
12. SOLUTIONS

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

- 12.1 An example of a solid solution prepared from a liquid and a solid is a dental filling made of liquid mercury and solid silver.
- 12.2 The $\text{C}_4\text{H}_9\text{OH}$ molecules will be more soluble in water because their $-\text{OH}$ ends can form hydrogen bonds with water.
- 12.3 The Na^+ ion has a larger energy of hydration because its ionic radius is smaller, giving Na^+ a more concentrated electric field than K^+ .
- 12.4 Write Henry's law ($S = k_{\text{H}}P$) for 159 mmHg (P_2), and divide it by Henry's law for one atm, or 760 mmHg (P_1). Then, substitute the experimental values of P_1 , P_2 , and S_1 to solve for S_2 .

$$\frac{S_2}{S_1} = \frac{k_{\text{H}}S_2}{k_{\text{H}}P_1} = \frac{P_2}{P_1}$$

(continued)

Solving for S_2 gives

$$S_2 = \frac{P_2 S_1}{P_1} = \frac{(159 \text{ mmHg})(0.0404 \text{ g O}_2/\text{L})}{760 \text{ mmHg}} = 8.452 \times 10^{-3}$$

$$= 8.45 \times 10^{-3} \text{ g O}_2/\text{L}$$

- 12.5 The mass of HCl in 20.2 percent HCl (0.202 = fraction of HCl) is

$$0.202 \times 35.0 \text{ g} = 7.070 = 7.07 \text{ g HCl}$$

The mass of H_2O in 20.2 percent HCl is

$$35.0 \text{ g solution} - 7.07 \text{ g HCl} = 27.93 = 27.9 \text{ g H}_2\text{O}$$

- 12.6 Calculate the moles of toluene using its molar mass of 92.14 g/mol:

$$35.6 \text{ g toluene} \times \frac{1 \text{ mol toluene}}{92.14 \text{ g toluene}} = 0.3863 \text{ mol toluene}$$

To calculate molality, divide the moles of toluene by the mass in kg of the solvent (C_6H_6):

$$\text{Molality} \times \frac{0.3863 \text{ mol toluene}}{0.125 \text{ kg solvent}} = 3.0904 = 3.09 \text{ m toluene}$$

- 12.7 The number of moles of toluene = 0.3863 (previous exercise); the number of moles of benzene is

$$125 \text{ g benzene} \times \frac{1 \text{ mol benzene}}{78.11 \text{ g benzene}} = 1.6003 \text{ mol benzene}$$

The total number of moles is $1.6003 + 0.3863 = 1.9866$, and the mole fractions are

$$\text{Mol fraction benzene} \times \frac{1.6003 \text{ mol benzene}}{1.9866 \text{ mol}} = 0.80554 = 0.806$$

$$\text{Mol fraction toluene} \times \frac{0.3863 \text{ mol toluene}}{1.9866 \text{ mol}} = 0.1944 = 0.194$$

The sum of the mole fractions = 1.000.

- 12.8 This solution contains 0.120 moles of methanol dissolved in 1.00 kg of ethanol. The number of moles in 1.00 kg of ethanol is

$$1.00 \times 10^3 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.07 \text{ g C}_2\text{H}_5\text{OH}} = 21.706 \text{ mol C}_2\text{H}_5\text{OH}$$

The total number of moles is $21.706 + 0.120 = 21.826$, and the mole fractions are

$$\text{Mol fraction C}_2\text{H}_5\text{OH} = \frac{21.706 \text{ mol C}_2\text{H}_5\text{OH}}{21.826 \text{ mol}} = 0.994501 = 0.995$$

$$\text{Mol fraction CH}_3\text{OH} = \frac{0.120 \text{ mol CH}_3\text{OH}}{21.826 \text{ mol}} = 0.005498 = 0.00550$$

The sum of the mole fractions is 1.000.

- 12.9 One mole of solution contains 0.250 moles methanol and 0.750 moles ethanol. The mass of this amount of ethanol, the solvent, is

$$0.750 \text{ mol C}_2\text{H}_5\text{OH} \times \frac{46.07 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} = 34.55 \text{ g C}_2\text{H}_5\text{OH} \text{ (0.03455 kg)}$$

The molality of methanol in the ethanol solvent is

$$\frac{0.250 \text{ mol CH}_3\text{OH}}{0.03455 \text{ kg C}_2\text{H}_5\text{OH}} = 7.2358 = 7.24 \text{ m CH}_3\text{OH}$$

- 12.10 Assume an amount of solution contains one kilogram of water. The mass of urea in this mass is

$$3.42 \text{ mol urea} \times \frac{60.05 \text{ g urea}}{1 \text{ mol urea}} = 205.4 \text{ g urea}$$

The total mass of solution is $205.4 + 1000.0 \text{ g} = 1205.4 \text{ g}$. The volume and molarity are

$$\text{Volume of solution} = 1205.4 \text{ g} \times \frac{1 \text{ mL}}{1.045 \text{ g}} = 1153.49 \text{ mL} = 1.15349 \text{ L}$$

$$\text{Molarity} = \frac{3.42 \text{ mol urea}}{1.15349 \text{ L solution}} = 2.9649 \text{ mol/L} = 2.96 \text{ M}$$

12.11 Assume a volume equal to 1.000 L of solution. Then

$$\text{Mass of solution} = 1.029 \text{ g/mL} \times (1.000 \times 10^3 \text{ mL}) = 1029 \text{ g}$$

$$\text{Mass of urea} = 2.00 \text{ mol urea} \times \frac{60.05 \text{ g urea}}{1 \text{ mol urea}} = 120.1 \text{ g urea}$$

$$\text{Mass of water} = (1029 - 120.1) \text{ g} = 908.9 \text{ g water} \text{ (0.9089 kg)}$$

$$\text{Molality} = \frac{2.00 \text{ mol urea}}{0.9089 \text{ kg solvent}} = 2.2004 = 2.20 \text{ m urea}$$

12.12 Calculate the moles of naphthalene and moles of chloroform:

$$0.515 \text{ g C}_{10}\text{H}_8 \times \frac{1 \text{ mol C}_{10}\text{H}_8}{128.17 \text{ g C}_{10}\text{H}_8} = 0.004018 \text{ mol C}_{10}\text{H}_8$$

$$60.8 \text{ g CHCl}_3 \times \frac{1 \text{ mol CHCl}_3}{119.38 \text{ g CHCl}_3} = 0.50929 \text{ mol CHCl}_3$$

The total number of moles is $0.004018 + 0.50929 = 0.5133$ mol, and the mole fraction of chloroform is

$$\text{Mol fraction CHCl}_3 = \frac{0.50929 \text{ mol CHCl}_3}{0.5133 \text{ mol}} = 0.9921$$

$$\text{Mol fraction C}_{10}\text{H}_8 = \frac{0.004018 \text{ mol C}_{10}\text{H}_8}{0.5133 \text{ mol}} = 0.007828$$

The vapor-pressure lowering is

$$\Delta P = P^\circ X_{\text{C}_{10}\text{H}_8} = (156 \text{ mmHg})(0.007828) = 1.221 = 1.22 \text{ mmHg}$$

Use Raoult's law to calculate the vapor-pressure of chloroform:

$$P = P^\circ X_{\text{CHCl}_3} = (156 \text{ mmHg})(0.9921) = 154.7 = 155 \text{ mmHg}$$

12.13 Solve for c_m in the freezing-point-depression equation ($\Delta T = K_f c_m$; K_f in Table 12.3):

$$c_m = \frac{\Delta T}{K_f} = \frac{0.150^\circ\text{C}}{1.858^\circ\text{C/m}} = 0.08073 \text{ m}$$

(continued)

Use the molal concentration to solve for the mass of ethylene glycol:

$$\frac{0.08073 \text{ m glycol}}{1 \text{ kg solvent}} \times 0.0378 \text{ kg solvent} = 0.003051 \text{ mol glycol}$$

$$0.003051 \text{ mol glycol} \times \frac{62.1 \text{ g glycol}}{1 \text{ mol glycol}} = 1.894 \times 10^{-1} = 1.89 \times 10^{-1} \text{ g glycol}$$

- 12.14 Calculate the moles of ascorbic acid (vit. C) from the molality, and then divide the mass of 0.930 g by the number of moles to obtain the molar mass:

$$\frac{0.0555 \text{ mol vit. C}}{1 \text{ kg H}_2\text{O}} \times 0.0950 \text{ kg H}_2\text{O} = 0.005272 \text{ mol vit C}$$

$$\frac{0.930 \text{ g vit. C}}{0.005272 \text{ mol vit. C}} = 176.4 = 176 \text{ g/mol}$$

The molecular weight of ascorbic acid, or vitamin C, is 176 amu.

- 12.15 The molal concentration of white phosphorus is

$$c_m = \frac{\Delta T}{K_b} = \frac{0.159^\circ\text{C}}{2.40^\circ\text{C/m}} = 0.06625 \text{ m}$$

The number of moles of white phosphorus (P_x) present in this solution is

$$\frac{0.06625 \text{ mol } P_x}{1 \text{ kg CS}_2} \times 0.0250 \text{ kg CS}_2 = 0.001656 \text{ mol } P_x$$

The molar mass of white phosphorus equals the mass divided by moles:

$$0.205 \text{ g} \div 0.001656 \text{ mol} = 123.77 = 124 \text{ g/mol}$$

Thus, the molecular weight of P_x is 124 amu. The number of P atoms in the molecule of white phosphorus is obtained by dividing the molecular weight by the atomic weight of P:

$$\frac{123.77 \text{ amu } P_x}{30.97 \text{ amu P}} = 3.9965 = 4.00$$

Hence, the molecular formula is P_4 ($x = 4$).

12.16 The number of moles of sucrose is

$$5.0 \text{ g sucrose} \times \frac{1 \text{ mol sucrose}}{342.3 \text{ g sucrose}} = 0.0146 \text{ mol sucrose}$$

The molarity of the solution is

$$\frac{0.0146 \text{ mol sucrose}}{0.100 \text{ L}} = 0.146 \text{ M sucrose}$$

The osmotic pressure, π , is equal to MRT and is calculated as follows:

$$\frac{0.146 \text{ mol sucrose}}{1 \text{ L}} \times \frac{0.0821 \text{ L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \times 293 \text{ K} = 3.51 = 3.5 \text{ atm}$$

12.17 The number of ions from each formula unit is i . Here,

$$i = 1 + 2 = 3$$

The boiling-point elevation is

$$\Delta T_b = K_b c_m = 3 \times \frac{0.512 \text{ }^\circ\text{C}}{\text{m}} \times 0.050 \text{ m} = 0.07681 = 0.077 \text{ }^\circ\text{C}$$

The boiling point of aqueous MgCl_2 is $100.077 \text{ }^\circ\text{C}$.

12.18 AlCl_3 would be most effective in coagulating colloidal sulfur because of the greater magnitude of charge on the Al ion (+3).

■ Answers to Concept Checks

12.1 In each case, the component present in the greatest amount is the solvent.

- Mo is the solute, and Cr is the solvent.
- MgCl_2 is the solute, and water is the solvent.
- N_2 and O_2 are the solutes, and Ar is the solvent.

- 12.2 The two main factors to consider when trying to determine the solubility of an ionic compound in water are ionic size and lattice energy. In this case, the lattice energy for the two compounds is the same; you can discount its effects. Since a smaller cation will have a more concentrated electric field leading to a larger energy of hydration, you would expect AB_2 to have a greater energy of hydration; AB_2 is the more soluble compound.
- 12.3 As the altitude increases, the percentage of oxygen in air decreases, and thus the partial pressure decreases. Above 3500 m, the partial pressure of oxygen in air has decreased to the point that not enough will dissolve in the water to sustain the fish.
- 12.4 In order to boil at a lower temperature than water, the vapor pressure of the solution (water + liquid) must be greater than that of water. To make this solution, you must add a liquid that is both soluble in water and chemically similar to water. It must have a higher vapor pressure than water and a boiling point lower than $100\text{ }^\circ\text{C}$. One possible liquid is ethanol, with a boiling point of $78.3\text{ }^\circ\text{C}$ (Table 12.3).
- 12.5 By the principle of osmosis, in a brine solution, water will flow out of the pickle (lower concentration of ions) into the brine (higher concentration of ions). If the pickles were stored in a water solution, the water (lower concentration of ions) would flow into the pickle (higher concentration of ions) and cause it to swell up and probably burst.
- 12.6 Each of these solutions is a water solution of identical volume (normal boiling point $100\text{ }^\circ\text{C}$) containing a different number of moles of solute. The boiling point of a solution can be determined by the formula $\Delta T_b = iK_b m$. The solution with the largest ΔT_b will have the highest boiling point. Since K_b is a constant, this will be the compound with the largest factor of $i \cdot m$. Also, since the volume is constant, the factor reduces to $i \cdot \text{moles}$. Ideally, all of the compounds will dissolve completely, so NaCl and KBr have $i = 2$, and Na_2SO_4 and MgCl_2 have $i = 3$. This gives

$$\text{For NaCl, } i \cdot \text{moles} = 2 \times 1.5 = 3.0$$

$$\text{For Na}_2\text{SO}_4, i \cdot \text{moles} = 3 \times 1.3 = 3.9$$

$$\text{For MgCl}_2, i \cdot \text{moles} = 3 \times 2.0 = 6.0$$

$$\text{For KBr, } i \cdot \text{moles} = 2 \times 2.0 = 4.0$$

The result is given from highest boiling point to lowest boiling point:



- 12.7 Iron(III) hydroxide is a hydrophobic colloid. As the colloid forms in water, an excess of iron(III) ion (Fe^{3+}) is present on the surface, giving each crystal an excess positive charge. These positively charged crystals repel one another, so aggregation to larger particles of iron(III) hydroxide is prevented. When the electrodes are dipped into the colloidal solution, iron(III) hydroxide precipitates because electrons from the negative electrode neutralize the excess positive charge on the iron(III) hydroxide, allowing larger particles to form (precipitate).

■ Answers to Review Questions

- 12.1 An example of a gaseous solution is air, in which nitrogen (78 percent) acts as a solvent for a gas such as oxygen (21 percent). Recall that the solvent is the component present in greater amount. An example of a liquid solution containing a gas is any carbonated beverage in which water acts as the solvent for carbon dioxide gas. Ethanol in water is an example of a liquid-liquid solution. An example of a solid solution is any gold-silver alloy.
- 12.2 The two factors that explain differences in solubilities are (1) the natural tendency of substances to mix together, or the natural tendency of substances to become disordered, and (2) the relative forces of attraction between solute species (or solvent species) compared to that between the solute and solvent species. The strongest interactions are always achieved.
- 12.3 This is not a case of "like dissolves like." There are strong hydrogen-bonding forces between the water molecules. For the octane to mix with water, hydrogen bonds must be broken and replaced by the much weaker London forces between water and octane. Thus, octane does not dissolve in water because the maximum forces of attraction among molecules are obtained if it does not dissolve.
- 12.4 In most cases, the wide differences in solubility can be explained in terms of the different energies of attraction between ions in the crystal and between ions and water. Hydration energy is used to measure the attraction of ions for water molecules, and lattice energy is used to measure the attraction of positive ions for negative ions in the crystal lattice. An ionic substance is soluble when the hydration energy is much larger than the lattice energy. An ionic substance is insoluble when the lattice energy is much larger than the hydration energy.
- 12.5 A sodium chloride crystal dissolves in water because of two factors. The positive Na^+ ion is strongly attracted to the oxygen (negative end of the water dipole) and dissolves as $\text{Na}^+(\text{aq})$. The negative Cl^- ion is strongly attracted to the hydrogens (positive end of the water dipole) and dissolves as $\text{Cl}^-(\text{aq})$.

- 12.6 When the temperature (energy of a solution) is increased, the solubility of an ionic compound usually increases. A number of salts are exceptions to this rule, particularly a number of calcium salts such as calcium acetate, calcium sulfate, and calcium hydroxide (although the solubilities of calcium bromide, calcium chloride, calcium fluoride, and calcium iodide all increase with temperature).
- 12.7 Calcium chloride is an example of a salt that releases heat when it dissolves (exothermic heat of solution). Ammonium nitrate is an example of a salt that absorbs heat when it dissolves (endothermic heat of solution).
- 12.8 As the temperature of the solution is increased by heating, the concentration of the dissolved gas would decrease.
- 12.9 A carbonated beverage must be stored in a closed container because carbonated beverages must contain more carbon dioxide than is soluble in water at atmospheric pressure. It is possible to add carbon dioxide under pressure to a closed container before it is sealed and increase the solubility of carbon dioxide. This is an illustration of Le Chatelier's principle, which states that the equilibrium between gaseous and dissolved carbon dioxide is shifted in favor of the dissolved carbon dioxide by an increase in pressure.
- 12.10 According to Le Chatelier's principle, a gas is more soluble in a liquid at higher pressures because, when the gas dissolves in the liquid, the system decreases in volume, tending to decrease the applied pressure. However, when a solid dissolves in a liquid, there is very little volume change. Thus, pressure has very little effect on the solubility of a solid in a liquid.
- 12.11 The four ways to express the concentration of a solute in a solution are (1) molarity, which is moles per liter; (2) mass percentage of solute, which is the percentage by mass of solute contained in a given mass of solution; (3) molality, which is the moles of solute per kilogram of solvent; and (4) mole fraction, which is the moles of the component substance divided by the total moles of solution.
- 12.12 The vapor pressure of the solvent of the dilute solution is larger than the vapor pressure of the solvent of the more concentrated solution (Raoult's law). Thus, the more dilute solution loses solvent and becomes more concentrated while the solvent molecules in the vapor state condense into the concentrated solution making it more dilute. After sufficient time has passed, the vapor pressure of the solvent in the closed container will reach a steady value (equilibrium), at which time the concentration of solute in the two solutions will be the same.
- 12.13 In fractional distillation, the vapor that first appears over a solution will have a greater mole fraction of the more volatile component. If a portion of this is vaporized and condensed, the liquid will be still richer in the more volatile component. After successive distillation stages, eventually the more volatile component will be obtained in pure form (Figure 12.19).

- 12.14 The boiling point of the solution is higher because the nonvolatile solute lowers the vapor pressure of the solvent. Thus, the temperature must be increased to a value greater than the boiling point of the pure solvent to achieve a vapor pressure equal to atmospheric pressure.
- 12.15 One application is the use of ethylene glycol in automobile radiators as antifreeze; the glycol-water mixture usually has a freezing point well below the average low temperature during the winter. A second application is spreading sodium chloride on icy roads in the winter to melt the ice. The ice usually melts because, at equilibrium, a concentrated solution of NaCl usually freezes at a temperature below that of the roads.
- 12.16 If a pressure greater than the osmotic pressure of the ocean water is applied, the natural osmotic flow can be reversed. Then, the water solvent flows from the ocean water through a membrane to a more dilute solution or to pure water, leaving behind the salt and other ionic compounds from the ocean in a more concentrated solution.
- 12.17 Part of the light from the sun is scattered in the direction of an observer by fine particles in the clouds (Tyndall effect) rather than being completely absorbed by the clouds. The scattered light becomes visible against the darker background of dense clouds.
- 12.18 The examples are: fog is an aerosol, whipped cream is a foam, mayonnaise is an emulsion, solid silver chloride dispersed in water is a sol, and fruit jelly is a gel.
- 12.19 The polar -OH group on glycerol allows it to interact (hydrogen bond) with the polar water molecules, which means it is like water and, therefore, will dissolve in water. Benzene is a nonpolar molecule, which indicates that it is not “like” water and, therefore, will not dissolve in water.
- 12.20 Soap removes oil from a fabric by absorbing the oil into the hydrophobic centers of the soap micelles and off the surface of the fabric. Rinsing removes the micelles from contact with the fabric and leaves only water on the fabric, which can then be dried.

■ Answers to Conceptual Problems

- 12.21 The amount of oxygen dissolved in water decreases as the temperature increases. Thus, at the lower temperatures, there is enough oxygen dissolved in the water to support both bass and trout. But, as the temperature rises above 23 °C, there is not enough dissolved oxygen in the warm water to support the trout, who need more O₂ than bass.

12.22 The salt that would best accomplish the task would be the salt that lowers the freezing point of water the most. This, in turn, will be the salt with the largest i factor. Ideally, if each salt dissolved completely, KCl , MgSO_4 , and AgCl would have $i = 2$. Similarly, CaCl_2 and PbS_2 would have $i = 3$. Of the latter two salts, CaCl_2 is more soluble than PbS_2 , so its i factor is closer to three. Therefore, the salt with the largest i factor is CaCl_2 , so it would lower the freezing point of water the most and would best accomplish the task.

12.23 The two main factors to consider when trying to determine the solubility of an ionic compound in water are ionic size and lattice energy. Ionic size is inversely related to the energy of hydration; the smaller the ion, the greater the energy of hydration. Keep in mind that the greater the energy of hydration, the more likely it is for a compound to dissolve. The amount of lattice energy is directly related to the solubility of the compound; the lower the lattice energy, the more likely it is for the compound to dissolve.

Taking into account these factors, in order to increase the solubility of a compound you need to decrease the ionic size and decrease the lattice energy. Since the energy of hydration of the Y^+ ion is greater than that of the X^+ ion (making XZ less soluble), in order for XZ to be more soluble than YZ , the lattice energy must be less for the XZ compound.

12.24 a. According to Raoult's law, the addition of a nonvolatile nonelectrolyte to a solvent will lower the vapor pressure of the solvent, so we would expect the vapor pressure of such a solution to be lower than that of the pure solvent (water in this case). When a volatile solute is added to a solvent, the vapor pressure of the solution is dependent upon the mole fraction of the solute and solvent and the vapor pressures of both the solute and solvent. Since the solute is volatile (a high vapor pressure relative to water), the solution must have a higher vapor pressure than pure water.

b. Keeping in mind that a solution will boil when the vapor pressure equals the pressure pushing on the surface of the solution, the solution with the greater vapor pressure will boil at a lower temperature. In this case, it is the solution with the volatile solute.

12.25 Smoke particles carry a small net charge that prevents them from forming larger particles that would settle to the bottom of the smokestack. The charged smoke particles are neutralized by the current, which then allows them to aggregate into large particles. These large particles are too big to be carried out of the stack.

12.26 a. Since beaker B contains more solute particles, according to Raoult's law, it will boil at a higher temperature than beaker A.

b. More particles in solution lead to lower vapor pressure, which in turn lowers the freezing point of a solution. Since beaker B contains more solute particles, it will freeze at a lower temperature than beaker A.

- c. When separated by a semipermeable membrane, the solvent from the less concentrated solution flows into the more concentrated solution. Because of this, the water will flow from beaker A to beaker B causing an increase in the concentration of NaCl in beaker A.
- 12.27 Vinegar is a solution of acetic acid (solute) and water (solvent). Because the salt concentration outside the lettuce leaf is higher than inside, water will pass out of the lettuce leaf into the dressing via osmosis. The result is that the lettuce will wilt.
- 12.28 As a solution freezes, pure solvent forms without any of the solute present. This means that, as ocean water freezes to make icebergs, it freezes as pure water without the salt present.
- 12.29 To answer these questions it is necessary to determine the relative concentrations of solute particles in each beaker. With equal moles and equal volumes, the solution concentration will be directly proportional to the van't Hoff factor, i , for that compound. Urea, a nonelectrolyte, has $i = 1$. NaCl, which is ionic, has $i = 2$, and CaCl_2 , which is also ionic, has $i = 3$.
- The highest boiling point will correspond to the largest boiling point elevation, ΔT_b . Since $\Delta T_b = iK_b m$, the largest ΔT_b will correspond to the highest value of i . Thus, the solution with CaCl_2 ($i = 3$) will have the highest boiling point.
 - The highest freezing point will correspond to the smallest freezing point depression, ΔT_f . Since $\Delta T_f = iK_f m$, the smallest ΔT_f will correspond to the smallest value of i . Thus, the solution with urea ($i = 1$) will have the highest freezing point.
- 12.30 To answer these questions, it is necessary to determine the relative concentrations of solute particles in each beaker. In order to compare volumes, use the lines on the sides of the beakers. Beaker A has a concentration of five atoms per two volume units, $5/2$ or $2.5/1$. Beaker B has a concentration of ten atoms per two volume units, $10/2$ or $5/1$. Beaker C has a concentration of five atoms per volume unit, $5/1$. Comparing the concentrations, beaker B = beaker C = 2 x beaker A.
- Since $P_A = X_A P_A^\circ$, the solution with the highest vapor pressure is the solution with the highest mole fraction of solvent, or conversely, the lowest concentration of solute particles. Thus, the solution in beaker A will have the highest vapor pressure.
 - The lowest boiling point will correspond to the smallest boiling point increase, ΔT_b . Since $\Delta T_b = K_b m$, the smallest ΔT_b will correspond to the smallest concentration of solute particles (molality here). Thus, the solution in beaker A will have the lowest boiling point.

- c. For each solution to have the same freezing point, they all need the same concentration of solute particles. This can be accomplished by doubling the volume in beakers B and C.

■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off. Atomic weights are rounded to two decimal places, except for that of hydrogen.

- 12.31 An example of a liquid solution prepared by dissolving a gas in a liquid is household ammonia, which consists of ammonia (NH_3) gas dissolved in water.
- 12.32 An example of a solid solution prepared from two solids is almost any alloy, such as 18-kt gold, which consists of 25 percent silver dissolved in 75 percent gold.
- 12.33 Boric acid would be more soluble in ethanol because this acid is polar and is more soluble in a more polar solvent. It can also hydrogen-bond to ethanol but not to benzene.
- 12.34 Naphthalene is more soluble in benzene because nonpolar naphthalene must break the strong hydrogen bonds between ethanol molecules and replace them with weaker London forces.
- 12.35 The order of increasing solubility is $\text{H}_2\text{O} < \text{CH}_2\text{OHCH}_2\text{OH} < \text{C}_{10}\text{H}_{22}$. The solubility in nonpolar hexane increases with the decreasing polarity of the solute.
- 12.36 Acetic acid is more soluble because it is polar like ethanol and can form hydrogen bonds to it.
- 12.37 The Al^{3+} ion has both a greater charge and a smaller ionic radius than Mg^{2+} , so Al^{3+} should have a greater energy of hydration.
- 12.38 The F^- ion has a smaller radius than the Cl^- ion, so F^- should have a greater hydration energy.

12.39 The order is $\text{Ba}(\text{IO}_3)_2 < \text{Sr}(\text{IO}_3)_2 < \text{Ca}(\text{IO}_3)_2 < \text{Mg}(\text{IO}_3)_2$. The iodate ion is fairly large, so the lattice energy for all these iodates should change to a smaller degree than the hydration energy of the cations. Therefore, solubility should increase with decreasing cation radius.

12.40 Fluorides (smaller anion): As the cation radius increases, the lattice energy of the fluoride salts decreases to a greater degree than the cation hydration energy decreases. Permanganates (larger anion): As the cation radius increases, the lattice energy of the permanganate salts decreases to a smaller degree than the cation hydration energy decreases. Therefore, the solubility of the fluorides increases with increasing cation radius, whereas the solubility of the permanganates decreases with increasing cation radius.

12.41 Using Henry's law, let S_1 = the solubility at 1.00 atm (P_1), and S_2 = the solubility at 5.50 atm (P_2).

$$S_2 = \frac{P_2 S_1}{P_1} = \frac{(5.50 \text{ atm})(0.161 \text{ g}/100 \text{ mL})}{1.00 \text{ atm}} = 0.8855$$

$$= 0.886 \text{ g}/100 \text{ mL}$$

12.42 The partial pressure of N_2 in air at 4.79 atm is

$$4.79 \text{ atm} \times 0.781 = 3.74099 \text{ atm}$$

Substituting this for P_2 in $S_2 = S_1 \times P_2/P_1$ (Henry's law), you obtain

$$S_2 = (0.00175 \text{ g}/100 \text{ mL}) \times \frac{3.74099 \text{ atm}}{1.00 \text{ atm}} = 0.006546$$

$$= 0.00655 \text{ g}/100 \text{ mL}$$

12.43 First, calculate the mass of KI in the solution; then calculate the mass of water needed.

$$\text{Mass KI} = 72.5 \text{ g} \times \frac{5.00 \text{ g KI}}{100 \text{ g solution}} = 3.6250 = 3.63 \text{ g KI}$$

$$\text{Mass H}_2\text{O} = 72.5 \text{ g soln} - 3.6250 \text{ g KI} = 68.875 = 68.9 \text{ g H}_2\text{O}$$

Dissolve 3.63 g KI in 68.9 g of water.

- 12.44 First, calculate the mass of Na_2SO_4 in the solution; then calculate the mass of water needed.

$$\text{Mass of Na}_2\text{SO}_4 = 455 \text{ g solution} \times \frac{6.50 \text{ g Na}_2\text{SO}_4}{100 \text{ g solution}} = 29.57 \text{ g Na}_2\text{SO}_4$$

$$\text{Mass of water} = 455 \text{ g soln} - 29.57 \text{ g Na}_2\text{SO}_4 = 425.43 \text{ g water}$$

Dissolve 29.57 g Na_2SO_4 in 425.43 g of water.

- 12.45 Multiply the mass of KI by 100 g of solution per 5.00 g KI (reciprocal of percentage).

$$0.258 \text{ g KI} \times \frac{100 \text{ g solution}}{5.00 \text{ g KI}} = 5.160 = 5.16 \text{ g solution}$$

- 12.46 Multiply the mass of sodium sulfate by 100 g of solution per 6.50 g Na_2SO_4 (reciprocal of percentage).

$$1.50 \text{ g Na}_2\text{SO}_4 \times \frac{100 \text{ g solution}}{6.50 \text{ g Na}_2\text{SO}_4} = 23.07 = 23.1 \text{ g solution}$$

- 12.47 Convert mass of vanillin ($\text{C}_8\text{H}_8\text{O}_3$, molar mass 152.14 g/mol) to moles, convert mg of ether to kg, and divide for molality.

$$0.0391 \text{ g vanillin} \times \frac{1 \text{ mol vanillin}}{152.14 \text{ g vanillin}} = 2.570 \times 10^{-4} \text{ mol vanillin}$$

$$168.5 \text{ mg ether} \times 1 \text{ kg}/10^6 \text{ mg} = 168.5 \times 10^{-6} \text{ kg ether}$$

$$\text{Molality} = \frac{2.570 \times 10^{-4} \text{ mol vanillin}}{168.5 \times 10^{-6} \text{ kg ether}} = 1.5252 = 1.53 \text{ m vanillin}$$

- 12.48 Convert mass of lauryl alcohol (LA, $\text{C}_{12}\text{H}_{25}\text{OH}$, molar mass 186.33 g/mol) to moles, convert g of ethanol to kg, and divide.

$$17.1 \text{ g LA} \times \frac{1 \text{ mol LA}}{186.3 \text{ g LA}} = 9.177 \times 10^{-2} \text{ mol LA}$$

$$165 \text{ g ethanol} \times 1 \text{ kg}/10^3 \text{ g} = 165 \times 10^{-3} \text{ kg ethanol}$$

$$\text{Molality} = \frac{9.177 \times 10^{-2} \text{ mol LA}}{165 \times 10^{-3} \text{ kg ethanol}} = 0.5561 = 0.556 \text{ m LA}$$

- 12.49 Convert mass of fructose ($\text{C}_6\text{H}_{12}\text{O}_6$, molar mass 180.16 g/mol) to moles, and then multiply by one kg H_2O per 0.125 mol fructose (the reciprocal of molality).

$$1.75 \text{ g fructose} \times \frac{1 \text{ mol fructose}}{180.16 \text{ g fructose}} \times \frac{1 \text{ kg H}_2\text{O}}{0.125 \text{ mol fructose}} \\ = 0.07770 \text{ kg (77.7 g H}_2\text{O)}$$

- 12.50 Multiply the kg of CHCl_3 by the molality to get moles of caffeine; then convert moles to mass.

$$0.0450 \text{ kg CHCl}_3 \times \frac{0.0946 \text{ mol caffeine}}{1 \text{ kg CHCl}_3} \times \frac{194.2 \text{ g caffeine}}{1 \text{ mol caffeine}} \\ = 0.8267 = 0.827 \text{ g caffeine}$$

- 12.51 Convert masses to moles, and then calculate the mole fractions.

$$65.0 \text{ g alc.} \times \frac{1 \text{ mol alc.}}{60.09 \text{ g alc.}} = 1.0817 \text{ mol alc.}$$

$$35.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 1.9422 = 1.94 \text{ mol H}_2\text{O}$$

$$\text{Mol fraction alc.} = \frac{\text{mol alc.}}{\text{total mol}} = \frac{1.0817 \text{ mol}}{3.0239 \text{ mol}} = 0.3577 = 0.358$$

$$\text{Mol fraction H}_2\text{O} = \frac{\text{mol H}_2\text{O}}{\text{total mol}} = \frac{1.9422 \text{ mol}}{3.0239 \text{ mol}} = 0.6422 = 0.642$$

- 12.52 Convert masses to moles, and then calculate the mole fractions.

$$2.50 \times 10^3 \text{ g glycol} \times \frac{1 \text{ mol glycol}}{62.07 \text{ g glycol}} = 40.277 \text{ mol glycol}$$

$$2.00 \times 10^3 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 110.98 \text{ mol H}_2\text{O}$$

$$\text{Total moles} = 40.277 + 110.98 = 151.26 \text{ mol}$$

$$\text{Mol fraction H}_2\text{O} = \frac{\text{mol H}_2\text{O}}{\text{total mol}} = \frac{110.98 \text{ mol}}{151.26 \text{ mol}} = 0.7336 = 0.734$$

$$\text{Mol fraction glycerol} = \frac{\text{mol glycerol}}{\text{total mol}} = \frac{40.277 \text{ mol}}{151.26 \text{ mol}} = 0.2662 = 0.266$$

12.53 In the solution, for every 0.650 mol of NaClO there is 1.00 kg, or 1.00×10^3 g, H_2O , so

$$1.00 \times 10^3 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 55.49 \text{ mol H}_2\text{O}$$

$$\text{Total mol} = 55.49 \text{ mol H}_2\text{O} + 0.650 \text{ mol NaClO} = 56.14 \text{ mol}$$

$$\text{Mol fraction NaClO} = \frac{\text{mol NaClO}}{\text{total mol}} = \frac{0.650 \text{ mol}}{56.14 \text{ mol}} = 0.01157 = 0.0116$$

12.54 In the solution, for every 0.600 mol of H_2O_2 , there is 1.00 kg, or 1.00×10^3 g, of H_2O , so

$$1.00 \times 10^3 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 55.49 \text{ mol H}_2\text{O}$$

$$\text{Total mol} = 55.49 \text{ mol H}_2\text{O} + 0.600 \text{ mol H}_2\text{O}_2 = 56.093 \text{ mol}$$

$$\text{Mol fraction H}_2\text{O}_2 = \frac{\text{mol H}_2\text{O}_2}{\text{mol solution}} = \frac{0.600 \text{ mol}}{56.093 \text{ mol}} = 0.01069 = 0.0107$$

12.55 The total moles of solution = $3.31 \text{ mol H}_2\text{O} + 1.00 \text{ mol HCl} = 4.31 \text{ mol}$.

$$\text{Mol fraction HCl} = \frac{1 \text{ mol HCl}}{4.31 \text{ mol}} = 0.23201 = 0.232$$

$$3.31 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times \frac{1 \text{ kg H}_2\text{O}}{10^3 \text{ g H}_2\text{O}} = 5.965 \times 10^{-2} \text{ kg H}_2\text{O}$$

$$\text{Molality} = \frac{1.00 \text{ mol HCl}}{5.965 \times 10^{-2} \text{ kg H}_2\text{O}} = 16.76 = 16.8 \text{ m}$$

12.56 The total moles of solution = $1.00 \text{ mol NH}_3 + 2.44 \text{ mol H}_2\text{O} = 3.44 \text{ mol}$.

$$\text{Mol fraction NH}_3 = \frac{1 \text{ mol NH}_3}{3.44 \text{ mol}} = 0.2906 = 0.291$$

$$2.44 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times \frac{1 \text{ kg H}_2\text{O}}{10^3 \text{ g H}_2\text{O}} = 4.397 \times 10^{-2} \text{ kg H}_2\text{O}$$

$$\text{Molality} = \frac{1.00 \text{ mol NH}_3}{4.397 \times 10^{-2} \text{ kg H}_2\text{O}} = 22.74 = 22.7 \text{ m}$$

- 12.57 The mass of 1.000 L of solution is 1.022 kg. In the solution, there are 0.580 moles of $\text{H}_2\text{C}_2\text{O}_4$ (OA) for every 1.0000 kg of water. Convert this number of moles to mass.

$$0.580 \text{ mol OA} \times \frac{90.04 \text{ g OA}}{1 \text{ mol OA}} \times \frac{1 \text{ kg OA}}{10^3 \text{ g OA}} = 0.05222 \text{ kg OA}$$

The total mass of the solution containing 1.000 kg H_2O and 0.580 moles of OA is calculated as follows:

$$\text{Mass} = 1.0000 \text{ kg H}_2\text{O} + 0.05222 \text{ kg OA} = 1.05222 \text{ kg}$$

Use this to relate the mass of 1.000 L (1.022 kg) of solution to the amount of solute:

$$1.022 \text{ kg solution} \times \frac{0.580 \text{ mol OA}}{1.05222 \text{ kg solution}} = 0.5633 \text{ mol OA}$$

$$\text{Molarity} = \frac{0.5633 \text{ mol OA}}{1.00 \text{ L solution}} = 0.5633 = 0.563 \text{ M}$$

- 12.58 In the solution, there is 0.688 mole of citric acid (CA) for every 1.000 kg of water. Convert this number of moles to mass.

$$0.688 \text{ mol CA} \times \frac{192.1 \text{ g CA}}{1 \text{ mol CA}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.1321 \text{ kg CA}$$

The total mass of solution containing 1.000 kg H_2O and 0.688 moles of CA is:

$$\text{Mass} = 1.0000 \text{ kg H}_2\text{O} + 0.1321 \text{ kg CA} = 1.1321 \text{ kg} = 1132.1 \text{ g}$$

Now, use the density of the solution (1.049 g/mL) to convert this mass to liters of solution.

$$\text{Volume} = 1132.1 \text{ g} \times \frac{1 \text{ mL}}{1.049 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 1.0792 \text{ L}$$

Since there are 0.688 moles of CA in this volume, you can now calculate the molarity of the solution.

$$\text{Molarity} = \frac{\text{mol CA}}{\text{Volume}} = \frac{0.688 \text{ mol}}{1.0792 \text{ L}} = 0.63746 = 0.637 \text{ M}$$

- 12.59 In 1.000 L of vinegar, there is 0.763 mole of acetic acid. The total mass of the 1.000 L solution is 1.004 kg. Start by calculating the mass of acetic acid (AA) in the solution.

$$0.763 \text{ mol AA} \times \frac{60.05 \text{ g AA}}{1 \text{ mol AA}} = 45.\underline{82} \text{ g AA} \text{ (0.045}\underline{82} \text{ kg AA)}$$

The mass of water may be found by difference:

$$\text{Mass H}_2\text{O} = 1.004 \text{ kg soln} - 0.04582 \text{ kg AA} = 0.9582 \text{ kg H}_2\text{O}$$

$$\text{Molality} = \frac{0.763 \text{ mol AA}}{0.9582 \text{ kg H}_2\text{O}} = 0.7962 = 0.796 \text{ m AA}$$

- 12.60 In 1.000 L of beverage, there is 0.265 mole of tartaric acid. The total mass of the 1.000 L beverage is 1.016 kg. Start by calculating the mass of tartaric acid (TA) in the beverage.

$$0.265 \text{ mol TA} \times \frac{150.1 \text{ g TA}}{1 \text{ mol TA}} = 39.\underline{77} \text{ g TA} \text{ (0.039}\underline{77} \text{ kg TA)}$$

The mass of water may be found by difference:

$$\text{Mass H}_2\text{O} = 1.016 \text{ kg soln} - 0.03977 \text{ kg TA} = 0.9762 \text{ kg H}_2\text{O}$$

$$\text{Molality} = \frac{0.265 \text{ mol TA}}{0.9762 \text{ kg H}_2\text{O}} = 0.2714 = 0.271 \text{ m TA}$$

- 12.61 To find the mole fraction of sucrose, first find the amounts of both sucrose (suc.) and water:

$$20.2 \text{ g sucrose} \times \frac{1 \text{ mol sucrose}}{342.30 \text{ g sucrose}} = 0.05901 \text{ mol sucrose}$$

$$70.1 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 3.890 \text{ mol H}_2\text{O}$$

$$X_{\text{sucrose}} = \frac{0.05901 \text{ mol sucrose}}{(3.890 + 0.05901) \text{ mol}} = 0.01494$$

From Raoult's law, the vapor pressure (P) and lowering (ΔP) are

$$\begin{aligned} P &= P_{\text{H}_2\text{O}}^{\circ} X_{\text{H}_2\text{O}} = P_{\text{H}_2\text{O}}^{\circ} (1 - X_{\text{suc.}}) = (42.2 \text{ mmHg})(1 - 0.01494) \\ &= 41.\underline{569} = 41.6 \text{ mmHg} \end{aligned}$$

$$\Delta P = P_{\text{H}_2\text{O}}^{\circ} X_{\text{suc.}} = (42.2 \text{ mmHg})(0.01494) = 0.6306 = 0.631 \text{ mmHg}$$

- 12.62 To find the mole fraction of benzene, first find the amounts of both benzene (ben.) and naphthalene (nap.):

$$1.20 \text{ g nap.} \times \frac{1 \text{ mol nap.}}{128.2 \text{ g nap.}} = 0.0093603 \text{ mol nap.}$$

$$25.6 \text{ g ben.} \times \frac{1 \text{ mol ben.}}{78.11 \text{ g ben.}} = 0.3277 \text{ mol ben.}$$

$$X_{\text{nap.}} = \frac{0.0093603 \text{ mol nap.}}{(0.3277 + 0.0093603) \text{ mol}} = 0.02777$$

From Raoult's law, the vapor pressure (P) and lowering (ΔP) are

$$\begin{aligned} P &= P_{\text{ben.}}^{\circ} (1 - X_{\text{nap.}}) = (86.0 \text{ mmHg})(1 - 0.02777) = 83.61 \\ &= 83.6 \text{ mmHg} \end{aligned}$$

$$\Delta P = P_{\text{ben.}}^{\circ} X_{\text{nap.}} = (86.0 \text{ mmHg})(0.02777) = 2.388 = 2.39 \text{ mmHg}$$

- 12.63 Find the molality of glycerol (gly.) in the solution first:

$$0.150 \text{ g gly.} \times \frac{1 \text{ mol gly.}}{92.095 \text{ g gly.}} = 0.001628 \text{ mol gly.}$$

$$\text{Molality} = \frac{0.001628 \text{ mol gly.}}{0.0200 \text{ kg solvent}} = 0.081437 \text{ m}$$

Substitute $K_b = 0.512 \text{ }^{\circ}\text{C/m}$ and $K_f = 1.858 \text{ }^{\circ}\text{C/m}$ (Table 12.3) into equations for ΔT_b and ΔT_f :

$$\Delta T_b = K_b m = 0.512 \text{ }^{\circ}\text{C/m} \times 0.081437 \text{ m} = 0.04169 \text{ }^{\circ}\text{C}$$

$$T_b = 100.000 + 0.04169 = 100.04169 = 100.042 \text{ }^{\circ}\text{C}$$

$$\Delta T_f = K_f m = 1.858 \text{ }^{\circ}\text{C/m} \times 0.081437 \text{ m} = 0.15131 \text{ }^{\circ}\text{C}$$

$$T_f = 0.000 - 0.15131 = -0.15131 = -0.151 \text{ }^{\circ}\text{C}$$

12.64 Find the molality of the solution first:

$$0.817 \text{ g S}_8 \times \frac{1 \text{ mol S}_8}{256.56 \text{ g S}_8} = 0.003184 \text{ mol S}_8$$

$$\text{Molality} = \frac{0.003184 \text{ mol S}_8}{0.1000 \text{ kg solvent}} = 0.03184 \text{ m}$$

Substitute $K_b = 3.08 \text{ }^\circ\text{C/m}$ and $K_f = 3.59 \text{ }^\circ\text{C/m}$ (Table 12.3) into equations for ΔT_b and ΔT_f :

$$\Delta T_b = K_b m = \frac{3.08 \text{ }^\circ\text{C}}{\text{m S}_8} \times 0.03184 \text{ m S}_8 = 0.09808 \text{ }^\circ\text{C}$$

$$T_b = 118.5 \text{ }^\circ\text{C} + 0.09808 \text{ }^\circ\text{C} = 118.598 = 118.6 \text{ }^\circ\text{C}$$

$$\Delta T_f = K_f m = \frac{3.59 \text{ }^\circ\text{C}}{\text{m S}_8} \times 0.03184 \text{ m S}_8 = 0.1143 \text{ }^\circ\text{C}$$

$$T_f = 16.60 \text{ }^\circ\text{C} - 0.1143 \text{ }^\circ\text{C} = 16.4856 = 16.49 \text{ }^\circ\text{C}$$

12.65 Calculate ΔT_f , the freezing-point depression, and, using $K_f = 1.858 \text{ }^\circ\text{C/m}$ (Table 12.3), the molality, c_m .

$$\Delta T_f = 0.000 \text{ }^\circ\text{C} - (-0.086 \text{ }^\circ\text{C}) = 0.086 \text{ }^\circ\text{C}$$

$$c_m = \frac{\Delta T_f}{K_f} = \frac{0.086 \text{ }^\circ\text{C}}{1.858 \text{ }^\circ\text{C/m}} = 4.628 \times 10^{-2} = 4.6 \times 10^{-2} \text{ m}$$

12.66 Calculate ΔT_f , the freezing-point depression, and, using $K_f = 1.858 \text{ }^\circ\text{C/m}$ (Table 12.3), the molality, c_m .

$$\Delta T_f = 0.000 \text{ }^\circ\text{C} - (-0.085 \text{ }^\circ\text{C}) = 0.085 \text{ }^\circ\text{C}$$

$$c_m = \frac{\Delta T_f}{K_f} = \frac{0.085 \text{ }^\circ\text{C}}{1.858 \text{ }^\circ\text{C/m}} = 0.0457 \text{ m}$$

(continued)

Now, find the amount of urea in the solution from the definition of molality:

$$\text{Mol}_{\text{urea}} = c_m \times (\text{kg H}_2\text{O}) = \frac{0.0457 \text{ mol urea}}{1 \text{ kg H}_2\text{O}} \times 0.1000 \text{ kg H}_2\text{O}$$

$$= 0.00457 \text{ mol}$$

$$0.00457 \text{ mol urea} \times \frac{60.06 \text{ g urea}}{1 \text{ mol urea}} = 0.274 = 0.27 \text{ g urea}$$

12.67 Find the moles of unknown solute from the definition of molality:

$$\text{Mol}_{\text{solute}} = m \times \text{kg solvent} = \frac{0.0698 \text{ mol}}{1 \text{ kg solvent}} \times 0.002135 \text{ kg solvent}$$

$$= 1.490 \times 10^{-4} \text{ mol}$$

$$\text{Molar mass} = \frac{0.0182 \text{ g}}{1.490 \times 10^{-4} \text{ mol}} = 122.1 = 122 \text{ g/mol}$$

The molecular weight is 122 amu.

12.68 Find the moles of the solute from the definition of molality:

$$\text{Mol}_{\text{solute}} = m \times \text{kg solvent} = \frac{0.0368 \text{ mol}}{1 \text{ kg solvent}} \times 0.00831 \text{ kg solvent}$$

$$= 3.058 \times 10^{-4} \text{ mol}$$

$$\text{Molar mass} = \frac{0.0653 \text{ g}}{3.058 \times 10^{-4} \text{ mol}} = 213.5 = 214 \text{ g/mol}$$

The molecular weight is 214 amu.

- 12.69 Calculate ΔT_f , the freezing-point depression, and, using $K_f = 1.858\text{ }^\circ\text{C/m}$ (Table 12.3), the molality, c_m .

$$\Delta T_f = 26.84\text{ }^\circ\text{C} - 25.70\text{ }^\circ\text{C} = 1.14\text{ }^\circ\text{C}$$

$$c_m = \frac{\Delta T_f}{K_f} = \frac{1.14\text{ }^\circ\text{C}}{8.00\text{ }^\circ\text{C/m}} = 0.1425\text{ m}$$

Find the moles of solute by rearranging the definition of molality:

$$\begin{aligned}\text{Mol} &= m \times \text{kg solvent} = \frac{0.1425\text{ mol}}{1\text{ kg solvent}} \times 103 \times 10^{-6}\text{ kg solvent} \\ &= 1.467 \times 10^{-5}\text{ mol}\end{aligned}$$

$$\text{Molar mass} = \frac{2.39 \times 10^{-3}\text{ g}}{1.467 \times 10^{-5}\text{ mol}} = 162.9 = 163\text{ g/mol}$$

The molecular weight is 163 amu.

- 12.70 Calculate ΔT_f , the freezing-point depression, and, using $K_f = 5.065\text{ }^\circ\text{C/m}$ (Table 12.3), the molality, c_m .

$$\Delta T_f = 5.455\text{ }^\circ\text{C} - 4.880\text{ }^\circ\text{C} = 0.575\text{ }^\circ\text{C}$$

$$c_m = \frac{\Delta T_f}{K_f} = \frac{0.575\text{ }^\circ\text{C}}{5.065\text{ }^\circ\text{C/m}} = 0.1135\text{ m}$$

Find the moles of solute by rearranging the definition of molality:

$$\text{Mol} = m \times \text{kg solvent} = \frac{0.1135\text{ mol}}{1\text{ kg solvent}} \times 0.1000\text{ kg solvent} = 0.01135\text{ mol}$$

$$\text{Molar mass} = \frac{2.500\text{ g}}{0.01135\text{ mol}} = 220.2 = 2.20 \times 10^2\text{ g/mol}$$

The molecular weight is 2.20×10^2 amu.

12.71 Use the equation for osmotic pressure (π) to solve for the molarity of the solution.

$$M = \frac{\pi}{RT} = \frac{1.47 \text{ mmHg} \times \left(\frac{1 \text{ atm}}{760 \text{ mmHg}} \right)}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(21 + 273)\text{K}} = 8.013 \times 10^{-5} \text{ mol/L}$$

Now, find the number of moles in 106 mL (0.106 L) using the molarity.

$$0.106 \text{ L} \times \frac{8.013 \times 10^{-5} \text{ mol}}{1 \text{ L}} = 8.494 \times 10^{-6} \text{ mol}$$

$$\text{Molar mass} = \frac{0.582 \text{ g}}{8.494 \times 10^{-6} \text{ mol}} = 6.851 \times 10^4 = 6.85 \times 10^4 \text{ g/mol}$$

The molecular weight is 6.85×10^4 amu.

12.72 Use the equation for osmotic pressure (π) to solve for the molarity of the solution.

$$M = \frac{\pi}{RT} = \frac{3.70 \text{ mmHg} \times \left(\frac{1 \text{ atm}}{760 \text{ mmHg}} \right)}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(25 + 273)\text{K}} = 1.989 \times 10^{-4} \text{ mol/L}$$

Now, find the number of moles in 100.0 mL (0.1000 L) using the molarity.

$$0.100 \text{ L} \times \frac{1.989 \times 10^{-4} \text{ mol}}{1 \text{ L}} = 1.989 \times 10^{-5} \text{ mol}$$

$$\text{Molar mass} = \frac{0.0216 \text{ g}}{1.989 \times 10^{-5} \text{ mol}} = 1.0859 \times 10^3 = 1.09 \times 10^3 \text{ g/mol}$$

The molecular weight is 1.09×10^3 amu.

12.73 Begin by noting that $i = 3$. Then, calculate ΔT_f from the product of $iK_f c_m$:

$$3 \times \frac{1.858^\circ\text{C}}{\text{m}} \times 0.0075 \text{ m} = 0.0418^\circ\text{C}$$

The freezing point = $0.000^\circ\text{C} - 0.0418^\circ\text{C} = -0.0418 = -0.042^\circ\text{C}$.

- 12.74 Start by noting that $i = 4$. Then, calculate ΔT_f from the product of $iK_f c_m$:

$$4 \times \frac{1.858\text{ }^\circ\text{C}}{\text{m}} \times 0.0091\text{ m} = 0.0676\text{ }^\circ\text{C}$$

$$\text{The freezing point} = 0.000\text{ }^\circ\text{C} - 0.0676\text{ }^\circ\text{C} = -0.0676 = -0.068\text{ }^\circ\text{C}.$$

- 12.75 Begin by calculating the molarity of $\text{Cr}(\text{NH}_3)_5\text{Cl}_3$.

$$\begin{aligned} 1.40 \times 10^{-2}\text{ g Cr}(\text{NH}_3)_5\text{Cl}_3 &\times \frac{1\text{ mol Cr}(\text{NH}_3)_5\text{Cl}_3}{243.5\text{ g Cr}(\text{NH}_3)_5\text{Cl}_3} \\ &= 5.749 \times 10^{-5}\text{ mol Cr}(\text{NH}_3)_5\text{Cl}_3 \end{aligned}$$

$$\text{Molarity} = \frac{5.749 \times 10^{-5}\text{ mol Cr}(\text{NH}_3)_5\text{Cl}_3}{0.0250\text{ L}} = 0.002299\text{ M}$$

Now find the hypothetical osmotic pressure, assuming $\text{Cr}(\text{NH}_3)_5\text{Cl}_3$ does not ionize:

$$\begin{aligned} \pi = MRT &= (2.30 \times 10^{-3}\text{ M}) \times \frac{0.0821\text{ L}\cdot\text{atm}}{\text{K}\cdot\text{mol}} \times 298\text{ K} \times \frac{760\text{ mmHg}}{1\text{ atm}} \\ &= 42.77\text{ mmHg} \end{aligned}$$

The measured osmotic pressure is greater than the hypothetical osmotic pressure. The number of ions formed per formula unit equals the ratio of the measured pressure to the hypothetical pressure:

$$i = \frac{119\text{ mmHg}}{42.77\text{ mmHg}} = 2.782 \approx 3\text{ ions/formula unit}$$

- 12.76 Begin by calculating the molality of the solution after rearranging $\Delta T_b = K_b c_m$:

$$c_m = \frac{\Delta T_b}{K_b} = \frac{5\text{ }^\circ\text{C}}{0.512\text{ }^\circ\text{C/m}} = 9.76\text{ m}$$

The 9.76 m is the molality of the ions, not of the NaCl. The molality of the NaCl will be $9.76\text{ m} \div 2 = 4.88\text{ m}$ because $i = 2$. Thus, 4.88 moles of NaCl must be dissolved in the one kg of water. This is converted to mass using the molar mass of NaCl:

$$4.88\text{ mol NaCl} \times \frac{58.5\text{ g NaCl}}{1\text{ mol NaCl}} = 285 = 3 \times 10^2\text{ g NaCl}$$

- 12.77 a. Aerosol (liquid water in air)
 b. Sol [solid $\text{Mg}(\text{OH})_2$ in liquid water]
 c. Foam (air in liquid soap solution)
 d. Sol (solid silt in liquid water)
- 12.78 a. Aerosol (liquid water in air)
 b. Foam (air in liquid egg white)
 c. Aerosol (solid dust particles in air)
 d. Emulsion (liquid oil in liquid vinegar)
- 12.79 Because the As_2S_3 particles are negatively charged, the effective coagulation requires a highly charged cation, so $\text{Al}_2(\text{SO}_4)_3$ is the best choice.
- 12.80 The K_3PO_4 contains highly charged PO_4^{3-} ions that will be strongly attracted to the positively charged solution and will cause it to coagulate.

■ Solutions to General Problems

- 12.81 Using Henry's law [$S_2 = S_1 \times (P_2/P_1)$, where $P_1 = 1.00 \text{ atm}$], find the solubility of each gas at P_2 , its partial pressure. For N_2 , $P_2 = 0.800 \times 1.00 \text{ atm} = 0.800 \text{ atm}$; for O_2 , $P_2 = 0.200 \text{ atm}$.

$$\text{N}_2: S_2 \times \frac{P_2}{P_1} = (0.0175 \text{ g/L H}_2\text{O}) \times \frac{0.800 \text{ atm}}{1.00 \text{ atm}} = 0.0140 \text{ g/L H}_2\text{O}$$

$$\text{O}_2: S_2 \times \frac{P_2}{P_1} = (0.0393 \text{ g/L H}_2\text{O}) \times \frac{0.200 \text{ atm}}{1.00 \text{ atm}} = 0.00786 \text{ g/L H}_2\text{O}$$

In 1.00 L of the water, there are 0.0140 g of N_2 and 0.00786 g of O_2 . If the water is heated to drive off both dissolved gases, the gas mixture that is expelled will contain 0.0140 g of N_2 and 0.00786 g of O_2 .

(continued)

Convert both masses to moles using the molar masses:

$$0.0140 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.01 \text{ g N}_2} = 4.998 \times 10^{-4} \text{ mol}$$

$$0.00786 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 2.456 \times 10^{-4} \text{ mol}$$

Now calculate the mole fractions of each gas:

$$X_{\text{N}_2} = \frac{\text{mol N}_2}{\text{total mol}} = \frac{4.998 \times 10^{-4} \text{ mol}}{(4.998 + 2.456) \times 10^{-4} \text{ mol}} = 0.67051 = 0.671$$

$$X_{\text{O}_2} = 1 - X_{\text{N}_2} = 1 - 0.67051 = 0.32949 = 0.329$$

- 12.82 Using Henry's law [$S_2 = S_1 \times (P_2/P_1)$, where $P_1 = 1.00 \text{ atm}$], find the solubility of each gas at P_2 , its partial pressure. For CH_4 , $P_2 = 0.900 \times 1.00 \text{ atm} = 0.900 \text{ atm}$; for C_2H_6 , $P_2 = 0.100 \times 1.00 \text{ atm} = 0.100 \text{ atm}$.

$$\text{CH}_4 : S_2 \times \frac{P_2}{P_1} = (0.023 \text{ g/L H}_2\text{O}) \times \frac{0.900 \text{ atm}}{1.00 \text{ atm}} = 0.0207 \text{ g/L H}_2\text{O}$$

$$\text{C}_2\text{H}_6 : S_2 \times \frac{P_2}{P_1} = (0.059 \text{ g/L H}_2\text{O}) \times \frac{0.100 \text{ atm}}{1.00 \text{ atm}} = 0.00590 \text{ g/L H}_2\text{O}$$

When 1.00 L of the water saturated with the gas mixture is heated, 0.0207 g of CH_4 and 0.0059 g of C_2H_6 will be expelled. Convert these to moles using molar masses:

$$0.0207 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} = 1.291 \times 10^{-3} \text{ mol}$$

$$0.0059 \text{ g C}_2\text{H}_6 \times \frac{1 \text{ mol C}_2\text{H}_6}{30.07 \text{ g C}_2\text{H}_6} = 1.962 \times 10^{-4} \text{ mol}$$

Now, find the mole fractions of each gas:

$$X_{\text{CH}_4} = \frac{\text{mol CH}_4}{\text{total mol}} = \frac{1.291 \times 10^{-3} \text{ mol}}{(1.291 + 0.196) \times 10^{-3} \text{ mol}} = 0.86819 = 0.87$$

$$X_{\text{C}_2\text{H}_6} = 1 - X_{\text{CH}_4} = 1 - 0.86819 = 0.13181 = 0.13$$

- 12.83 Assume a volume of 1.000 L whose mass is then 1.024 kg. Use the percent composition given to find the mass of each of the components of the solution.

$$1.024 \text{ kg soln} \times \frac{8.50 \text{ kg NH}_4\text{Cl}}{100.00 \text{ kg soln}} = 0.08704 \text{ kg NH}_4\text{Cl}$$

$$\text{Mass of H}_2\text{O} = 1.024 \text{ kg soln} - 0.08704 \text{ kg NH}_4\text{Cl} = 0.9370 \text{ kg H}_2\text{O}$$

Convert mass of NH_4Cl and water to moles:

$$87.04 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} = 1.627 \text{ mol NH}_4\text{Cl}$$

$$937.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.015 \text{ g H}_2\text{O}} = 52.01 \text{ mol H}_2\text{O}$$

$$\text{Molarity} = \frac{\text{mol NH}_4\text{Cl}}{\text{L solution}} = \frac{1.627 \text{ mol}}{1.00 \text{ L}} = 1.627 = 1.63 \text{ M}$$

$$\text{Molality} = \frac{\text{mol NH}_4\text{Cl}}{\text{kg H}_2\text{O}} = \frac{1.627 \text{ mol}}{0.9370 \text{ kg H}_2\text{O}} = 1.736 = 1.74 \text{ m}$$

$$X_{\text{NH}_4\text{Cl}} = \frac{\text{mol NH}_4\text{Cl}}{\text{total moles}} = \frac{1.627 \text{ mol}}{(52.01 + 1.627) \text{ mol}} = 0.03033 = 0.0303$$

- 12.84 Assume a volume of 1.000 L whose mass is then 1.127 kg. Use the percent composition given to find the mass of each of the components of the solution.

$$1.127 \text{ kg soln} \times \frac{27.0 \text{ kg LiCl}}{100.00 \text{ kg soln}} = 0.3042 \text{ kg} = 304.2 \text{ g LiCl}$$

$$\text{Mass of H}_2\text{O} = 1.127 \text{ kg soln} - 0.3042 \text{ kg LiCl} = 0.8227 \text{ kg} = 822.7 \text{ g H}_2\text{O}$$

Convert mass of LiCl and water to moles:

$$304.2 \text{ g LiCl} \times \frac{1 \text{ mol LiCl}}{42.39 \text{ g LiCl}} = 7.178 \text{ mol LiCl}$$

$$822.7 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.015 \text{ g H}_2\text{O}} = 45.66 \text{ mol H}_2\text{O}$$

$$\text{Total moles} = 7.178 + 45.66 = 52.838 \text{ mol}$$

(continued)

The various quantities can now be calculated.

$$\text{Molarity} = \frac{\text{mol LiCl}}{\text{L soln}} = \frac{7.178 \text{ mol LiCl}}{1.000 \text{ L soln}} = 7.178 = 7.18 \text{ M}$$

$$\text{Molality} = \frac{\text{mol LiCl}}{\text{kg H}_2\text{O}} = \frac{7.178 \text{ mol LiCl}}{0.8227 \text{ kg H}_2\text{O}} = 8.7249 = 8.72 \text{ m}$$

$$X_{\text{LiCl}} = \frac{\text{mol LiCl}}{\text{total moles}} = \frac{7.178 \text{ mol LiCl}}{52.838 \text{ mol}} = 0.1358 = 0.136$$

- 12.85 In 1.00 mol of gas mixture, there are 0.51 mol of propane (pro.) and 0.49 mol of butane (but.). First, calculate the masses of these components.

$$0.51 \text{ mol pro.} \times \frac{44.10 \text{ g pro.}}{1 \text{ mol pro.}} = 22.491 \text{ g pro.}$$

$$0.49 \text{ mol but.} \times \frac{58.12 \text{ g but.}}{1 \text{ mol but.}} = 28.478 \text{ g but.}$$

The mass of 1.00 mol of gas mixture is the sum of the masses of the two components:

$$22.491 \text{ g pro.} + 28.478 \text{ g but.} = 50.969 \text{ g mixture}$$

Therefore, in 50.969 g of the mixture, there are 22.491 g of propane and 28.478 g of butane. For a sample with a mass of 55 g:

$$55 \text{ g mixture} \times \frac{22.491 \text{ g pro.}}{50.969 \text{ g mixture}} = 24.26 = 24 \text{ g pro.}$$

$$55 \text{ g mixture} \times \frac{28.478 \text{ g but.}}{50.969 \text{ g mixture}} = 30.73 = 31 \text{ g but.}$$

- 12.86 In 1.000 mol of gas mixture, there are 0.036 mol O₂, 0.056 mol N₂, and 0.908 mol He. Calculate the masses of these components first.

$$0.036 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 1.\underline{15} \text{ g O}_2$$

$$0.056 \text{ mol N}_2 \times \frac{28.01 \text{ g N}_2}{1 \text{ mol N}_2} = 1.\underline{57} \text{ g N}_2$$

$$0.908 \text{ mol He} \times \frac{4.0026 \text{ g He}}{1 \text{ mol He}} = 3.\underline{634} \text{ g He}$$

The mass of 1.000 mol of the gas mixture is the sum of the masses of the components:

$$1.15 \text{ g O}_2 + 1.57 \text{ g N}_2 + 3.634 \text{ g He} = 6.35 \text{ g mixture}$$

For a sample with a mass of 7.84 g,

$$7.84 \text{ g mixture} \times \frac{1.15 \text{ g O}_2}{6.35 \text{ g mixture}} = 1.\underline{41} = 1.4 \text{ g O}_2$$

$$7.84 \text{ g mixture} \times \frac{1.57 \text{ g N}_2}{6.35 \text{ g mixture}} = 1.\underline{93} = 1.9 \text{ g N}_2$$

$$7.84 \text{ g mixture} \times \frac{3.634 \text{ g He}}{6.35 \text{ g mixture}} = 4.\underline{48} = 4.5 \text{ g He}$$

$$12.87 \quad P_{\text{ED}} = P^{\circ}_{\text{ED}}(X_{\text{ED}}) = 173 \text{ mmHg} (0.25) = 4\underline{3}.25 \text{ mmHg}$$

$$P_{\text{PD}} = P^{\circ}_{\text{PD}}(E_{\text{PD}}) = 127 \text{ mmHg} (0.75) = 9\underline{5}.25 \text{ mmHg}$$

$$P = P_{\text{ED}} + P_{\text{PD}} = 4\underline{3}.25 + 9\underline{5}.25 = 138.\underline{50} = 139 \text{ mmHg}$$

$$12.88 \quad P_{\text{B}} = P^{\circ}_{\text{B}}(X_{\text{B}}) = 75 \text{ mmHg} (0.30) = 2\underline{2}.5 \text{ mmHg}$$

$$P_{\text{T}} = P^{\circ}_{\text{T}}(E_{\text{T}}) = 22 \text{ mmHg} (0.70) = 1\underline{5}.4 \text{ mmHg}$$

$$P = P_{\text{B}} + P_{\text{T}} = 22.5 + 15.4 = 3\underline{7}.9 = 38 \text{ mmHg}$$

12.89 Calculate the moles of $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ using its molar mass of 474.4 g/mol, and use this to calculate the three concentrations.

- a. The moles of $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ are calculated below, using the abbreviation of "Hyd" for the formula of $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$.

$$\text{mol Hyd.} = 0.1186 \text{ g Hyd.} \times \frac{1 \text{ mol Hyd}}{474.4 \text{ g Hyd}} = 0.0002500 = 0.0002500 \text{ mol}$$

Note that one mol of $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ contains one mole of $\text{KAl}(\text{SO}_4)_2$, so calculating the molarity of $\text{KAl}(\text{SO}_4)_2$ can be performed using the moles of $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$.

$$\frac{\text{mol}}{\text{L}} = \frac{0.0002500 \text{ mol Hyd}}{1.000 \text{ L soln.}} \times \frac{1 \text{ mol KAl}(\text{SO}_4)_2}{1 \text{ mol Hyd}} = 0.0002500 \text{ M KAl}(\text{SO}_4)_2$$

- b. The molarity of the SO_4^{2-} ion will be twice that of the $\text{KAl}(\text{SO}_4)_2$.

$$\frac{\text{mol SO}_4^{2-}}{\text{L}} = 0.0002500 \text{ M KAl}(\text{SO}_4)_2 \times \frac{2 \text{ mol SO}_4^{2-}}{1 \text{ mol KAl}(\text{SO}_4)_2} = 0.0005000 \text{ M}$$

- c. Since the density of the solution is 1.00 g/mL, the mass of 1.000 L of solution is 1000 g, or 1.000 kg. Since molality is moles per 1.000 kg of solvent, the molality of $\text{KAl}(\text{SO}_4)_2$ equals 0.0002500 moles divided by 1.000 kg or 0.0002500 m (the same as the molarity).

12.90 Calculate the moles of $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$ using its molar mass of 666.46 g/mol, and use this to calculate the three concentrations.

- a. The moles of $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$ are calculated below, using the abbreviation "Hyd" for the formula of $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$.

$$\text{mol Hyd} = 0.1593 \text{ g Hyd} \times \frac{1 \text{ mol Hyd}}{666.46 \text{ g Hyd}} = 2.3902 \times 10^{-4} \text{ mol}$$

Note that one mole of $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$ contains one mole of $\text{Al}_2(\text{SO}_4)_3$, so calculating the molarity of $\text{Al}_2(\text{SO}_4)_3$ can be performed using the moles of $\text{Al}_2(\text{SO}_4)_3 \cdot 18\text{H}_2\text{O}$.

$$\text{Molarity} = \frac{\text{mol Hyd}}{\text{L soln}} = \frac{2.3902 \times 10^{-4} \text{ mol}}{1.000 \text{ L soln}} = 2.3902 \times 10^{-4} = 2.390 \times 10^{-4} \text{ M}$$

(continued)

- b. The molarity of the SO_4^{2-} ion will be three times that of the $\text{Al}_2(\text{SO}_4)_3$:

$$\begin{aligned} 2.3902 \times 10^{-4} \text{ mol Al}_2(\text{SO}_4)_3 \times \frac{3 \text{ mol SO}_4^{2-}}{1 \text{ mol Al}_2(\text{SO}_4)_3} &= 7.1707 \times 10^{-4} \\ &= 7.171 \times 10^{-4} \text{ M} \end{aligned}$$

- c. Since the density of the solution is 1.00 g/mL, the mass of 1.000 L of the solution is 1000. g. Of this, 0.1593 g is solute, so the mass of water in the solution is

$$1000. - 0.1593 = 999.84 \text{ g} = 0.99984 \text{ kg H}_2\text{O}$$

The molality of $\text{Al}_2(\text{SO}_4)_3$ can now be calculated.

$$\begin{aligned} \text{Molality} &= \frac{\text{mol Al}_2(\text{SO}_4)_3}{\text{kg H}_2\text{O}} = \frac{2.3902 \times 10^{-4} \text{ mol}}{0.99984 \text{ kg}} = 2.3906 \times 10^{-4} \\ &= 2.391 \times 10^{-4} \text{ M} \end{aligned}$$

- 12.91 In 1.00 kg of a saturated solution of urea, there are 0.41 kg of urea (a molecular solute) and 0.59 kg of water. First, convert the mass of urea to moles.

$$0.41 \times 10^3 \text{ g urea} \times \frac{1 \text{ mol urea}}{60.06 \text{ g urea}} = 6.826 \text{ mol urea}$$

Then, find the molality of the urea in the solution:

$$\text{Molality} = \frac{\text{mol urea}}{\text{kg H}_2\text{O}} \times \frac{6.826 \text{ mol urea}}{0.59 \text{ kg}} = 11.57 \text{ m}$$

$$\Delta T_f = K_f c_m = (1.858 \text{ }^\circ\text{C/m})(11.57 \text{ m}) = 21.4 \text{ }^\circ\text{C}$$

$$T_f = 0.0 \text{ }^\circ\text{C} - 21.4 \text{ }^\circ\text{C} = -21.4 = -21 \text{ }^\circ\text{C}$$

- 12.92 In 1.00 kg of a saturated solution of CaCl_2 , there are 0.32 kg of CaCl_2 (ionic) and 0.68 kg of water. First, convert the mass of CaCl_2 to moles.

$$0.32 \times 10^3 \text{ g CaCl}_2 \times \frac{1 \text{ mol CaCl}_2}{111.0 \text{ g CaCl}_2} = 2.88 \text{ mol CaCl}_2$$

(continued)

Then, find the molality of the CaCl_2 in the solution:

$$\text{Molality} = \frac{\text{mol CaCl}_2}{\text{kg H}_2\text{O}} \times \frac{2.88 \text{ mol CaCl}_2}{0.68 \text{ kg}} = 4.24 \text{ m}$$

For ionic solutions,

$$\Delta T_f = iK_f c_m$$

Calcium chloride dissolves to give three ions; thus $i = 3$, and

$$\Delta T_f = (1.858^\circ\text{C/m})(3)(4.24 \text{ m}) = 23.6^\circ\text{C}$$

$$T_f = 0.0^\circ\text{C} - 23.6^\circ\text{C} = -24^\circ\text{C}$$

$$12.93 \quad M = \frac{\pi}{RT} = \frac{7.7 \text{ atm}}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(37 + 273)\text{K}} = 0.302 = 0.30 \text{ mol/L}$$

$$12.94 \quad M = \frac{\pi}{RT} = \frac{5.50 \text{ atm}}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(25 + 273)\text{K}} = 0.2248 = 0.225 \text{ mol/L}$$

12.95 Consider the equation $\Delta T_f = iK_f c_m$. For CaCl_2 , $i = 3$; for glucose $i = 1$. Because $c_m = 0.10$ for both solutions, the product of i and c_m will be larger for CaCl_2 , as will ΔT_f . The solution of CaCl_2 will thus have the lower freezing point.

12.96 Consider the equation $\Delta T_b = iK_b c_m$. For CaCl_2 , $i = 3$; for KCl it is 2. The product of i and c_m will be larger for CaCl_2 , as will ΔT_b . The solution of KCl will thus have the lower boiling point.

12.97 Assume there is 1.000 L of the solution, which will contain eighteen mol H_2SO_4 , molar mass 98.09 g/mol. The mass of the solution is

$$\text{Mass solution} = \frac{18 \text{ mol} \times 98.09 \text{ g/mol}}{0.98} = 1802 \text{ g}$$

Thus, the density of the solution is

$$d = \frac{1802 \text{ g}}{1000 \text{ mL}} = 1.802 = 1.8 \text{ g/mL}$$

(continued)

The mass of water in the solution is

$$\text{mass H}_2\text{O} = 1802 \text{ g} \times 0.02 = \underline{36.0} \text{ g} = 0.0360 \text{ kg}$$

Thus, the molality of the solution is

$$m = \frac{18 \text{ mol H}_2\text{SO}_4}{0.0360 \text{ Kg}} = \underline{500} = 5 \times 10^2 \text{ m}$$

- 12.98 Assume there is 1.000 L of the solution, which will contain fifteen mol H_3PO_4 , molar mass 97.99 g/mol. The mass of the solution is

$$\text{Mass solution} = \frac{15 \text{ mol} \times 97.99 \text{ g/mol}}{0.85} = \underline{1729} \text{ g}$$

Thus, the density of the solution is

$$d = \frac{1729 \text{ g}}{1000 \text{ mL}} = \underline{1.729} = 1.7 \text{ g/mL}$$

The mass of water in the solution is

$$\text{mass H}_2\text{O} = 1729 \text{ g} \times 0.15 = \underline{259} \text{ g} = 0.259 \text{ kg}$$

Thus, the molality of the solution is

$$m = \frac{15 \text{ mol H}_3\text{PO}_4}{0.0259 \text{ Kg}} = \underline{57.9} = 58 \text{ m}$$

- 12.99 Use the freezing point depression equation to find the molality of the solution. The freezing point of pure cyclohexane is 6.55°C , and $K_f = 20.2^\circ\text{C/m}$. Thus

$$m = \frac{\Delta T_f}{K_f} = \frac{(6.55 - 5.28)^\circ\text{C}}{20.2^\circ\text{C/m}} = \underline{0.06287} \text{ m}$$

The moles of the compound are

$$\text{mol compound} = \frac{0.06287 \text{ mol}}{1 \text{ kg solvent}} \times 0.00538 \text{ kg} = \underline{3.382 \times 10^{-4}} \text{ mol}$$

(continued)

The molar mass of the compound is

$$\text{molar mass} = \frac{0.125 \text{ g}}{3.382 \times 10^{-4} \text{ mol}} = 369.6 \text{ g/mol}$$

The moles of the elements in 100 g of the compound are

$$\text{mol Mn} = 28.17 \text{ g Mn} \times \frac{1 \text{ mol}}{54.94 \text{ g Mn}} = 0.51274 \text{ mol}$$

$$\text{mol C} = 30.80 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g C}} = 2.5645 \text{ mol}$$

$$\text{mol O} = 41.03 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g O}} = 2.5644 \text{ mol}$$

This gives mole ratios of one mol Mn to five mol C to five mol O. Therefore, the empirical formula of the compound is MnC_5O_5 . The weight of this formula unit is approximately 195 amu. Since the molar mass of the compound is 370. g/mol, the value of n is

$$n = \frac{370 \text{ g/mol}}{195 \text{ g/unit}} = 2.00, \text{ or } 2$$

Therefore, the formula of the compound is $\text{Mn}_2\text{C}_{10}\text{O}_{10}$.

12.100 Use the freezing point depression equation to find the molality of the solution. The freezing point of pure cyclohexane is 6.55°C , and $K_f = 20.2^\circ\text{C/m}$. Thus

$$m = \frac{\Delta T_f}{K_f} = \frac{(6.55 - 5.23)^\circ\text{C}}{20.2^\circ\text{C/m}} = 0.06534 \text{ m}$$

The moles of the compound are

$$\text{mol compound} = \frac{0.06534 \text{ mol}}{1 \text{ kg solvent}} \times 0.0672 \text{ kg} = 4.391 \times 10^{-4} \text{ mol}$$

The molar mass of the compound is

$$\text{molar mass} = \frac{0.147 \text{ g}}{4.391 \times 10^{-4} \text{ mol}} = 334.8 \text{ g/mol}$$

(continued)

The moles of the elements in 100 g of the compound are

$$\text{mol Co} = 34.47 \text{ g Co} \times \frac{1 \text{ mol}}{58.93 \text{ g Co}} = 0.58493 \text{ mol}$$

$$\text{mol C} = 28.10 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g C}} = 2.3397 \text{ mol}$$

$$\text{mol O} = 37.43 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g O}} = 2.3394 \text{ mol}$$

This gives mole ratios of one mol Co to four mol C to four mol O. Therefore, the empirical formula of the compound is CoC_4O_4 . The weight of this formula unit is approximately 171 amu. Since the molar mass of the compound is 335. g/mol, the value of n is

$$n = \frac{335 \text{ g/mol}}{171 \text{ g/unit}} = 1.96, \text{ or } 2$$

Therefore, the formula of the compound is $\text{Co}_2\text{C}_8\text{O}_8$.

- 12.101 a. Use the freezing point depression equation to find the molality of the solution. The freezing point of pure water is 0°C and $K_f = 1.858^\circ\text{C/m}$. Thus

$$m = \frac{\Delta T_f}{K_f} = \frac{[0 - (-2.2)]^\circ\text{C}}{1.858^\circ\text{C/m}} = 1.184 \text{ m}$$

The moles of the compound are

$$\text{mol compound} = \frac{1.184 \text{ mol}}{1 \text{ kg solvent}} \times 0.100 \text{ kg} = 0.1184 \text{ mol}$$

The molar mass of the compound is

$$\text{molar mass} = \frac{18.0 \text{ g}}{0.1184 \text{ mol}} = 152.0 \text{ g/mol}$$

(continued)

The moles of the elements in 100 g of the compound are

$$\text{mol C} = 48.64 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g C}} = 4.05\text{00 mol}$$

$$\text{mol H} = 8.16 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g H}} = 8.0\text{95 mol}$$

$$\text{mol O} = 43.20 \text{ g O} \times \frac{1 \text{ mol}}{16.00 \text{ g O}} = 2.70\text{00 mol}$$

This gives mole ratios of 1.5 mol C to three mol H to one mol O. Multiplying by two gives ratios of three mol C to six mol H to two mol O. Therefore, the empirical formula of the compound is $\text{C}_3\text{H}_6\text{O}_2$. The weight of this formula unit is approximately 74 amu. Since the molar mass of the compound is 152.0 g/mol, the value of n is

$$n = \frac{152.0 \text{ g/mol}}{74 \text{ g/unit}} = 2.05, \text{ or } 2$$

Therefore, the formula of the compound is $\text{C}_6\text{H}_{12}\text{O}_4$.

b. The molar mass of the compound is (to the nearest tenth of a gram):

$$6(12.01) + 12(1.008) + 4(16.00) = 148.\text{156} = 148.2 \text{ g/mol}$$

12.102a. Use the freezing point depression equation to find the molality of the solution. The freezing point of naphthalene is 80.0°C and $K_f = 6.8^\circ\text{C/m}$. Thus

$$m = \frac{\Delta T_f}{K_f} = \frac{(80.0 - 78.0)^\circ\text{C}}{6.8^\circ\text{C/m}} = 0.29\text{4 m}$$

The moles of the compound are

$$\text{mol compound} = \frac{0.294 \text{ mol}}{1 \text{ kg solvent}} \times 0.00750 \text{ kg} = 2.2\text{1} \times 10^{-3} \text{ mol}$$

The molar mass of the compound is

$$\text{molar mass} = \frac{0.855 \text{ g}}{2.21 \times 10^{-3} \text{ mol}} = 38\text{7 g/mol}$$

(continued)

The moles of the elements in 100 g of the compound are

$$\text{mol C} = 39.50 \text{ g C} \times \frac{1 \text{ mol}}{12.01 \text{ g C}} = 3.2889 \text{ mol}$$

$$\text{mol H} = 2.21 \text{ g H} \times \frac{1 \text{ mol}}{1.008 \text{ g H}} = 2.192 \text{ mol}$$

$$\text{mol Cl} = 58.30 \text{ g Cl} \times \frac{1 \text{ mol}}{35.45 \text{ g Cl}} = 1.6446 \text{ mol}$$

This gives mole ratios of two mol C to 1.33 mol H to one mol Cl. Multiplying by three gives ratios of six mol C to four mol H to three mol Cl. Therefore, the empirical formula of the compound is $\text{C}_6\text{H}_4\text{Cl}_3$. The weight of this formula unit is approximately 182.5 amu. Since the molar mass of the compound is 387. g/mol, the value of n is

$$n = \frac{387 \text{ g/mol}}{182.5 \text{ g/unit}} = 2.12, \text{ or } 2$$

Therefore, the formula of the compound is $\text{C}_{12}\text{H}_8\text{Cl}_6$.

b. The molar mass of the compound is (to the nearest tenth of a gram):

$$12(12.01) + 8(1.008) + 6(35.45) = 364.884 = 364.9 \text{ g/mol}$$

12.103 Use the freezing point depression equation to find the molality of the solution. The freezing point of pure water is 0°C , and $K_f = 1.86^\circ\text{C/m}$. Thus

$$m = \frac{\Delta T_f}{K_f} = \frac{[0 - (-2.3)]^\circ\text{C}}{1.86^\circ\text{C/m}} = 1.24 \text{ m}$$

Since this is a relatively dilute solution, we can assume the molarity and molality of the fish blood are approximately equal. The calculated molarity is for the total number of particles, assuming they behave ideally.

$$\begin{aligned} \pi &= MRT = 1.24 \text{ mol/L} \times 0.08206 \text{ L}\cdot\text{atm/K}\cdot\text{mol} \times 298.2 \text{ K} \\ &= 30.3 = 30. \text{ atm} \end{aligned}$$

12.104 Use the osmotic pressure equation to find the molarity of the solution.

$$M = \frac{\pi}{RT} = \frac{18 \text{ atm}}{0.08206 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol} \times 295.2 \text{ K}} = 0.743 \text{ M}$$

Since this is a relatively dilute solution, we can assume the molality and molarity of the solution are approximately equal and the electrolyte (salt) is behaving ideally. The calculated temperature change is

$$\Delta T = K_f m = 1.86 \text{ }^\circ\text{C}/m \times 0.743 m = 1.38 \text{ }^\circ\text{C}$$

Therefore, the solution will freeze at $-1.4 \text{ }^\circ\text{C}$.

■ Solutions to Cumulative-Skills Problems

12.105 First, determine the initial moles of each ion present in the solution before any reaction has occurred. There are five ions present: Na^+ , $2 \times 0.375 = 0.750 \text{ mol}$; CO_3^{2-} , 0.375 mol ; Ca^{2+} , 0.125 mol ; Ag^+ , 0.200 mol ; and NO_3^- , $0.200 + 2 \times 0.125 = 0.450 \text{ mol}$. Two precipitates form, CaCO_3 and Ag_2CO_3 . The balanced net ionic equations for their formation are



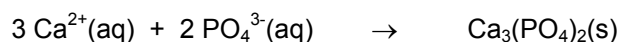
There is sufficient CO_3^{2-} to precipitate all of the Ca^{2+} and all of the Ag^+ ions. The moles of excess CO_3^{2-} are calculated as follows: $0.375 - 0.125 - 1/2 \times 0.200 = 0.150 \text{ mol}$ remaining. Since the volume is 2.000 L , the molarities of the various ions are

$$M \text{ of } \text{CO}_3^{2-} \text{ left} = 0.150 \text{ mol} \div 2.000 \text{ L} = 0.0750 \text{ M}$$

$$M \text{ of } \text{NO}_3^- \text{ left} = 0.450 \text{ mol} \div 2.000 \text{ L} = 0.225 \text{ M}$$

$$M \text{ of } \text{Na}^+ \text{ left} = 0.750 \text{ mol} \div 2.000 \text{ L} = 0.375 \text{ M}$$

- 12.106 First, determine the initial moles of each ion present in the solution before any reaction has occurred. There are five ions present: Na^+ , $3 \times 0.310 = 0.930$ mol; PO_4^{3-} , 0.310 mol; Ca^{2+} , 0.100 mol; Ag^+ , 0.150 mol; and NO_3^- , $0.150 + 2 \times 0.100 = 0.350$ mol. Two precipitates form, $\text{Ca}_3(\text{PO}_4)_2$ and Ag_3PO_4 . The balanced net ionic equations for their formation are

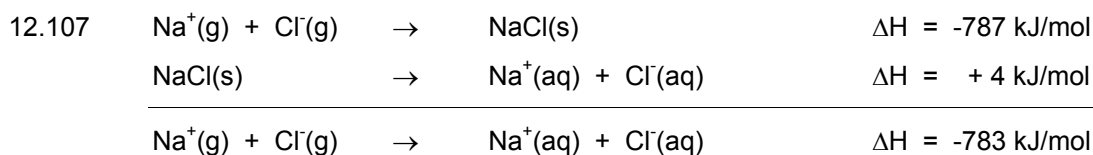


There is sufficient PO_4^{3-} to precipitate all of the Ca^{2+} and all of the Ag^+ ions. The moles of excess PO_4^{3-} is calculated as follows: $0.310 - 2/3 \times 0.100 - 1/3 \times 0.150 = 0.1933$ mol remaining. Since the volume is 4.000 L, the molarities of the various ions are

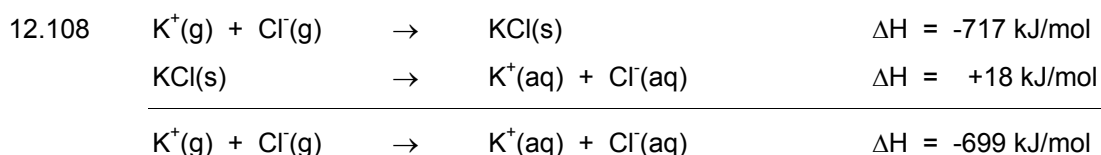
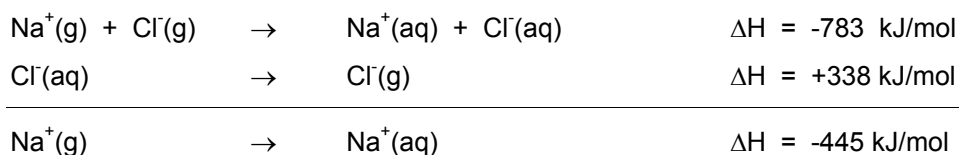
$$\text{M of } \text{PO}_4^{3-} \text{ left} = 0.1933 \text{ mol} \div 4.000 \text{ L} = 0.04833 = 0.0483 \text{ M}$$

$$\text{M of } \text{NO}_3^- \text{ left} = 0.350 \text{ mol} \div 4.000 \text{ L} = 0.0875 \text{ M}$$

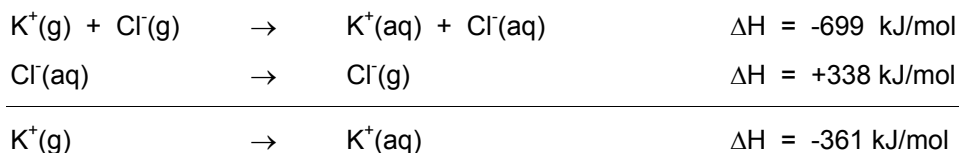
$$\text{M of } \text{Na}^+ \text{ left} = 0.930 \text{ mol} \div 4.000 \text{ L} = 0.2325 = 0.233 \text{ M}$$



The heat of hydration of Na^+ is



The heat of hydration of K^+ is



Less heat is evolved in the hydration of $\text{K}^+(\text{g})$ than in that of $\text{Na}^+(\text{g})$ because smaller ions have larger hydration energies.

$$12.109 \quad 15.0 \text{ g MgSO}_4 \cdot 7\text{H}_2\text{O} \times \frac{1 \text{ mol}}{246.5 \text{ g}} = 0.060854 \text{ mol MgSO}_4 \cdot 7\text{H}_2\text{O}$$

$$0.060854 \text{ mol MgSO}_4 \cdot 7\text{H}_2\text{O} \times \frac{7 \text{ mol H}_2\text{O}}{1 \text{ mol hydrate}} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 7.667 \text{ g H}_2\text{O}$$

$$\text{kg H}_2\text{O} = (100.0 \text{ g H}_2\text{O} + 7.667 \text{ g H}_2\text{O}) \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 0.10766 \text{ kg H}_2\text{O}$$

$$m = \frac{0.060854 \text{ mol MgSO}_4}{0.10766 \text{ kg H}_2\text{O}} = 0.5652 \text{ mol/kg} = 0.565 \text{ m}$$

$$12.110 \quad 15.0 \text{ g Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} \times \frac{1 \text{ mol}}{286.2 \text{ g}} = 0.05241 \text{ mol Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}, \text{ or Na}_2\text{CO}_3$$

$$0.05241 \text{ mol Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} \times \frac{10 \text{ mol H}_2\text{O}}{1 \text{ mol hydrate}} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 9.434 \text{ g H}_2\text{O}$$

$$\text{kg H}_2\text{O} = (100 \text{ g H}_2\text{O} + 9.434 \text{ g H}_2\text{O}) \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 0.10943 \text{ kg H}_2\text{O}$$

$$m = \frac{0.05241 \text{ mol Na}_2\text{CO}_3}{0.10943 \text{ kg H}_2\text{O}} = 0.4789 \text{ mol/kg} = 0.479 \text{ m}$$

$$12.111 \quad 15.0 \text{ g CuSO}_4 \cdot 5\text{H}_2\text{O} \times \frac{1 \text{ mol}}{249.7 \text{ g}} = 0.06007 \text{ mol CuSO}_4 \cdot 5\text{H}_2\text{O}, \text{ or CuSO}_4$$

$$100 \text{ g soln} \times \frac{1 \text{ mL soln}}{1.167 \text{ g soln}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.085689 \text{ L}$$

$$M = \frac{0.06007 \text{ mol CuSO}_4}{0.085689 \text{ L}} = 0.70102 = 0.701 \text{ mol/L}$$

$$12.112 \quad 20.0 \text{ g Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} \times \frac{1 \text{ mol}}{248.2 \text{ g}} = 0.08058 \text{ mol Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O, or Na}_2\text{S}_2\text{O}_3$$

$$100 \text{ g soln} \times \frac{1 \text{ mL soln}}{1.174 \text{ g soln}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.08517 \text{ L}$$

$$M = \frac{0.08058 \text{ mol Na}_2\text{S}_2\text{O}_3}{0.08517 \text{ L}} = 0.9461 = 0.946 \text{ mol/L}$$

$$12.113 \quad 0.159^\circ\text{C} \times \frac{m}{1.858^\circ\text{C}} = 0.08557 m = \frac{0.08557 \text{ mol AA} + \text{H}^+}{1000 \text{ g (or 1000 mL) H}_2\text{O}}$$

Note that $0.0830 \text{ mol AA} + \text{mol H}^+ = 0.08557 \text{ mol/L (AA} + \text{H}^+)$.

$$\text{Mol H}^+ = 0.08557 - 0.0830 = 0.00257 \text{ mol}$$

$$\begin{aligned} \text{Percent dissociation} &= \frac{\text{mol H}^+}{\text{mol AA}} \times 100\% = \frac{0.00257 \text{ mol}}{0.0830 \text{ mol}} \times 100\% \\ &= 3.09 = 3.1 \text{ percent} \end{aligned}$$

$$12.114 \quad 0.210^\circ\text{C} \times \frac{m}{1.858^\circ\text{C}} = 0.1130 m = \frac{0.1130 \text{ mol FA} + \text{H}^+}{1000 \text{ g (or 1000 mL) H}_2\text{O}}$$

Note that $0.109 \text{ mol FA} + \text{mol H}^+ = 0.113 \text{ mol/L (FA} + \text{H}^+)$.

$$\text{mol H}^+ = 0.113 - 0.109 = 0.0040 \text{ mol}$$

$$\text{percent dissociation} = \frac{\text{mol H}^+}{\text{mol FA}} \times 100\% = \frac{0.0040 \text{ mol}}{0.109 \text{ mol}} \times 100\% = 3.69 = 4 \text{ percent}$$

12.115 Calculate the empirical formula first, using the masses of C, O, and H in 1.000 g:

$$1.434 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 0.39132 \text{ g C}$$

$$0.783 \text{ g H}_2\text{O} \times \frac{2.016 \text{ g H}}{18.016 \text{ g H}_2\text{O}} = 0.08761 \text{ g H}$$

$$\text{g O} = 1.000 \text{ g} - 0.39132 \text{ g} - 0.08761 \text{ g} = 0.5211 \text{ g O}$$

(continued)

$$\text{Mol C} = 0.39132 \text{ g C} \times 1 \text{ mol}/12.01 \text{ g} = 0.03258 \text{ mol C (lowest integer} = 1)$$

$$\text{Mol H} = 0.08761 \text{ g H} \times 1 \text{ mol}/1.008 \text{ g} = 0.08691 \text{ mol H (lowest integer} = 8/3)$$

$$\text{Mol O} = 0.5211 \text{ g O} \times 1 \text{ mol}/16.00 \text{ g} = 0.03257 \text{ mol O (lowest integer} = 1)$$

Therefore, the empirical formula is $\text{C}_3\text{H}_8\text{O}_3$. The formula weight from the freezing point is calculated by first finding the molality:

$$0.0894 \text{ }^\circ\text{C} \times (m/1.858 \text{ }^\circ\text{C}) = 0.04811 = (0.04811 \text{ mol}/1000 \text{ g H}_2\text{O})$$

$$(0.04811 \text{ mol}/1000 \text{ g H}_2\text{O}) \times 25.0 \text{ g H}_2\text{O} = 0.001203 \text{ mol (in 25.0 g H}_2\text{O})$$

$$\text{Molar mass} = M_m = 0.1107 \text{ g}/0.001203 \text{ mol} = 92.02 \text{ g/mol}$$

Because this is also the formula weight, the molecular formula is also $\text{C}_3\text{H}_8\text{O}_3$.

12.116 Calculate the empirical formula first, using the masses of C, O, and H in 1.000 g:

$$1.418 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 0.3870 \text{ g C}$$

$$0.871 \text{ g H}_2\text{O} \times \frac{2.016 \text{ g H}}{18.016 \text{ g H}_2\text{O}} = 0.09746 \text{ g H}$$

$$\text{g O} = 1.000 \text{ g} - 0.3870 \text{ g} - 0.09746 \text{ g} = 0.5155 \text{ g O}$$

$$\text{Mol C} = 0.3870 \text{ g C} \times 1 \text{ mol}/12.01 \text{ g} = 0.0322 \text{ mol C (lowest integer} = 1)$$

$$\text{Mol H} = 0.09746 \text{ g H} \times 1 \text{ mol}/1.008 \text{ g} = 0.09668 \text{ mol H (lowest integer} = 3)$$

$$\text{Mol O} = 0.5155 \text{ g O} \times 1 \text{ mol}/16.00 \text{ g} = 0.03222 \text{ mol O (lowest integer} = 1)$$

The empirical formula is CH_3O . For the freezing-point calculation, the molality is calculated as follows:

$$0.0734 \text{ }^\circ\text{C} \times (m/1.858 \text{ }^\circ\text{C}) = 0.039505 = (0.39505 \text{ mol}/1000 \text{ g H}_2\text{O})$$

$$(0.39504 \text{ mol}/1000 \text{ g H}_2\text{O}) \times 45.0 \text{ g H}_2\text{O} = 0.001778 \text{ mol}$$

$$\text{Molar mass} = M_m = 0.1103 \text{ g}/0.0017778 \text{ mol} = 62.04 \text{ g/mol}$$

Because the formula weight of CH_3O is 31.02, the molecular formula is $\text{C}_2\text{H}_6\text{O}_2$.

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As of Last Complete Printing
Number of Pages: 43
Number of Words: 8,934 (approx.)
Number of Characters: 50,928 (approx.)