} = \frac{\text{moles } C_2H_6O_2}{\text{kilograms } H_2O} = \left(\frac{250 \text{ g } C_2H_6O_2}{750 \text{ g } H_2O}\right)\left(\frac{1 \text{ mol } C_2H_6O_2}{62.1 \text{ g } C_2H_6O_2}\right)\left(\frac{1000 \text{ g } H_2O}{1 \text{ kg } H_2O}\right)$$

$$= 5.37\ m$$

We can now use Equations 13.12 and 13.13 to calculate the changes in the boiling and freezing points:

$$\Delta T_b = K_b m \qquad [13.12]$$

$$\Delta T_f = K_f m \qquad [13.13]$$

$\Delta T_b = K_b m = (0.51\ °\text{C}/m)(5.37\ m) = 2.7\ °\text{C}$

$\Delta T_f = K_f m = (1.86\ °\text{C}/m)(5.37\ m) = 10.0\ °\text{C}$

Hence, the boiling and freezing points of the solution are

Boiling point = (normal bp of solvent) + ΔT_b
= 100.0 °C + 2.7 °C = 102.7 °C

Freezing point = (normal fp of solvent) $-\Delta Tf$
= 0.0 °C – 10.0 °C = –10.0 °C

Comment Notice that the solution is a liquid over a larger temperature range than the pure solvent.

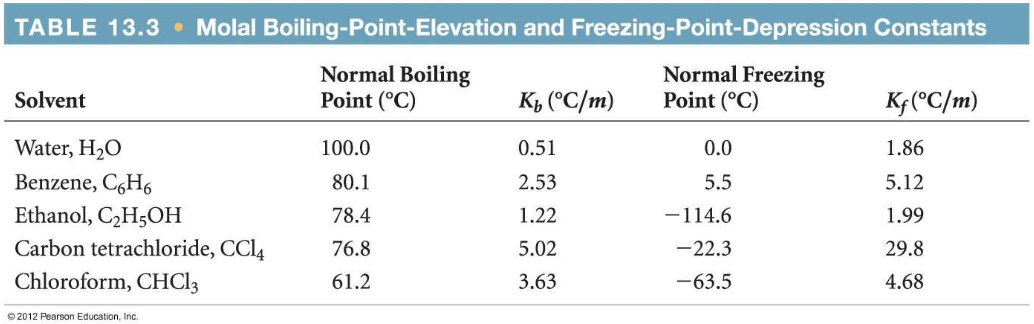

TABLE 13.3 • Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H_2O	100.0	0.51	0.0	1.86
Benzene, C_6H_6	80.1	2.53	5.5	5.12
Ethanol, C_2H_5OH	78.4	1.22	−114.6	1.99
Carbon tetrachloride, CCl_4	76.8	5.02	−22.3	29.8
Chloroform, $CHCl_3$	61.2	3.63	−63.5	4.68

Practice Exercise

Calculate the freezing point of a solution containing 0.600 kg of $CHCl_3$ and 42.0 g of eucalyptol ($C_{10}H_{18}O$), a fragrant substance found in the leaves of eucalyptus trees. (See Table 13.3.)

Answer: –65.6 °C

Sample Exercise 13.9 Freezing-Point Depression in Aqueous Solutions

List the following aqueous solutions in order of their expected freezing point: 0.050 m $CaCl_2$, 0.15 m NaCl, 0.10 m HCl, 0.050 m CH_3COOH, 0.10 m $C_{12}H_{22}O_{11}$.

Solution

Analyze We must order five aqueous solutions according to expected freezing points, based on molalities and the solute formulas.

Plan The lowest freezing point will correspond to the solution with the greatest concentration of solute particles. To determine the total concentration of solute particles in each case, we must determine whether the substance is a nonelectrolyte or an electrolyte and consider the number of ions formed when an electrolyte ionizes.

Solve $CaCl_2$, NaCl, and HCl are strong electrolytes, CH_3COOH (acetic acid) is a weak electrolyte, and $C_{12}H_{22}O_{11}$ is a nonelectrolyte. The molality of each solution in total particles is as follows:

$$0.050\ m\ \mathrm{CaCl_2} \Rightarrow 0.050\ m \text{ in } \mathrm{Ca^{2+}} \text{ and } 0.10\ m \text{ in } \mathrm{Cl^-} \Rightarrow 0.15\ m \text{ in particles}$$

$$0.15\ m\ \mathrm{NaCl} \Rightarrow 0.15\ m\ \mathrm{Na^+} \text{ and } 0.15\ m \text{ in } \mathrm{Cl^-} \Rightarrow 0.30\ m \text{ in particles}$$

$$0.10\ m\ \mathrm{HCl} \Rightarrow 0.10\ m\ \mathrm{H^+} \text{ and } 0.10\ m \text{ in } \mathrm{Cl^-} \Rightarrow 0.20\ m \text{ in particles}$$

$$0.050\ m\ \mathrm{CH_3COOH} \Rightarrow \text{weak electrolyte} \Rightarrow \text{between } 0.050\ m \text{ and } 0.10\ m \text{ in particles}$$

$$0.10\ m\ \mathrm{C_{12}H_{22}O_{11}} \Rightarrow \text{nonelectrolyte} \Rightarrow 0.10\ m \text{ in particles}$$

Because the freezing points depend on the total molality of particles in solution, the expected ordering is 0.15 m NaCl (lowest freezing point), 0.10 m HCl, 0.050 m $CaCl_2$, 0.10 m $C_{12}H_{22}O_{11}$, and 0.050 m CH_3COOH (highest freezing point).

Sample Exercise 13.9 Freezing-Point Depression in Aqueous Solutions

Continued

Practice Exercise

Which of the following solutes will produce the largest increase in boiling point upon addition to 1 kg of water: 1 mol of $Co(NO_3)_2$, 2 mol of KCl, 3 mol of ethylene glycol ($C_2H_6O_2$)?

Answer: 2 mol of KCl because it contains the highest concentration of particles, 2 *m* K^+ and 2 *m* Cl^-, giving 4 *m* in all

Sample Exercise 13.10 Calculating Involving Osmotic Pressure

The average osmotic pressure of blood is 7.7 atm at 25 °C. What molarity of glucose ($C_6H_{12}O_6$) will be isotonic with blood?

Solution

Analyze We are asked to calculate the concentration of glucose in water that would be isotonic with blood, given that the osmotic pressure of blood at 25 °C is 7.7 atm.

Plan Because we are given the osmotic pressure and temperature, we can solve for the concentration, using Equation 13.14.

$$\Pi = \left(\frac{n}{V}\right)RT = MRT \qquad [13.14]$$

Solve

$$\Pi = MRT$$

$$M = \frac{\Pi}{RT} = \frac{7.7\text{ atm}}{\left(0.0821\frac{\text{L-atm}}{\text{mol-K}}\right)(298\text{ K})} = 0.31\ M$$

Comment In clinical situations the concentrations of solutions are generally expressed as mass percentages. The mass percentage of a 0.31 *M* solution of glucose is 5.3%. The concentration of NaCl that is isotonic with blood is 0.16 *M*, because NaCl ionizes to form two particles, Na^+ and Cl^- (a 0.155 *M* solution of NaCl is 0.310 *M* in particles). A 0.16 *M* solution of NaCl is 0.9 mass % in NaCl. This kind of solution is known as a physiological saline solution.

Sample Exercise 13.10 Calculating Involving Osmotic Pressure

Continued

Practice Exercise

What is the osmotic pressure at 20 °C of a 0.0020 *M* sucrose ($C_{12}H_{22}O_{11}$) solution?

Answer: 0.048 atm, or 37 torr

Sample Exercise 13.11 Molar Mass from Freezing-Point Depression

A solution of an unknown nonvolatile nonelectrolyte was prepared by dissolving 0.250 g of the substance in 40.0 g of CCl_4. The boiling point of the resultant solution was 0.357 °C higher than that of the pure solvent. Calculate the molar mass of the solute.

Solution

Analyze Our goal is to calculate the molar mass of a solute based on knowledge of the boiling-point elevation of its solution in CCl_4, $\Delta T_b = 0.357$ °C, and the masses of solute and solvent. Table 13.3 gives K_b for the solvent (CCl_4), $K_b = 5.02$ °C/m.

Plan We can use Equation 13.12, $\Delta T_b = K_b m$, to calculate the Molality of the solution. Then we can use molality and the quantity of solvent (40.0 g CCl_4) to calculate the number of moles of solute. Finally, the molar mass of the solute equals the number of grams per mole, so we divide the number of grams of solute (0.250 g) by the number of moles we have just calculated.

$$\Delta T_b = K_b m \qquad [13.12]$$

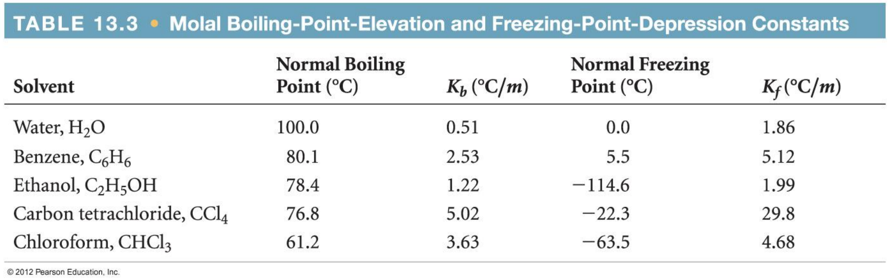

TABLE 13.3 • Molal Boiling-Point-Elevation and Freezing-Point-Depression Constants

Solvent	Normal Boiling Point (°C)	K_b (°C/m)	Normal Freezing Point (°C)	K_f (°C/m)
Water, H_2O	100.0	0.51	0.0	1.86
Benzene, C_6H_6	80.1	2.53	5.5	5.12
Ethanol, C_2H_5OH	78.4	1.22	−114.6	1.99
Carbon tetrachloride, CCl_4	76.8	5.02	−22.3	29.8
Chloroform, $CHCl_3$	61.2	3.63	−63.5	4.68

Sample Exercise 13.11 Molar Mass from Freezing-Point Depression

Continued

Solve From Equation 13.12 we have

$$\Delta T_b = K_b m \qquad [13.12]$$

$$\text{Molality} = \frac{\Delta T_b}{K_b} = \frac{0.357\ ^\circ\text{C}}{5.02\ ^\circ\text{C}/m} = 0.0711\ m$$

Thus, the solution contains 0.0711 mol of solute per kilogram of solvent. The solution was prepared using of solvent (CCl_4). The number of moles of solute in the solution is therefore

$$(0.0400\ \text{kg}\ CCl_4)\left(0.0711\ \frac{\text{mol solute}}{\text{kg}\ CCl_4}\right) = 2.84 \times 10^{-3}\ \text{mol solute}$$

The molar mass of the solute is the number of grams per mole of the substance:

$$\text{Molar mass} = \frac{0.250\ \text{g}}{2.84 \times 10^{-3}\ \text{mol}} = 88.0\ \text{g/mol}$$

Practice Exercise

Camphor ($C_{10}H_{16}O$) melts at 179.8 °C, and it has a particularly large freezing-point-depression constant, $K_f = 40.0\ ^\circ\text{C}/m$. When 0.186 g of an organic substance of unknown molar mass is dissolved in 22.01 g of liquid camphor, the freezing point of the mixture is found to be 176.7 °C. What is the molar mass of the solute?

Answer: 110 g/mol

Sample Exercise 13.12 Molar Mass from Osmotic Pressure

The osmotic pressure of an aqueous solution of a certain protein was measured to determine the protein's molar mass. The solution contained 3.50 mg of protein dissolved in sufficient water to form 5.00 mL of solution. The osmotic pressure of the solution at 25 °C was found to be 1.54 torr. Treating the protein as a nonelectrolyte, calculate its molar mass.

Solution

Analyze Our goal is to calculate the molar mass of a high-molecular-mass protein, based on its osmotic pressure and a knowledge of the mass of protein and solution volume.

Plan The temperature ($T = 25\ °\text{C}$) and osmotic pressure ($\Pi = 1.54$ torr) are given, and we know the value of R so we can use Equation 13.14 to calculate the molarity of the solution, M. In doing so, we must convert temperature from °C to K and the osmotic pressure from torr to atm. We then use the molarity and the volume of the solution (5.00 mL) to determine the number of moles of solute. Finally, we obtain the molar mass by dividing the mass of the solute (3.50 mg) by the number of moles of solute.

$$\Pi = \left(\frac{n}{V}\right)RT = MRT \qquad [13.14]$$

Solve Solving Equation 13.14 for molarity gives

$$\text{Molarity} = \frac{\Pi}{RT} = \frac{(1.54\ \text{torr})\left(\dfrac{1\ \text{atm}}{760\ \text{torr}}\right)}{\left(0.0821\dfrac{\text{L-atm}}{\text{mol-K}}\right)(298\ \text{K})} = 8.28 \times 10^{-5}\ \frac{\text{mol}}{\text{L}}$$

Sample Exercise 13.12 Molar Mass from Osmotic Pressure

Continued

Because the volume of the solution is 5.00 mL = 5.00×10^{-3} L, the number of moles of protein must be

$$\text{Moles} = (8.28 \times 10^{-5}\ \text{mol/L})(5.00 \times 10^{-3}\ \text{L}) = 4.14 \times 10^{-7}\ \text{mol}$$

The molar mass is the number of grams per mole of the substance. The sample has a mass of 3.50 mg = 3.50×10^{-3} g. The molar mass is the number of grams divided by the number of moles:

$$\text{Molar mass} = \frac{\text{grams}}{\text{moles}} = \frac{3.50 \times 10^{-3}\ \text{g}}{4.14 \times 10^{-7}\ \text{mol}} = 8.45 \times 10^{3}\ \text{g/mol}$$

Comment Because small pressures can be measured easily and accurately, osmotic pressure measurements provide a useful way to determine the molar masses of large molecules.

Practice Exercise

A sample of 2.05 g of polystyrene of uniform polymer chain length was dissolved in enough toluene to form 0.100 L of solution. The osmotic pressure of this solution was found to be 1.21 kPa at 25 °C. Calculate the molar mass of the polystyrene.

Answer: 4.20 × 104 g/mol