Practice Exercises

13.1 \[ C_{H_2S} = k_H P_{H_2S} \]

\[ P_{H_2S} = 1.0 \text{ atm} \]

\[ C_{H_2S} = \left( \frac{0.11 \text{ mol } H_2S}{1 \text{ mol } H_2S} \right) \frac{34.08 \text{ g } H_2S}{1 \text{ L}} = 3.7 \text{ g L}^{-1} \]

Hydrogen sulfide is more soluble in water than nitrogen and oxygen. Hydrogen sulfide reacts with the water to form hydronium ions and HS⁻.

13.2 With one atmosphere of air, the concentrations of the gases in the water depends on the partial pressures of the gases.

Oxygen: \[ C_{O_2} = \left( \frac{0.00430 \text{ g } O_2}{100 \text{ mL } H_2O} \right) \frac{159 \text{ mm Hg}}{760 \text{ mm Hg}} = 0.899 \text{ mg } O_2/100 \text{ mL} \]

\[ g \text{ O}_2 \text{ in } 125 \text{ g of water} = 0.899 \text{ mg } O_2/100 \text{ mL} \times 125 \text{ mL} = 1.12 \text{ g} \]

Nitrogen: \[ C_{N_2} = \left( \frac{0.00190 \text{ g } N_2}{100 \text{ mL } H_2O} \right) \frac{593 \text{ mm Hg}}{760 \text{ mm Hg}} = 1.48 \text{ mg } N_2/100 \text{ mL} \]

\[ g \text{ N}_2 \text{ in } 125 \text{ mL of water} = 1.48 \text{ mg } N_2/100 \text{ mL} \times 125 \text{ mL} = 1.85 \text{ g} \]

13.3 A 10% w/w solution of sucrose will need 10 grams of sucrose for each 100 g of solution. For a solution with 45.0 g of sucrose:

\[ 10\% \text{ solution} = \frac{45.0 \text{ g sucrose}}{x \text{ g solution}} \]

\[ x = 450 \text{ g solution} \]

\[ g \text{ water} = 450 \text{ g solution} - 45.0 \text{ g sucrose} \]

\[ g \text{ water} = 405 \text{ g water} \]

\[ \text{mL water} = (405 \text{ g water}) \left( \frac{1 \text{ cm}^3}{0.9982 \text{ g}} \right) = 405.7 \text{ mL water} \]

13.4 The total mass of the solution is to be 25.0 g. If the solution is to be 1.00 % (w/w) NaBr, then the mass of NaBr will be: 25.0 g \times 1.00 g NaBr/100 g solution = 0.250 g NaBr. We therefore need 0.250 g of NaBr and (25.0 – 0.250) = 24.75 g H₂O. The volume of water that is needed is: 24.75 g/0.988 g/mL = 25.0 mL H₂O.

13.5 An HCl solution that is 37 % (w/w) has 37 grams of HCl for every 1.0 \times 10^2 grams of solution.

\[ g \text{ solution} = (7.5 \text{ g HCl}) \left( \frac{1.0 \times 10^2 \text{ g solution}}{37 \text{ g HCl}} \right) = 2.0 \times 10^1 \text{ g solution} \]

13.6 We need to know the number of moles of Na₂SO₄ and the number of kg of water.

44.00 g Na₂SO₄ \div 142.0 \text{ g/mol} = 0.3099 \text{ mol Na₂SO₄} \]

250 g H₂O \times 1 \text{ kg/1000 g} = 0.250 \text{ kg H₂O} \]

The molality is thus given by:

\[ m = 0.3099 \text{ mol/0.25 kg} = 1.239 \text{ mol Na₂SO₄/kg H₂O} = 1.239 \text{ m} \]

Molarity is moles solute per liters of solution. The moles of solute is the same for molarity and molality but the volume of solution would be larger than the kilograms of solvent so \( M < m \).
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13.7 \[ g \text{ CH}_3\text{OH for } 0.050 \text{ m} = 0.200 \text{ kg} \left( \frac{0.050 \text{ mol CH}_3\text{OH}}{\text{kg H}_2\text{O}} \right) \left( \frac{32.0 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right) \]

= 0.320 g

\[ g \text{ CH}_3\text{OH for } 0.100 \text{ m} = 0.200 \text{ kg} \left( \frac{0.100 \text{ mol CH}_3\text{OH}}{\text{kg H}_2\text{O}} \right) \left( \frac{32.0 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right) \]

= 0.640 g

\[ g \text{ CH}_3\text{OH for } 0.150 \text{ m} = 0.200 \text{ kg} \left( \frac{0.150 \text{ mol CH}_3\text{OH}}{\text{kg H}_2\text{O}} \right) \left( \frac{32.0 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right) \]

= 0.960 g

\[ g \text{ CH}_3\text{OH for } 0.200 \text{ m} = 0.200 \text{ kg} \left( \frac{0.200 \text{ mol CH}_3\text{OH}}{\text{kg H}_2\text{O}} \right) \left( \frac{32.0 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right) \]

= 1.28 g

\[ g \text{ CH}_3\text{OH for } 0.250 \text{ m} = (0.200 \text{ kg H}_2\text{O}) \left( \frac{0.250 \text{ mol CH}_3\text{OH}}{\text{kg H}_2\text{O}} \right) \left( \frac{32.0 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right) \]

= 1.60 g CH$_3$OH

13.8 First we need to find the number of grams of Fe(NO$_3$)$_3$ for each kg of solvent.

\[ 0.853 \text{ m Fe(NO}_3\text{)}_3 = 0.853 \text{ mol Fe(NO}_3\text{)}_3 \left( \frac{241.86 \text{ g Fe NO}_3\text{)}_3}{1 \text{ mol Fe NO}_3\text{)}_3} \right) = 206.3 \text{ g Fe(NO}_3\text{)}_3 \]

Then we need to find how the ratio of the moles of Fe(NO$_3$)$_3$ to the mass of the solution:

\[ \text{ratio} = \frac{0.853 \text{ mol Fe NO}_3\text{)}_3}{1000 \text{ g H}_2\text{O} + 206.3 \text{ g Fe NO}_3\text{)}_3} = 7.072 \times 10^{-4} \text{ mol Fe(NO}_3\text{)}_3/\text{g solution} \]

(a) \[ g \text{ solution} = (0.0200 \text{ mol Fe(NO}_3\text{)}_3) \left( \frac{1 \text{ g solvent}}{7.07 \times 10^{-4} \text{ mol Fe NO}_3\text{)}_3} \right) = 28.3 \text{ g solution} \]

(b) \[ g \text{ sol'n} = (0.0500 \text{ mol Fe}^{3+}) \left( \frac{1 \text{ mol Fe NO}_3\text{)}_3}{1 \text{ mol Fe}^{3+}} \right) \left( \frac{1 \text{ g solvent}}{7.07 \times 10^{-4} \text{ mol Fe NO}_3\text{)}_3} \right) = 70.7 \text{ g sol'n} \]

(c) \[ g \text{ sol'n} = (0.00300 \text{ mol NO}_3^-) \left( \frac{1 \text{ mol Fe NO}_3\text{)}_3}{3 \text{ mol NO}_3^-} \right) \left( \frac{1 \text{ g solvent}}{7.07 \times 10^{-4} \text{ mol Fe NO}_3\text{)}_3} \right) = 1.41 \text{ g sol'n} \]

13.9 If a solution is 52\% NaOH, then it has 52 g of NaOH for each 100 g of solution. The mass of water is 48 g of water for 52 g of NaOH. To calculate the moality of the solution, we need to find the moles of NaOH for each kilogram of water.

\[ m = \left( \frac{52 \text{ g NaOH}}{48 \text{ g H}_2\text{O}} \right) \left( \frac{1 \text{ mol NaOH}}{40.0 \text{ g NaOH}} \right) \left( \frac{1000 \text{ g H}_2\text{O}}{1 \text{ kg H}_2\text{O}} \right) = 27 \text{ mol NaOH/kg H}_2\text{O} = 27 \text{ m NaOH} \]

13.10 If a solution is 37.0\% (w/w) HCl, then 37.0\% of the mass of any sample of such a solution is HCl and (100.0 – 37.0) = 63.0\% of the mass is water. In order to determine the moality of the solution, we can conveniently choose 100.0 g of the solution as a starting point. Then 37.0 g of this solution are HCl and 63.0 g are H$_2$O. For moality, we need to know the number of moles of HCl and the mass in kg of the solvent.

\[ 37.0 \text{ g HCl} \div 36.46 \text{ g/mol} = 1.01 \text{ mol HCl} \]

\[ 63.0 \text{ g H}_2\text{O} \times 1 \text{ kg/1000 g} = 0.0630 \text{ kg H}_2\text{O} \]
molality = mol HCl/kg H₂O = 1.01 mol/0.0630 kg = 16.1 m

13.11 40.0 % HBr means 40 g of HBr per 100 g of solution. Since molarity is defined as moles of solute per liter of solution we need to determine the volume of solution. The density of the solution allows us to determine the volume of the solution

\[
\frac{40 \text{ g HBr}}{100 \text{ g solution}} \times \frac{1 \text{ mol HBr}}{80.91 \text{ g}} \times \frac{1.38 \text{ g}}{\text{mL}} \times \frac{1000 \text{ mL}}{L} = 6.82M
\]

13.12 First determine the number of moles of Al(NO₃)₃ dissolved in the liter of water.

\[
\text{mol Al(NO₃)₃} = (1.00 \text{ g Al(NO₃)₃}) \left(\frac{1 \text{ mol Al NO₃}₃}{212.996 \text{ g Al NO₃}₃}\right) = 0.00469 \text{ mol Al(NO₃)₃}
\]

Next find the mass of the water:

\[
\text{g H}_₂\text{O} = (1.00 \text{ L H}_₂\text{O}) \left(\frac{1000 \text{ mL H}_₂\text{O}}{1 \text{ L H}_₂\text{O}}\right) \left(\frac{0.9982 \text{ g H}_₂\text{O}}{1 \text{ mL H}_₂\text{O}}\right) = 998.2 \text{ g H}_₂\text{O}
\]

To find the molarity of the solution, first we need to find the mass of the solution, and then the volume of the solution:

\[
\text{g solution} = 998.2 \text{ g H}_₂\text{O} + 1.00 \text{ g Al(NO₃)₃} = 999.2 \text{ g solution}
\]

\[
\text{L solution} = (999.2 \text{ solution}) \left(\frac{1 \text{ mL solution}}{0.9989 \text{ g solution}}\right) \left(\frac{1 \text{ L solution}}{1000 \text{ mL solution}}\right) = 1.0003 \text{ L}
\]

\[
M \text{ of solution} = \frac{0.00469 \text{ mol Al NO₃}₃}{1.0003 \text{ L solution}} = 0.00469 M \text{ Al(NO₃)₃}
\]

The molality of the solution can also be determined

\[
m \text{ of solution} = \frac{0.00469 \text{ mol Al NO₃}₃}{0.9982 \text{ kg H}_₂\text{O}} = 0.00470 m \text{ Al(NO₃)₃}
\]

13.13 First determine the number of moles of each component of the solution:

For C₁₆H₁₉₂O₄, 20.0 g/278 g/mol = 0.0719 mol

For C₅H₁₂, 50.0 g/72.2 g/mol = 0.692 mol

The mole fraction of solvent is:

0.692 mol/(0.692 mol + 0.0719 mol) = 0.906

Using Raoult’s Law, we next find the vapor pressure to expect for the solution, which arises only from the solvent (since the solute is known to be nonvolatile):

\[
P_{\text{solvent}} = \chi_{\text{solvent}} \times P^o_{\text{solvent}} = 0.906 \times 541 \text{ torr} = 4.90 \times 10^2 \text{ torr}
\]

13.14

\[
P_{\text{acetone}} = \chi_{\text{acetone}} \times P^o_{\text{acetone}}
\]

mol acetone = (156 g acetone) \left(\frac{1 \text{ mol acetone}}{58.0 \text{ mol acetone}}\right) = 2.690 \text{ mol acetone}

Do not round your answers until the end. The moles of stearic acid is small compared to the moles of acetone and rounding error may give you too high of a mass of stearic acid.

\[
155 \text{ torr} = \left(\frac{2.690 \text{ mol acetone}}{2.690 \text{ mol acetone} + x \text{ mol stearic acid}}\right) \times 162 \text{ torr}
\]

\[
0.957 = \left(\frac{2.690 \text{ mol acetone}}{2.690 \text{ mol acetone} + x \text{ mol stearic acid}}\right)
\]

2.574 mol acetone + 0.957x mol stearic acid = 2.690 mol acetone

0.957x mol stearic acid = 0.116

x mol stearic acid = 0.121 mol stearic acid
Finally solve to find the number of grams of stearic acid
\[
g \text{ stearic acid} = (0.121 \text{ mol stearic acid}) \left( \frac{284.5 \text{ g stearic acid}}{1 \text{ mol stearic acid}} \right) = 34.5 \text{ g stearic acid}
\]

13.15\[P_{\text{cyclohexane}} = \chi_{\text{cyclohexane}} \times P^o_{\text{cyclohexane}} = 0.750 \times 66.9 \text{ torr} = 50.2 \text{ torr} \]
\[P_{\text{toluene}} = \chi_{\text{toluene}} \times P^o_{\text{toluene}} = 0.250 \times 21.1 \text{ torr} = 5.28 \text{ torr} \]
\[P_{\text{total}} = P_{\text{cyclohexane}} + P_{\text{toluene}} = 33.5 \text{ torr} + 10.6 \text{ torr} = 55.4 \text{ torr} \]

32.16 First we need to find the moles of the cyclohexane and the moles of toluene.
\[\text{mol cyclohexane} = (122 \text{ g cyclohexane}) \left( \frac{1 \text{ mol cyclohexane}}{84.15 \text{ g cyclohexane}} \right) = 1.450 \text{ mol cyclohexane} \]
\[\text{mol toluene} = (122 \text{ g toluene}) \left( \frac{1 \text{ mol toluene}}{92.14 \text{ g toluene}} \right) = 1.324 \text{ mol toluene} \]

Now, find the cyclohexane and the toluene
\[\chi_{\text{cyclohexane}} = \frac{1.450 \text{ mol cyclohexane}}{1.450 \text{ mol cyclohexane} + 1.324 \text{ mol toluene}} = 0.523 \]
\[\chi_{\text{toluene}} = 1 - \chi_{\text{cyclohexane}} = 1 - 0.523 = 0.477 \]
\[P_{\text{cyclohexane}} = \chi_{\text{cyclohexane}} \times P^o_{\text{cyclohexane}} = 0.523 \times 66.9 \text{ torr} = 35.0 \text{ torr} \]
\[P_{\text{toluene}} = \chi_{\text{toluene}} \times P^o_{\text{toluene}} = 0.477 \times 21.1 \text{ torr} = 10.1 \text{ torr} \]
\[P_{\text{total}} = P_{\text{cyclohexane}} + P_{\text{toluene}} = 35.0 \text{ torr} + 10.1 \text{ torr} = 45.1 \text{ torr} \]

13.17 First convert \(^{0}\text{F}\) to \(^{0}\text{C}\). \(235 \, {^{0}\text{F}} = 112.78 \, {^{0}\text{C}}\) and \(240 \, {^{0}\text{F}} = 115.56 \, {^{0}\text{C}}\).
The corresponding boiling point elevations are then 12.78 \(^{0}\text{C}\) and 15.56 \(^{0}\text{C}\) respectively.
The molality of the two solutions is given as:
\[m_1 = \frac{12.78 \, {^{0}\text{C}}}{0.51 \, {^{0}\text{C}} \text{ m}^{-1}} = 25.06 \text{ m}\]
\[m_2 = \frac{15.56 \, {^{0}\text{C}}}{0.51 \, {^{0}\text{C}} \text{ m}^{-1}} = 30.51 \text{ m}\]

The solution that boils at 235 \(^{0}\text{F}\) has the following mass percent of sugar:
\[
\frac{25.06 \text{ mol} \text{ C}_2\text{H}_2\text{O}_{11}}{1000 \text{ g H}_2\text{O}} \times \frac{342 \text{ g}}{\text{mol C}_2\text{H}_2\text{O}_{11}} = \frac{8.57 \text{ g C}_2\text{H}_2\text{O}_{11}}{\text{g H}_2\text{O}}
\]
Total grams of the solution = \(8.57 \) g C\(_2\)H\(_2\)O\(_{11}\) + \(1.00 \) g H\(_2\)O = \(9.57 \) g

Percent sugar = \(\frac{8.57 \text{ g}}{9.57 \text{ g}} \times 100 \% = 89.5\%\)

For the solution that boils at 240 \(^{0}\text{F}\)
\[
\frac{30.51 \text{ mol} \text{ C}_2\text{H}_2\text{O}_{11}}{1000 \text{ g H}_2\text{O}} \times \frac{342 \text{ g}}{\text{mol C}_2\text{H}_2\text{O}_{11}} = \frac{10.43 \text{ g C}_2\text{H}_2\text{O}_{11}}{\text{g H}_2\text{O}}
\]
Percent sugar = \(\frac{10.43 \text{ g}}{11.43 \text{ g}} \times 100 \% = 91.2\%\)

The mass percent range for the solutions is 89.5 % to 91.2 %

13.18 \(\Delta T_b = K_b \times m = 0.51 \, {^{0}\text{C}} \text{ m}^{-1} \times x \, m = 2.36 \, {^{0}\text{C}}\)
x \(m = 4.627 \text{ m}\)
To find the number of grams of glucose, first we need to find the number of moles of glucose.
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mol glucose = \((m\text{ solution})(kg\text{ solvent})\)

\[
g\text{ glucose} = (1.18\text{ mol glucose})\left(\frac{180.9\text{ g glucose}}{1\text{ mol glucose}}\right) = 213\text{ g glucose}
\]

13.19 It is first necessary to obtain the values of the freezing point of pure benzene and the value of \(K_f\) for benzene from Table 13.4 of the text. We proceed to determine the number of moles of solute that are present and that have caused this depression in the freezing point: \(\Delta T = K_f m\)

\[
\therefore m = \frac{\Delta T}{K_f} = \frac{(5.45\text{ °C} - 4.13\text{ °C})/(5.07\text{ °C kg mol}^{-1})}{0.260}\text{ m}
\]

Next, use this molality to determine the number of moles of solute that must be present:

\[
0.260\text{ mol solute/kg solvent} \times 0.0850\text{ kg solvent} = 0.0221\text{ mol solute}
\]

Last, determine the formula mass of the solute:

\[
3.46\text{ g/0.0221 mol} = 157\text{ g mol}^{-1}
\]

13.20 To find the molar mass of the substance, first, we need to find the molality of the solution from the freezing point depression, and then using the 5.0% (wt/wt) amount, determine the moles of the solute.

\[
m = \frac{\Delta T}{K_f} = \frac{80.2\text{ °C} - 77.3\text{ °C}}{6.9\text{ °C m}^{-1}} = 0.420\text{ m}
\]

Assume there is 100 g of solution:

\[
5.0\%\ (\text{wt/wt}) = \frac{5\text{ g unknown substance}}{5\text{ g unknown substance} + 95\text{ g naphthalene}}
\]

We have 95 g of naphthalene, or 0.095 kg naphthalene and 5 g of the unknown.

Using the equation for molality, we can determine the number of moles of the unknown

\[
\text{mol unknown} = (m\text{ solution})(kg\text{ solvent}) = (0.420\text{ m})(0.095\text{ kg naphthalene}) = 0.0399\text{ mol unknown}
\]

\[
\text{molar mass} = \frac{5.0\text{ g unknown}}{0.0399\text{ mol unknown}} = 125\text{ g mol}^{-1}
\]

13.21 Use the equation \(\Pi = MRT\):

\[
\frac{5\text{ g protein}}{1 \text{ mol protein}}\left(\frac{1}{235,000\text{ g protein}}\right) = 2.13 \times 10^{-4}\text{ M solution}
\]

\[
R = 0.0821\text{ L atm/K mol}
\]

\[
T = 4.0 + 273.2 = 277.2\text{ K}
\]

\[
\Pi = (2.13 \times 10^{-4}\text{ M solution})(0.0821\text{ L atm/K mol})(277.2\text{ K}) = 4.84 \times 10^{-3}\text{ atm}
\]

\[
\text{mm Hg} = 4.84 \times 10^{-3}\text{ atm}\left(\frac{760\text{ mm Hg}}{1\text{ atm}}\right) = 3.68\text{ mm Hg}
\]

\[
\text{mm H}_2\text{O} = 3.68\text{ mm Hg}\left(\frac{13.6\text{ mm H}_2\text{O}}{1\text{ mm Hg}}\right) = 50.0\text{ mm H}_2\text{O}
\]

13.22 We can use the equation \(\Pi = MRT\):

\[
\Pi = (0.0115\text{ M})(0.0821\text{ L atm/K mol})(310\text{ K})
\]

\[
\Pi = 0.293\text{ atm}
\]

\[
\Pi = 0.293\text{ atm} \times 760\text{ torr atm}^{-1} = 222\text{ torr}
\]

To determine the boiling and freezing temperatures of the solution we can assume that the molality is equal to the molarity. At low concentrations the two values are nearly identical.

\[
T_f = 0\text{ °C} - mK_f = -0.0115\text{ m} \times 1.86\text{ °C m}^{-1} = -0.021\text{ °C}
\]

\[
T_{bp} = 100\text{ °C} + mK_{bp} = 100 + 0.0115\text{ m} \times 0.51\text{ °C m}^{-1} = 100.006\text{ °C}
\]

Note that significant figures rules were not used for the boiling point answer.
13.23 \[ \Pi = MRT \]
\[ \Pi = 6.45 \text{ cm water} \left( \frac{10 \text{ mm H}_2\text{O}}{1 \text{ cm H}_2\text{O}} \right) \left( \frac{1.00 \text{ g mL}^{-1}}{13.6 \text{ g mL}^{-1}} \right) \left( \frac{1 \text{ atm}}{760 \text{ mm Hg}} \right) = 6.24 \times 10^{-3} \text{ atm} \]

\[ R = 0.0821 \text{ L atm / mol K} \]
\[ T = 277 \text{ K} \]

\[ M = \frac{\Pi}{RT} = \frac{6.24 \times 10^{-3} \text{ atm}}{0.0821 \text{ L atm/mol K} \cdot 277 \text{ K}} = 2.74 \times 10^{-4} \text{ mol L}^{-1} \]

\[ \text{mol protein} = (2.74 \times 10^{-3} \text{ mol L}^{-1})(0.1000 \text{ L}) = 2.74 \times 10^{-5} \text{ mol} \]

\[ \text{molar mass} = \frac{0.1372 \text{ g protein}}{2.74 \times 10^{-5} \text{ mol protein}} = 5.00 \times 10^{3} \text{ g mol}^{-1} \]

13.24 We can use the equation \( \Pi = MRT \), remembering to convert pressure to atm:

\[ \text{atm} = \left( \frac{25.0 \text{ torr}}{1 \text{ atm}} \right) \left( \frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.0329 \text{ atm} \]

\[ \Pi = 0.0329 \text{ atm} = M \times (0.0821 \text{ L atm/K mol})(298 \text{ K}) \]

\[ M = 1.34 \times 10^{-3} \text{ mol L}^{-1} \]

\[ \text{mol} = 1.34 \times 10^{-3} \text{ mol L}^{-1} \times 0.100 \text{ L} = 1.34 \times 10^{-4} \text{ mol} \]

\[ \text{formula mass} = \frac{72.4 \times 10^{-3} \text{ g}}{1.34 \times 10^{-4} \text{ mol}} = 5.38 \times 10^{2} \text{ g mol}^{-1} \]

13.25 For the solution as if the solute were 100% dissociated:

\[ \Delta T = (1.86 \text{ °C m}^{-1})(2 \times 0.237 \text{ m}) = 0.882 \text{ °C} \]

For the solution as if the solute were 0% dissociated:

\[ \Delta T = (1.86 \text{ °C m}^{-1})(1 \times 0.237 \text{ m}) = 0.441 \text{ °C} \]

13.26 Use the freezing point depression equation:

\[ \Delta T = K_f m \]

Remember that there are two moles of ions for each mole of MgSO₄.

\[ K_f \text{ water} = 1.86 \text{ °C m}^{-1} \]

(a) For 0.1 \text{ m MgSO}_4 \quad m = 0.2 \text{ m}

\[ \Delta T = (1.86 \text{ °C m}^{-1})(0.2 \text{ m}) = 0.372 \text{ °C} \quad \text{thus, } T_f = -0.372 \text{ °C} \]

(b) For 0.01 \text{ m MgSO}_4 \quad m = 0.02 \text{ m}

\[ \Delta T = (1.86 \text{ °C m}^{-1})(0.02 \text{ m}) = 0.0372 \text{ °C} \quad \text{thus, } T_f = -0.0372 \text{ °C} \]

(c) For 0.001 \text{ m MgSO}_4 \quad m = 0.002 \text{ m}

\[ \Delta T = (1.86 \text{ °C m}^{-1})(0.002 \text{ m}) = 0.00372 \text{ °C} \quad \text{thus, } T_f = -0.00372 \text{ °C} \]

The first freezing point depression could be measured using a laboratory thermometer that can measure 1 °C increments.

**Review Questions**

13.1 This event, diagrammed in Figure 13.1, is due to the tendency for all systems to proceed spontaneously towards a state with a higher degree of randomness (disorder).

13.2 First, the tendency towards randomness drives the solution process, and second, the new forces of attraction between solute and solvent molecules drive the process. Thus the relative degree of solute–solute, solvent–solvent and solute–solvent interactions will determine if a solute is soluble in a solvent.
13.3 Since water and methanol both have OH groups, there can be hydrogen bonding between a water molecule and a methanol molecule. This allows any proportion of methyl alcohol in water to be nearly as stable as either separate water samples or separate methyl alcohol samples.

13.4 Water molecules are tightly linked to one another by hydrogen bonding. In hexane, however, which is a nonpolar organic substance, we have only weak London forces of attraction. This means that hexane as a solute in water offers no advantage in attraction to individual water molecules, and the solvent is therefore not disrupted to allow the solute to dissolve.

13.5 The dipole moments of water molecules can be oriented so as to stabilize both the dissolved cation and the dissolved anion.

13.6 There is no solvating force provided by carbon tetrachloride that can overcome and offset the very strong ion–ion forces of the solid KCl sample.

13.7 Since the enthalpy of solution is positive, the process is endothermic. The system thus requires heat for the dissolving process, and the heat flow should cause the temperature to decrease as the solute dissolves.

13.8 The lattice energy is numerically larger since that step is endothermic, that is it requires energy to separate the particles.

13.9 The $\text{Al}^{3+}$ ion, having the greater positive charge, should have the larger hydration energy.

13.10 When a gas dissolves in a liquid, there is no endothermic step analogous to the lattice energy of a solid. The only enthalpy change is the one associated with hydration, and this is always negative.

13.11 There is a greater attraction between water and acetone molecules in the resulting solution than there is among acetone molecules in the starting pure solute or water molecules in the starting pure solvent.

13.12 The disruption of ethyl alcohol and the disruption of hexane together cost more energy than is gained on formation of the solution. This is because the two liquids are not alike; ethyl alcohol is a polar substance with hydrogen bonding, whereas hexane is a nonpolar liquid having only London forces.

13.13 If the solution becomes cool, it is an endothermic process and the $\Delta H_{\text{obs}}$ is positive. This is because the solution is absorbing heat from the surroundings. The solubility is likely to increase with increasing temperature because heat is required for the reaction to proceed, so increasing the temperature increases the amount of heat available.

13.14 The fact that the $\Delta H_{\text{obs}}$ value for the formation of a mixture of $\text{A}$ and $\text{B}$ is zero, implies that the relative strengths of $\text{A}–\text{A}$, $\text{B}–\text{B}$, and $\text{A}–\text{B}$ intermolecular attractions are similar.

13.15 We can estimate from Figure 13.10 that the solubility of $\text{NH}_4\text{NO}_3$ in 100 g of $\text{H}_2\text{O}$ is 550 g at 75 °C and 165 g at 15 °C. The amount of solid that will crystallize is the difference between these two solubilities, namely $550 – 165 = 385$ g in 100 g of solvent. However, we have 125 g of solvent.

\[
\frac{385 \text{ g} \text{NH}_4\text{NO}_3}{100 \text{ g solvent}} \times 125 \text{ g solvent} = 481 \text{ g NH}_4\text{NO}_3 \text{ will separate out of the solution.}
\]

13.16 Oxygen solubility increases as the temperature decreases. The larger fish will need more oxygen and will be found in the colder areas of lake bottoms.

13.17 Henry's Law is the statement, applied to the dissolving of a gas in a solvent, that at a given temperature, the concentration ($C_g$) of the gas in a solution is directly proportional to the partial pressure ($P_g$) of the gas on the solution, where $k$ in the following equation is the constant of proportionality: $C_g = k \times P_g$. As
discussed in the text, an alternate statement expresses the relationship of concentration at one pressure \( P_1 \) to the concentration that would exist at some new pressure \( P_2 \):  
\[
C_1 / P_1 = C_2 / P_2
\]

13.18 The atmospheric pressure on a mountain is less than the atmospheric pressure at sea level. From Henry's Law, as the partial pressure of oxygen decreases, the concentration of the oxygen also decreases. Therefore, there is less oxygen to sustain life in mountain streams.

13.19 Ammonia is more soluble in water than nitrogen because ammonia is able to hydrogen bond with solvent molecules, whereas nitrogen cannot. Nitrogen is a nonpolar molecular substance, whereas ammonia is a polar substance capable of hydrogen bonding. Also, ammonia reacts with water to form nonvolatile ions:

\[
\text{NH}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^- (aq)
\]

Hydrogen chloride would have a high solubility in water because it ionizes in water:

\[
\text{HCl}(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^- (aq)
\]

13.20 When the cap is removed from a bottle of carbonated beverage, the liquid fizzes because \( \text{CO}_2 \) is being released from the liquid. When the cap is on, the \( \text{CO}_2 \) fills the space above the liquid until equilibrium is established between the gas and the liquid. After the cap is removed, the equilibrium is disrupted and more of the gas leaves the solution. This is the fizzing.

13.21 (a) mole fraction = moles component/total number of moles  
(b) mole percent = mole fraction \times 100\%  
(c) molality = moles solute/kg solvent  
(d) percent by mass = (mass component/total mass) \times 100\%

The maximum value for mole fraction is 1, and the maximum value for mole percent and percent by mass is 100 \%. 

13.22 Molality is independent of temperature. Molarity decreases with increasing temperature because the volume of the solvent increases with increasing temperature.

13.23 The molarity will be greater than 1.0. Since the density of the solution is greater than one, the mass of the solution in kg will be greater than its volume in liters.

13.24 A colligative property of a solution is one that depends only on the molal concentration of the solute particles, and not on the identity of the solute.

13.25 A solution is ideal if the sum of the partial pressures of the components of the solution equals the observed vapor pressure of the solution, i.e., if the solution obeys Raoult's Law. Also, it should be true that the heat of solution is nearly zero.

13.26 A positive deviation indicates that the vapor pressure of the real solution is greater than expected if the solution behaved ideally. Positive deviations result when mixtures with weaker intermolecular forces of attraction between the two substances as compared to the intermolecular forces of the pure substances are formed.

13.27 When a solute is dissolved in a solvent, the vapor pressure is lowered. As a result, the boiling point is increased to a temperature where the vapor pressure is high enough to once again allow boiling to occur. This affect also reduces the triple point and the entire solid-liquid equilibrium curve on a phase diagram shifts to lower temperatures. The net result is a lowering of the freezing point.

13.28 These are semipermeable because only certain substances are able to pass through the membrane. A nonpermeable material would allow nothing to pass through.
An osmotic membrane allows only solvent to pass, whereas a dialyzing membrane allows solvated ions of a certain minimum size to pass as well as solvent molecules. A dialyzing membrane prevents the passage of only certain solute particles, usually those of large size, such as colloid particles.

The side of the membrane less concentrated in solute will be more concentrated in solvent. Therefore, the escaping tendency of the solvent will be greater than on the side of the membrane less concentrated in solute. The solvent will shift through the membrane from the side less concentrated in solute to the side more concentrated in solute.

The solution that loses solvent into the other solution is the one with the lower molarity.

In each case, the osmotic pressure $\Pi$ is given by the equation: $\Pi = M \times R \times T$. Since we do not know either the density of the solution or the volume of the solution, we cannot convert values for % by mass into molarities. However, we do know that glucose, having the smaller molecular mass, has the higher molarity, and we conclude that it will have the larger osmotic pressure.

By the "association of solute particles" we mean that some particles are attracted to others, or that solvent does not perfectly insulate solute particles from attachment to one another. This is another way of saying that there is less than 100% dissociation or dissolution of solute in such a solution.

Colligative properties depend on the number of particles in solution. If association of solute occurs, this decreases the number of independent particles in solution. Therefore, the freezing point will not be lowered as much as expected, and the boiling point and osmotic pressure will not increase as much as expected.

If a cell is placed in a solution, the concentration of salts in the solution will affect the cell. If the solution is hypertonic, the concentration of salts is higher than the concentration of salts in the cell. If the solution is hypotonic, then the concentration of salts is lower than the concentration of salts in the cell.

Ionic compounds dissociate in solution. The dissociation results in an increase in the number of particles in the solution, i.e., one NaCl "molecule" will dissociate creating two ions; Na$^+$ and Cl$^-$. Colligative properties depend upon the concentration of particles so any compound that dissociates into multiple particles will have pronounced effects on colligative properties.

The van't Hoff factor is the ratio of the value for a colligative property as actually measured to that value of the colligative property that is expected in the complete absence of any solute dissociation. A van't Hoff factor of one is expected for all nondissociating molecular solutes. A van’t Hoff factor greater than one indicates a dissociation of the solute. A van’t Hoff factor less than one indicates association of the solute. If the van’t Hoff factor is 0.5, then this indicates the formation of dimers.

The solute that dissolves to produce the greater number of ions, Na$_2$CO$_3$, gives the solution with the larger boiling point elevation and, thus, the higher boiling point.

(a) suspension
(b) solution or a suspension. It depends on the type of apple juice
(c) solution
(d) solution
(e) suspension. The oils from the lemon are not soluble on water.
(f) suspension

The laser can be used to test light scattering by the liquid.

Sodium stearate is a soap and will stabilize a water in oil emulsion. The hydrophilic head (carboxylate group) of the molecule attaches to water and the hydrophobic tail (non-polar section) of the molecule attaches to the oil molecules.
13.40 A colloidal particle of BaSO₄ in solution has a charged surface due to the adsorption of excess positive or negative ions on its surface as it grows. This ionic atmosphere has a defined volume. Consider a colloidal particle that has an excess of Ba²⁺ ions adsorbed on its surface. The colloidal particle will be repelled by other positively charged colloidal particles when they come within each others ionic boundary. The addition of an electrolyte, such as an acid or salt, to this solution decreases the volume of the ionic boundary. The negative ions of the electrolyte will attach to a positively charged colloidal particle reducing its overall charge. As a result, colloidal particles can come closer together and thus form larger aggregates. Also, heating the solution can be used to increase the kinetic energy of the particles allowing them to collide with enough energy to overcome repulsion and thus increasing particle size.

13.41 A micelle is a cluster of surfactant molecules dispersed in a colloidal suspension. Micelles form due to the structure of surfactant molecules. They have a hydrophilic head and a hydrophobic tail. Surfactant molecules aggregate in a manner that will minimize contact between the hydrophobic end with water molecules. This results in a spherical structure.

Review Problems

13.42 This is to be very much like that shown in Figure 13.5:
The lattice energy is defined as: K⁺(g) + Cl⁻(g) → KCl(s) ΔH° = -715 kJ mol⁻¹

(a) KCl(s) → K⁺(g) + Cl⁻(g), ΔH° = +715 kJ mol⁻¹
(b) K⁺(g) + Cl⁻(g) → K⁺(aq) + Cl⁻(aq), ΔH° = -686 kJ mol⁻¹
KCl(s) → K⁺(aq) + Cl⁻(aq), ΔH° = +29 kJ mol⁻¹

13.43 ΔHsoln = ΔHlattice energy + ΔHhydration
ΔHlattice energy = ΔHsoln - ΔHhydration
ΔHlattice energy = -56 kJ mol⁻¹ - (-894 kJ mol⁻¹) = 838 kJ mol⁻¹

13.44 \[ \frac{C_1}{P_1} = \frac{C_2}{P_2} \]
\[ C_2 = \frac{C_1 \times P_2}{P_1} = \frac{0.025 \text{ g L}^{-1} \times 1.5 \text{ atm}}{1.0 \text{ atm}} = 0.038 \text{ g/L.} \]

13.45 We can compare the solubility that is actually observed with the predicted solubility based on Henry's Law. If the actual and the predicted solubilities are the same, we conclude that the gas obeys Henry's Law. We proceed as in Review Problem 12.40:
\[ \frac{C_1}{P_1} = \frac{C_2}{P_2} \]
\[ C_2 = \frac{C_1 \times P_2}{P_1} = \frac{0.018 \text{ g L}^{-1} \times 620 \text{ torr}}{740 \text{ torr}} = 0.015 \text{ g/L.} \]
The calculated value of C₂ is the same as the observed value, and we conclude that over this pressure range, nitrogen does obey Henry's Law.

13.46 \[ \frac{C_1}{P_1} = \frac{C_2}{P_2} \]
\[ C_2 = \frac{C_1 \times P_2}{P_1} = \frac{0.010 \text{ g L}^{-1} \times 2.0 \text{ atm}}{1.0 \text{ atm}} = 0.020 \text{ g L}^{-1} \]

13.47 \[ C_{\text{gas}} = k_H \times P_{\text{gas}} \]
Chapter 13

\[ C_{O_2} = 0.0039 \ \text{g O}_2/100 \ \text{mL solution} \quad P_{O_2} = 1.0 \ \text{atm} \]

\[ k_H = \frac{3.9 \times 10^{-5} \ \text{g mL}^{-1} \text{ atm}^{-1}}{1.0 \ \text{atm}} = 3.9 \times 10^{-5} \ \text{g mL}^{-1} \text{ atm}^{-1} \]

13.48 One liter of solution has a mass of:

\[ g \ \text{solution} = 1 \ \text{L solution} \left( \frac{1000 \ \text{mL solution}}{1 \ \text{L solution}} \right) \left( \frac{1.07 \ \text{g solution}}{1 \ \text{mL solution}} \right) = 1,070 \ \text{g} \]

According to the given molarity, it contains 3.000 mol NaCl. This has a mass of:

\[ g \ \text{NaCl} = 3.000 \ \text{mol NaCl} \left( \frac{58.45 \ \text{g NaCl}}{1 \ \text{mol NaCl}} \right) = 175.4 \ \text{g NaCl} \]

Thus, the mass of water in 1 L solution must be:

\[ 1,070 \ \text{g} - 175.4 \ \text{g} = 895 \ \text{g water} \]

\[ m = \left( \frac{3.000 \ \text{mol NaCl}}{0.895 \ \text{kg solvent}} \right) = 3.35 \ m \]

13.49 Since the density of the solution is 1.00 g mL\(^{-1}\), the molarity and molality are the same

\[ \text{molality of the acetic acid solution} = \left( \frac{0.143 \ \text{mol CH}_2\text{O}_2\text{H}}{1 \ \text{L soln}} \right) \left( \frac{1 \ \text{L soln}}{1000 \ \text{mL soln}} \right) \left( \frac{1 \ \text{mL soln}}{1.00 \ \text{g soln}} \right) \]

\[ \times \left( \frac{1000 \ \text{g soln}}{1 \ \text{kg soln}} \right) = 0.143 \ m \]

13.50 24.0 g glucose \( \div \) 180 g/mol = 0.133 mol glucose

molality = 0.133 mol glucose/1.00 kg solvent = 0.133 \( m \)

mole fraction = moles glucose/total moles

moles glucose = 0.133

\[ \text{moles H}_2\text{O} = (1.00 \times 10^3 \ \text{g H}_2\text{O}) \left( \frac{1 \ \text{mole H}_2\text{O}}{18.01 \ \text{g H}_2\text{O}} \right) = 55.5 \ \text{mol H}_2\text{O} \]

\[ \chi_{\text{glucose}} = \frac{0.133}{55.5 + 0.133} = 2.39 \times 10^{-3} \]

mass % = \( \frac{24.0 \ \text{g glucose}}{1000 \ \text{g H}_2\text{O} + 24 \ \text{g glucose}} \times 100\% = 2.34\% \)

13.51 mol of NaCl = 11.5 g NaCl \( \left( \frac{1 \ \text{mol NaCl}}{58.44 \ \text{g NaCl}} \right) = 0.197 \ \text{mol NaCl} \)

molality = 0.197 mol NaCl/1.00 kg H\(_2\)O = 0.197 molal

mass % = 11.5/1011.5 \times 100\% = 1.14\%

mole % = \( \frac{\text{mol NaCl}}{\text{total mol}} \times 100\% \)

moles NaCl = 0.197 mol

mole % = \( \frac{0.197 \ \text{mol NaCl}}{55.5 \ \text{mol H}_2\text{O} + 0.197 \ \text{mol NaCl}} \times 100\% = 0.354\% \)
Since the density of water is 1.00 g/mL, the volume of 1 kg is 1 L. Thus, the molarity is: 0.197 mol/1.00 L = 0.197 M. A solvent must have a density close to 1 g/mL for this to happen. Also, the volume of the solvent must not change appreciably on addition of the solute.

13.52 We need to know the mole amounts of both components of the mixture. It is convenient to work from an amount of solution that contains 1.25 mol of ethyl alcohol and, therefore, 1.00 kg of solvent. Convert the number of moles into mass amounts as follows:

For CH₃CH₂OH:

\[
\frac{\text{g ethanol}}{\text{g solution}} = \frac{1.25 \text{ mol ethanol}}{1 \text{ kg water}} \left(\frac{46.08 \text{ g ethanol}}{1 \text{ mol ethanol}}\right) \left(\frac{1 \text{ kg water}}{1000 \text{ g water}}\right) = 57.6 \text{ g ethanol/1000 g water}
\]

Mass % ethanol = (mass ethanol/(total solution mass) × 100% = 5.45%

13.53 If we have 100.0 g of the solution, then 19.5 g is NaCl and the remainder, 80.5 g, is water. We need to know the number of moles of NaCl and the number of kg of water:

\[
\text{mol NaCl} = \left(\frac{19.5 \text{ g NaCl}}{58.44 \text{ g NaCl}}\right) = 0.334 \text{ mol NaCl};
\]

\[
\text{kg H}_2\text{O} = \left(\frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}}\right) = 8.05 \times 10^{-2} \text{ kg H}_2\text{O}.
\]

Molality = \[
\frac{0.334 \text{ mol NaCl}}{8.05 \times 10^{-2} \text{ kg H}_2\text{O}} = 4.14 \text{ m NaCl}.
\]

13.54 If we assume 100 g of solution we have 7.50 g NH₃ and 92.50 g H₂O.

\[
\text{mol NH}_3 = \left(\frac{7.50 \text{ g NH}_3}{17.03 \text{ g NH}_3}\right) = 0.440 \text{ mol NH}_3,
\]

\[
\text{kg H}_2\text{O} = \left(\frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g}}\right) = 0.0925 \text{ kg H}_2\text{O}.
\]

\[
\text{m} = \frac{0.440 \text{ moles NH}_3}{0.0925 \text{ kg H}_2\text{O}} = 4.76 \text{ m}.
\]

\[
\text{mol H}_2\text{O} = \left(\frac{92.5 \text{ g H}_2\text{O}}{18.0 \text{ g H}_2\text{O}}\right) = 5.14 \text{ mol H}_2\text{O}.
\]

Mole percent = \[
\frac{0.440 \text{ mol NH}_3}{0.440 \text{ mol NH}_3 + 5.14 \text{ mol H}_2\text{O}} \times 100% = 7.89%.
\]

13.55 Assume 1 mole total.

\[
\text{g C}_3\text{H}_8\text{O} = 0.250 \text{ mol}\left(\frac{60.10 \text{ g}}{1 \text{ mol}}\right) = 15.02 \text{ g C}_3\text{H}_8\text{O}
\]

\[
\text{mol H}_2\text{O} = 0.750 \text{ mol}
\]

\[
\text{g H}_2\text{O} = (0.750 \text{ mol H}_2\text{O})\left(\frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}}\right) = 13.5 \text{ g H}_2\text{O}.
\]

\[
\text{C}_3\text{H}_8\text{O mass \%} = \frac{15.02 \text{ g}}{13.5 \text{ g} + 15.02 \text{ g}} \times 100% = 52.7%.
\]
13.56 If we choose, for convenience, an amount of solution that contains 1 kg of solvent, then it also contains 0.363 moles of NaNO₃. The number of moles of solvent is:
\[
\text{mol H}_2\text{O} = \frac{(1000 \text{ g H}_2\text{O})}{18.02 \text{ g/mol}} = 55.6 \text{ mol H}_2\text{O}
\]
Now, convert the number of moles to a number of grams: for NaNO₃, \(0.363 \text{ mol} \times 85.0 \text{ g/mol} = 30.9 \text{ g}\); for H₂O, 1000 g was assumed and the percent (w/w) values are:
\[
\% \text{ NaNO}_3 = \frac{30.9 \text{ g}}{1030.9 \text{ g}} \times 100\% = 3.00\%
\]
\[
\% \text{ H}_2\text{O} = \frac{1000 \text{ g}}{1030.9 \text{ g}} \times 100 = 97.0\%
\]
To determine the molar concentration of NaNO₃ assume 1 kg of solvent which would then contain 0.363 mole of NaNO₃ or 30.9 g NaNO₃. The total mass of the solution would be 1000 g + 30.9 g = 1031 g of solution. Now, the ratio of moles of solute to grams of solution is \(0.363 \text{ mol NaNO}_3/1031 \text{ g solution}\).
\[
\text{Molarity} = \frac{0.363 \text{ mol NaNO}_3}{1031 \text{ g solution}} = 0.359 \text{ M NaNO}_3
\]
13.57 (a) If the sample is 1.89 mol% H₂SO₄, then an amount of the solution that contains 1.89 mol of H₂SO₄ also contains \((100 - 1.89) = 98.11\) mol water. We can calculate the molality if we know the number of moles of H₂SO₄ and the number of kg of solvent. The latter is determined as follows:
\[
\text{kg H}_2\text{O} = \frac{98.11 \text{ mol H}_2\text{O} \times 18.02 \text{ g/mol} \times 1 \text{ kg/1000 g}}{1.768 \text{ kg H}_2\text{O}} = 1.07 \text{ m H}_2\text{SO}_4
\]
(b) The mass of H₂SO₄ in the above sample is: 1.89 mol × 98.1 g/mol = 185 g H₂SO₄. The total mass of the solution is then equal to \([185 \text{ g} + (1.768 \times 10^3 \text{ g})] = 1.953 \times 10^3 \text{ g}\), and the % (w/w) values are: for H₂SO₄,
\[
\text{Mass % H}_2\text{SO}_4 = \frac{185 \text{ g H}_2\text{SO}_4}{1953 \text{ g total mass}} \times 100 = 9.47\% ;
\]
for H₂O,
\[
\text{Mass % water} = \frac{1768 \text{ g H}_2\text{O}}{1953 \text{ g total mass}} \times 100 = 90.53\% .
\]
(c) If we have on hand 100 mL (0.100 L) of this solution, it will have a mass that can be determined using its known density:
\[
\text{mass solution} = (100.0 \text{ mL}) \left( \frac{1.0645 \text{ g solution}}{1 \text{ mL solution}} \right) = 106.4 \text{ g of solution}.
\]
Since this solution has 9.49% (w/w) H₂SO₄, the mass of H₂SO₄ in 0.100 L of the solution is: Mass H₂SO₄ = 106.4 g × 0.0949 = 10.1 g H₂SO₄.
The number of moles of H₂SO₄ is thus:
\[
\text{mol H}_2\text{SO}_4 = \left( \frac{10.1 \text{ g H}_2\text{SO}_4}{98.1 \text{ g H}_2\text{SO}_4} \right) = 0.103 \text{ mol H}_2\text{SO}_4.
\]
The molarity is the number of moles of H₂SO₄ divided by the volume of solution:
\[
\text{Molarity} = \left( \frac{0.103 \text{ mol H}_2\text{SO}_4}{0.100 \text{ L solution}} \right) = 1.03 \text{ M H}_2\text{SO}_4.
\]
Chapter 13

13.58 \[ P_{\text{solution}} = P^o_{\text{solvent}} \times \chi_{\text{solvent}} \]

We need to determine \( \chi_{\text{solvent}} \):

\[
\text{mol glucose} = (65.0 \text{ g}) \left( \frac{1 \text{ mol}}{180.2 \text{ g}} \right) = 0.361 \text{ mol glucose}
\]

\[
\text{mol H}_2\text{O} = (150 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 8.32 \text{ mol H}_2\text{O}
\]

The total number of moles is thus: 8.32 mol + 0.361 mol = 8.69 mol and the mole fraction of the solvent is:

\[
\chi_{\text{solvent}} = \left( \frac{8.32 \text{ mol solvent}}{8.69 \text{ mol solution}} \right) = 0.957.
\]

Therefore,

\[ P_{\text{solution}} = 23.8 \text{ torr} \times 0.957 = 22.8 \text{ torr}. \]

13.59 In 100 g of the mixture we have the following mole amounts:

- 65.0 g H\(_2\)O \(\times\) 18.02 g/mol = 3.61 mol H\(_2\)O
- 35 g C\(_2\)H\(_4\)O\(_2\) \(\times\) 62.07 g/mol = 0.564 mol ethylene glycol

\[
\chi_{\text{H}_2\text{O}} = \frac{3.61 \text{ mol H}_2\text{O} + 0.564 \text{ mol ethylene glycol}}{3.61 \text{ mol H}_2\text{O}} = 0.865
\]

\[ P_{\text{solution}} = P^o_{\text{solvent}} \times \chi_{\text{solvent}} = 17.5 \text{ torr} \times 0.865 = 15.1 \text{ torr}. \]

13.60 \[ P_{\text{benzene}} = \chi_{\text{benzene}} \times P^o_{\text{benzene}} \]

\[ P_{\text{toluene}} = \chi_{\text{toluene}} \times P^o_{\text{toluene}} \]

\[ P_{\text{Tot}} = P_{\text{benzene}} + P_{\text{toluene}} \]

\[
\text{mol benzene} = (35.0 \text{ g}) \left( \frac{1 \text{ mol}}{78.11 \text{ g}} \right) = 0.448 \text{ mol benzene}
\]

\[
\text{mol toluene} = (65.0 \text{ g}) \left( \frac{1 \text{ mol}}{92.14 \text{ g}} \right) = 0.705 \text{ mol toluene}
\]

\[
\chi_{\text{benzene}} = \frac{0.448}{0.448 + 0.705} = 0.389
\]

\[
\chi_{\text{toluene}} = \frac{0.705}{0.448 + 0.705} = 0.611
\]

\[ P_{\text{benzene}} = (0.389)(93.4 \text{ torr}) = 36.3 \text{ torr} \]

\[ P_{\text{toluene}} = (0.611)(26.9 \text{ torr}) = 16.4 \text{ torr} \]

\[ P_{\text{Tot}} = 36.3 \text{ torr} + 16.4 \text{ torr} = 52.7 \text{ torr} \]

13.61 Assume 50 g each of pentane:

\[
\text{mol pentane} = (50 \text{ g pentane}) \left( \frac{1 \text{ mol pentane}}{72.15 \text{ g pentane}} \right) = 0.693 \text{ mol}
\]

\[
\text{mol heptane} = (50 \text{ g heptane}) \left( \frac{1 \text{ mol heptane}}{100.21 \text{ g heptane}} \right) = 0.499 \text{ mol heptane}
\]

\[
\chi_{\text{pentane}} = \frac{0.693}{0.693 + 0.499} = 0.581
\]

\[
\chi_{\text{heptane}} = \frac{0.499}{0.693 + 0.499} = 0.419
\]

\[ P_{\text{pentane}} = \chi_{\text{pentane}} \times P^o_{\text{pentane}} = 0.581 \times 422 \text{ torr} = 245 \text{ torr} \]
Chapter 13

\[ P_{\text{heptane}} = \chi_{\text{heptane}} \times P^\circ_{\text{heptane}} = 0.419 \times 36 \text{ torr} = 15.1 \text{ torr} \]
\[ P_{\text{Total}} = P_{\text{pentane}} + P_{\text{heptane}} = (245 + 15.1) \text{ torr} = 260 \text{ torr} \]

13.62 The following relationships are to be established: \( P_{\text{Total}} = 96 \text{ torr} = (P^\circ_{\text{benzene}} \times \chi_{\text{benzene}}) + (P^\circ_{\text{toluene}} \times \chi_{\text{toluene}}) \).

The relationship between the two mole fractions is: \( \chi_{\text{benzene}} = 1 - \chi_{\text{toluene}} \), since the sum of the two mole fractions is one. Substituting this expression for \( \chi_{\text{benzene}} \) into the first equation gives:

\[ 96 \text{ torr} = [P^\circ_{\text{benzene}} \times (1 - \chi_{\text{toluene}})] + [P^\circ_{\text{toluene}} \times \chi_{\text{toluene}}]. \]

Solving for \( \chi_{\text{toluene}} \) we get: \( 120 \times \chi_{\text{toluene}} = 84 \), \( \chi_{\text{toluene}} = 0.70 \) and \( \chi_{\text{benzene}} = 0.30 \). The mole % values are to be 70 mol% toluene and 30 mol% benzene.

13.63 \( \chi_{\text{CH}_3\text{OH}} = P/P^\circ = 145 \text{ torr}/164 \text{ torr} = 0.884 \)

\[ \text{mol } \text{CH}_3\text{OH} = (105 \text{ g}) \left( \frac{1 \text{ mol } \text{CH}_3\text{OH}}{32.0 \text{ g } \text{CH}_3\text{OH}} \right) = 3.28 \text{ mol } \text{CH}_3\text{OH} \]

\[ 0.884 = \frac{3.28 \text{ mol } \text{CH}_3\text{OH}}{3.28 \text{ mol } \text{CH}_3\text{OH} + x \text{ mol } \text{C}_3\text{H}_5(\text{OH})_3} \]

\[ 3.28 \text{ mol } \text{CH}_3\text{OH} + x \text{ mol } \text{C}_3\text{H}_5(\text{OH})_3 = \frac{3.28 \text{ mol } \text{CH}_3\text{OH}}{0.884} \]

\[ x = \frac{3.28 \text{ mol}}{0.884} - 3.28 \text{ mol} = 0.430 \text{ mol} \]

\[ g \text{ C}_3\text{H}_5(\text{OH})_3 = (0.430 \text{ mol}) \left( \frac{92.1 \text{ g}}{1 \text{ mol}} \right) = 39.6 \text{ g } \text{C}_3\text{H}_5(\text{OH})_3 \]

13.64 (a) \( \chi_{\text{solvent}} = \frac{P}{P^\circ} = \frac{511 \text{ torr}}{526 \text{ torr}} = 0.971 \)

\( \chi_{\text{solvent}} = 1 - \chi_{\text{solvent}} = 0.029 \)

(b) We know 0.971 = \( \frac{1 \text{ mol}}{1 \text{ mol} + x \text{ mol}} \) \( x = 2.99 \times 10^{-2} \) moles

(c) molar mass = \( \frac{8.3 \text{ g}}{2.99 \times 10^{-2} \text{ mol}} = 278 \text{ g/mol} \)

13.65 \( P_{\text{solvent}} = \chi_{\text{solvent}} \times P^\circ_{\text{solvent}} \)

\[ 336.0 \text{ torr} = \chi_{\text{solvent}} \times 400.0 \text{ torr} \]

\( \chi_{\text{solvent}} = 0.8400 \)

\( \chi_{\text{solvent}} = 1 - 0.8400 = 0.1600 \)

The number of moles of solvent is: \( 33.25 \text{ g} \div 109.0 \text{ g/mol} = 0.3050 \text{ mol} \) and the following expression for mole fraction of solvent can be solved to determine the number of moles of solute

We know 0.8400 = 0.3050 mol/(0.3050 mol + x) \( x = 5.81 \times 10^{-2} \) moles

Molar mass = 18.26 g/5.81 \( \times 10^{-2} \text{ mol} = 314.3 \text{ g/mol} \)

13.66 \( \Delta T_f = K_m \)

\( m = \Delta T_f/K_m = 3.00 \text{ °C}/1.86 \text{ °C kg/mol} = 1.61 \text{ mol/kg} \)

\[ \text{kg} = (125 \text{ g}) \left( \frac{1 \text{ kg}}{1000 \text{ g}} \right) = 0.125 \text{ kg} \]
Chapter 13

mol = \left( \frac{1.61 \text{ mol}}{1 \text{ kg}} \right) (0.125 \text{ kg}) = 0.201 \text{ mol}

g = (0.201 \text{ mol}) \left( \frac{342.3 \text{ g}}{1 \text{ mol}} \right) = 68.9 \text{ g}

13.67 \quad \Delta T_b = K_b m

T_c = (270 \text{ °F} - 32 \text{ °F}) \left( \frac{5 \text{ °C}}{9 \text{ °C}} \right) = 132 \text{ °C}

(132 \text{ °C} - 100 \text{ °C}) = (0.51 \text{ °C kg/mol}) \times m

m = 63 m

\chi_{\text{sucrose}} = \left( \frac{63 \text{ mol sucrose}}{1000 \text{ g water} \left( \frac{1 \text{ mol H}_2\text{O}}{18.01 \text{ g H}_2\text{O}} \right) + 63 \text{ mol sucrose}} \right) = 0.53

13.68 \quad \Delta T_b = K_b m

\Delta T_b = 81.7 \text{ °C} - 80.2 \text{ °C} = 1.5 \text{ °C}

m = \Delta T_b/K_b = 1.5 \text{ °C}/2.53 \text{ °C kg/mol} = 0.593 \text{ mol/kg}

mol solute = (0.593 \text{ mol/kg})(0.100 \text{ kg}) = 0.0593 \text{ mol}

molar mass = (14 \text{ g})/0.0593 \text{ mol} = 240 \text{ g/mol}

13.69 \quad \Delta T = (5.45 - 3.45) = 2.00 \text{ °C} = K_f \times m = 5.07 \text{ °C kg mol}^{-1} \times m

m = 0.394 \text{ mol solute/kg solvent}

0.394 \text{ mol/kg benzene} \times 0.200 \text{ kg benzene} = 0.0788 \text{ mol solute and the molecular mass is: 12.00 g/0.0788 mol} = 152 \text{ g/mol}

13.70 \quad \Delta T_f = K_m

m = \Delta T_f/K_f = 0.307 \text{ °C}/5.07 \text{ °C kg/mol} = 0.0606 \text{ mol/kg}

mol = \left( \frac{0.0606 \text{ mol}}{1 \text{ kg}} \right) (0.5 \text{ kg}) = 0.0303 \text{ mol}

molar mass = \frac{3.84 \text{ g}}{0.0303 \text{ mol}} = 127 \text{ g/mol}

The empirical formula has a mass of 64.1 g/mol. So the molecular formula is C_8H_4N_2.

13.71 \quad (a) \quad \text{For convenience we choose to work with 100 g of the compound, and then to convert the mass amounts of each element found in this compound into mole amounts:}

\text{for C, 42.86 g ÷ 12.01 g/mol} = 3.569 \text{ mol C}

\text{for H, 2.40 g ÷ 1.01 g/mol} = 2.38 \text{ mol H}

\text{for N, 16.67 g ÷ 14.01 g/mol} = 1.190 \text{ mol N}

\text{for O, 38.07 g ÷ 16.00 g/mol} = 2.379 \text{ mol O}

\text{The relative mole amounts that represent the empirical formula are determined by dividing the above mole amounts each by the smallest mole amount:}

\text{for C: 3.569 mol ÷ 1.190 mol} = 2.999

\text{for H: 2.379 mol ÷ 1.190 mol} = 1.99

\text{for N: 1.190 mol ÷ 1.190 mol} = 1.000

\text{for O: 2.379 mol ÷ 1.190 mol} = 1.999

\text{and the empirical formula is C_8H_4N_2.}

(b) \quad \Delta T_b = 1.84 \text{ °C} = K_b \times m = 2.53 \text{ °C kg mol}^{-1} \times m

m = 0.727 \text{ mol solute/kg benzene.}
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The number of moles of solute is: 0.727 mol/kg benzene \times 0.045 kg benzene = 0.0327 mol, and the formula mass is: 5.5g/0.0327 mol = 168 g/mol. Since the mass of the empirical unit is 84, the molecular formula must be twice the empirical formula, namely \( C_6H_4N_2O_4 \).

13.72 (a) If the equation is correct, the units on both sides of the equation should be g/mol. The units on the right side of this equation are:

\[
\frac{(g) \times (L \text{ atm mol}^{-1} \text{ K}^{-1}) \times (K)}{L \times \text{ atm}} = \text{g/mol}
\]

which is correct.

(b) \( \Pi = MRT = (n/V)RT \), \( n = \Pi V/RT \)
This means that we can calculate the number of moles of solute in one L of solution, as follows:

\[
n = \frac{0.021 \text{ torr} \times 1 \text{ atm/760 torr}}{0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 298 \text{ K}} \times 1.0 \text{ L} = 1.1 \times 10^{-6} \text{ mol}
\]

The molecular mass is the mass in 1 L divided by the number of moles in 1 L:

\[
2.0 \text{ g/1.1} \times 10^{-6} \text{ mol} = 1.8 \times 10^6 \text{ g/mol}
\]

13.73 \( \Pi = MRT = (n/V)RT \), \( n = \Pi V/RT \)

\[
n = \frac{3.74 \text{ torr} \times 1 \text{ atm/760 torr}}{0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}} \times 1 \text{ L} = 2.00 \times 10^{-4} \text{ mol}
\]

\[
\text{molar mass} = \frac{0.400 \text{ g}}{2.00 \times 10^{-4} \text{ mol}} = 2.00 \times 10^3 \text{ g/mol}
\]

13.74 The equation for the vapor pressure is:

\[
P_{\text{solution}} = P_0^{H_2O} \times \chi^{H_2O}
\]

Where \( P_0^{H_2O} \) is 17.5 torr. To calculate the vapor pressure we need to find the mole fraction of water first.

\[
\chi^{H_2O} = \text{moles H}_2\text{O}/(\text{moles H}_2\text{O} + \text{moles NaCl})
\]

Calculate the moles of NaCl in 23.0 g

\[
\text{mol NaCl} = 23.0 \text{ g NaCl/58.44 g NaCl} = 0.394 \text{ moles NaCl}
\]

When NaCl dissolves in water, Na\(^+\) and Cl\(^-\) are formed. So, for every mole of NaCl that dissolves, two moles of ions are formed. For this solution, the number of moles of ions is 0.788.

The number of moles of solute (water) is:

\[
\text{mol H}_2\text{O} = 100 \text{ g H}_2\text{O}/18.02 \text{ g H}_2\text{O} = 5.55 \text{ moles H}_2\text{O}
\]

Calculate the mole fraction as

\[
\chi^{H_2O} = \frac{\text{moles H}_2\text{O}}{(\text{moles H}_2\text{O} + \text{moles NaCl})} = \frac{5.55 \text{ mol}}{(5.55 \text{ mol} + 0.788 \text{ mol})} = 0.876
\]

The vapor pressure is then \( P_{\text{solution}} = P_0^{H_2O} \times \chi^{H_2O} = 17.5 \text{ torr} \times 0.876 = 15.3 \text{ torr} \)

13.75 \( \chi^{H_2O} = \frac{P}{P^0} = \frac{38.7 \text{ torr}}{42.2 \text{ torr}} = 0.917 \)
\[ \text{mol H}_2\text{O} = (150 \text{ mL}) \left( \frac{1 \text{ g}}{1 \text{ mL}} \right) \left( \frac{1 \text{ mol}}{18.02 \text{ g}} \right) = 8.32 \text{ mol} \]

\[ 0.917 = \frac{8.32 \text{ mol}}{8.32 \text{ mol} + x \text{ mol}} \]

\[ X = 0.753 \text{ mole} \]

Since the van’t Hoff factor for AlCl\(_3\) is 4, we need:

\[ 0.753 \text{ mol} / 4 = 0.188 \text{ mol AlCl}_3 \]

\[ \text{g AlCl}_3 = (0.188 \text{ mol}) \left( \frac{133.3 \text{ g}}{1 \text{ mol}} \right) = 25.1 \text{ g AlCl}_3 \]

13.76 Assume 100 mL of solution, that is, 2.0 g NaCl and 0.100 L of solution:

\[ \Pi = M \text{RT} \]

\[ (2.0 \text{ g NaCl}) \left( \frac{1 \text{ mol NaCl}}{58.45 \text{ g NaCl}} \right) = 0.34 \text{ M} \]

For every NaCl there are two ions produces so \( M = 0.68 \text{ M} \)

\[ \Pi = (0.68 \text{ M})(0.0821 \text{ L atm/mol K})(298 \text{ K}) \left( \frac{760 \text{ torr}}{1 \text{ atm}} \right) = 1.3 \times 10^4 \text{ torr} \]

13.77 \[ \Pi = M \text{RT} \]

For each ion, multiply the concentration by 24.47 L atm/mol

<table>
<thead>
<tr>
<th>Ion</th>
<th>Molality (mol/L)</th>
<th>( \Pi ) (atm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl</td>
<td>0.566</td>
<td>13.9</td>
</tr>
<tr>
<td>Na(^+)</td>
<td>0.486</td>
<td>11.9</td>
</tr>
<tr>
<td>Mg(^{2+})</td>
<td>0.055</td>
<td>1.35</td>
</tr>
<tr>
<td>SO(_4^{2-})</td>
<td>0.029</td>
<td>0.710</td>
</tr>
<tr>
<td>Ca(^{2+})</td>
<td>0.011</td>
<td>0.269</td>
</tr>
<tr>
<td>K(^+)</td>
<td>0.011</td>
<td>0.269</td>
</tr>
<tr>
<td>HCO(_3^{-})</td>
<td>0.002</td>
<td>0.0489</td>
</tr>
</tbody>
</table>

Adding these together we get \( \Pi = 28.4 \text{ atm} \). Thus, a pressure greater than 28.4 atm is needed to desalinate seawater by reverse osmosis.

13.78 CaCl\(_2\) \( \rightarrow \) Ca\(^{2+}\) + 2Cl\(^-\); \quad \text{van’t Hoff factor, } i = 3

\[ \Delta T_f = i \times K_f \times m = (3)(1.86 \text{ °C m}^{-1})(0.20 \text{ m}) = 1.1 \text{ °C} \]

The freezing point is \(-1.1 \text{ °C} \).

13.79 If we assume that mercury(I) nitrate has the formula HgNO\(_3\), we predict a freezing point of \(-0.37 \text{ °C} \),

\[ \Delta T_f = i \times K_f \times m = 2 \times 1.86 \text{ °C m}^{-1} \times 0.10, \text{ m} = 0.37 \text{ °C} \]

However, the observed freezing point depression is lower than this. So, assume that the correct formula of the compound is Hg\(_2\)(NO\(_3\))\(_2\) where the mercury ion is dimeric and divalent, i.e., Hg\(_2^{2+}\).

Assuming that the person who prepared the solution thought it was HgNO\(_3\) when they calculated the molality, the concentration of the solution based on the correct formula for the compound would be recalculated correctly as 0.050 m since the true formula mass is twice that assumed, and \( \Delta T_f = i \times K_f \times m = 3 \times 1.86 \text{ °C m}^{-1} \times 0.050 \text{ m} = 0.28 \text{ °C} \).

Therefore, the dissociation produces three ions and the equation is:

Hg\(_2\)(NO\(_3\))\(_2\) \( \rightarrow \) Hg\(_2^{2+}\) + 2NO\(_3^{-}\)
13.80 Any electrolyte such as NiSO\(_4\), that dissociated to give 2 ions, if fully dissociated should have a van't Hoff factor of 2.

13.81 K\(_2\)SO\(_4\) \(\rightarrow\) 2K\(^+\) + SO\(_4^{2-}\), the van’t Hoff factor is expected to be three.

13.82 \(\Delta T_f = i \times K_f \times m\)

\(i = \Delta T_f / K_f \times m = 0.415°C/(1.86 °C m\(^{-1}\))(0.118 m) = 1.89\)

13.83 To solve this problem, we need to assume the density of the solution is 1 g/mL. From problem 13.82 we know that most of the LiCl has dissociated. As a result, the affect of the dissociated ions will increase the osmotic pressure.

\[\Pi = MRT\]

If we consider the dissociation

\[\Pi = iMRT\]

\[\Pi = (1.89)(0.118 \text{ mol L}^{-1})(0.0821 \text{ L atm/mol K})(283 \text{ K})(760 \text{ torr}/1 \text{ atm}) = 3.94 \times 10^3 \text{ torr}\]

**Additional Exercises**

13.84 The partial pressure of N\(_2\) in air is:

\[P_{N_2} = 1.00 \text{ atm}(78 \text{ mol%}) = 0.78 \text{ atm}\]

Therefore, according to Henry’s Law, the amount of N\(_2\) dissolved per liter of blood at 1.00 atm is:

\[(1 \text{ L})(0.015 \text{ g/L})(0.78/1.00) = 0.012 \text{ g N}_2\]

\[0.012 \text{ g N}_2 (1 \text{ mol N}_2/28.0 \text{ g N}_2) = 0.00043 \text{ mol N}_2\]

The amount of N\(_2\) dissolved per liter of blood at 4.00 atm would be four times that, or: 0.0017 mol N\(_2\)

The amount of nitrogen released per liter of blood upon quickly surfacing is the difference between the two, or (0.0017 mol – 0.00043 mol) = 0.0013 mol N\(_2\). The volume of that gas at 1.00 atm and 37 °C would be given by the ideal gas law:

\[PV = nRT\]

\[V = nRT/P\]

\[V = (0.0013 \text{ mol N}_2)(0.0821 \text{ L atm/mol K})(273+37) \text{ K}/1 \text{ atm}\]

\[V = 0.033 \text{ L} = 33 \text{ mL N}_2 \text{ per liter of blood}\]

13.85 Let A = CCl\(_4\) and B = unknown

\[P_{Tot} = P_A + P_B\]

\[P_{Tot} = \chi_A P_A^0 + \chi_B P_B^0\]

We also know that \(\chi_A + \chi_B = 1\)

So, \(P_{Tot} = (1 - \chi_B)P_A^0 + \chi_B P_B^0\)

\[P_{Tot} - P_A^0 = \chi_B (P_B^0 - P_A^0)\]

\[\chi_B = \left(\frac{P_{Tot} - P_A^0}{P_B^0 - P_A^0}\right) = \left(\frac{137 \text{ torr} - 143 \text{ torr}}{85 \text{ torr} - 143 \text{ torr}}\right) = 0.103\]

\[\chi_A = 0.897\]

\[\chi_{CCl_4} = \frac{\text{mol CCl}_4}{\text{mol CCl}_4 + \text{mol unknown}},\] rearranging, we get
moles unknown = \text{mol CCl}_4 \left( \frac{1}{7 \text{CCl}_4} - 1 \right)

\text{mol CCl}_4 = (400 \text{ g CCl}_4 \text{ mol}^{-1}) \left( \frac{1 \text{ mol CCl}_4}{153.8 \text{ g CCl}_4} \right) = 2.60 \text{ mol CCl}_4

\text{mol unknown} = 2.60 \text{ mol CCl}_4 \left( \frac{1}{0.897} - 1 \right) = 0.299 \text{ mol unknown}

\text{molar mass} = \frac{43.3 \text{ g}}{0.299 \text{ moles}} = 145 \text{ g/mol}

13.86 \quad \Pi = MRT = (0.0100 \text{ mol/L})(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K}) = 0.245 \text{ atm}

0.245 \text{ atm} \times 760 \text{ torr/atm} = 186 \text{ torr}.

13.87 (a) Since $-45^\circ F$ is equal to $-43^\circ C$, the following expression applies: $\Delta T = k_m n$, so

$40^\circ C = (1.86^\circ C \text{ kg mol}^{-1}) \times m$, \quad m = 43/1.86 \text{ mol/kg} = 23 \text{ molal}$

Therefore, 22 moles must be added to 1 kg of water.

(b) $\text{mL} = (22 \text{ moles}) \left( \frac{62.1 \text{ g}}{1 \text{ mol}} \right) \left( \frac{1.00 \text{ mL}}{1.11 \text{ g}} \right) = 1.3 \times 10^3 \text{ mL}$

(c) There are 946 mL in one quart. Thus, for 1 qt of water we are to have 946 mL, and the required number of quarts of ethylene glycol is:

$$\frac{\text{qt C}_2\text{H}_6\text{O}_2}{\text{1 qt H}_2\text{O}} = \left( \frac{1.3 \times 10^3 \text{ mL C}_2\text{H}_6\text{O}_2}{1000 \text{ g H}_2\text{O}} \right) \left( \frac{1 \text{ g H}_2\text{O}}{1 \text{ mL H}_2\text{O}} \right) \left( \frac{946 \text{ mL H}_2\text{O}}{1 \text{ qt H}_2\text{O}} \right) \left( \frac{1 \text{ qt C}_2\text{H}_6\text{O}_2}{946 \text{ mL C}_2\text{H}_6\text{O}_2} \right)$$

$= 1.3 \text{ qt C}_2\text{H}_6\text{O}_2$

The proper ratio of ethylene glycol to water is 1.3 qt to 1 qt.

13.88 (a) The height difference is proportional to the osmotic pressure, therefore $\Pi$ may be calculated by converting the height difference to the height of a mercury column in mm, which is equal to the pressure in torr (1 mm Hg = 1 torr):

$h_{\text{Hg}} = h_{\text{solution}} \times (d_{\text{solute}}/d_{\text{Hg}}) = (12.6 \text{ mm}) \times (1.00 \text{ g/mL}/13.6 \text{ g/mL}) = 0.926 \text{ mm Hg}$

$P = 0.926 \text{ torr}$

(b) $\Pi = MRT$

$M = \Pi / RT = \left( \frac{0.926 \text{ torr}}{760 \text{ torr}} \right) \left( \frac{1 \text{ atm}}{60 \text{ torr}} \right) = 4.98 \times 10^{-5} \text{ M}$

(c) Since this is a dilute solution and the solute does not dissociate, we can assume that the molarity and molality are equivalent. So,

$\Delta T_i = k_m = (1.86 \text{ C m}^{-1}) (4.98 \times 10^{-5} \text{ m})$

$= 9.26 \times 10^{-5} \text{ C}$

Freezing point will be $-9.26 \times 10^{-5} \text{ C}$

(d) The magnitude of the temperature change is too small to measure.

13.89 (a) Since the molarity of the solution is 4.613 mol/L, then one L of this solution contains:

$4.613 \text{ mol} \times 46.07 \text{ g/mol} = 212.5 \text{ g C}_2\text{H}_5\text{OH}$. 
The mass of the total 1 L of solution is:

\[1000 \text{ mL} \times 0.9677 \text{ g/mL} = 967.7 \text{ g.}\]

The mass of water is thus 967.7 g – 212.5 g = 755.2 g H₂O, and the molality is:

\[4.613 \text{ mol C}_2\text{H}_5\text{OH}/0.7552 \text{ kg H}_2\text{O} = 6.108 \text{ m}.\]

(b) \( \% \text{ (w/w)} \text{ C}_2\text{H}_5\text{OH} = (212.5 \text{ g}/967.7 \text{ g}) \times 100\% = 21.96\% \)

13.90 (a) \( \Delta T_b = k_b m = (0.51 \text{ °C m}^{-1})(1.00 \text{ m}) \)

\[= 0.51 \text{ °C} \]

\( \Delta T_b = 100 \text{ °C} + 0.51 \text{ °C} = 100.51 \text{ °C} \)

(b) \( \Delta T_b = i k_b m = (4)(0.51 \text{ °C m}^{-1})(1.00 \text{ m}) \)

\[= 2.04 \text{ °C} \]

\( \Delta T_b = 100 \text{ °C} + 2.04 \text{ °C} = 102.04 \text{ °C} \)

(c) \(i = 0.183 \text{ °C}/0.51 \text{ °C} = 0.36\)

13.91 Colligative properties are based on the number of particles in solution. In each of the two cases we are told to assume complete dissociation. Thus, one formula unit of AlCl₃ contributes 4 ions and one formula unit of Na₂SO₄ contributes 3 ions to the solution.

The total molality, \( m_T \), of the solution is the sum of the molalities of the two salts.

\( m_T = m_1(\text{AlCl}_3) + m_2(\text{Na}_2\text{SO}_4) \)

The molality of AlCl₃ is \( 4/7 \) of \( m_T \) and the molality of Na₂SO₄ is \( 3/7 \) of \( m_T \).

\[\Delta T_f = i \times K_f \times m_T\]

\[2.65 \text{ °C} = 7 \times 1.86 \frac{\text{°C}}{\text{molal}} \times m_T\]

\( m_T = 0.2035 \text{ m} \)

Since the total molality is small we can assume that molarity and molality are approximately the same value.

\( M(\text{AlCl}_3) = \frac{4}{7} \times 0.2035 \text{ m} = 0.116 \text{ M} \)

\( M(\text{Na}_2\text{SO}_4) = \frac{3}{7} \times 0.2035 \text{ m} = 0.0872 \text{ M} \)

Multi-Concept Problems

13.92 From the boiling point elevation we can determine the total molality of the solution.

\[\Delta T_f = i \times K_f \times m_T\]

\[4.6 \text{ °C} = 5 \times 0.51 \frac{\text{°C}}{\text{molal}} \times m_T\]

\( m_T = 1.804 \text{ m} \)

The molality of the KCl solution is:

\[\frac{2}{5} \times 1.804 \text{ m} = 0.7216 \text{ m} \]
The molality of the Fe(NO\textsubscript{3})\textsubscript{2} is:

\[
\frac{3}{5} \times 1.804 \text{ m} = 1.0824 \text{ m}
\]

From the titration we can determine the molarity of the iron(II) nitrate.

\[
6\text{Fe}^{2+} + \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 6\text{Fe}^{3+} + 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

\[
36.3 \text{ mL} \times \frac{0.220 \text{ mmol Cr}_2\text{O}_7^{2-}}{\text{mL}} \times \frac{6 \text{ mmol Fe}^{2+}}{\text{mmol Cr}_2\text{O}_7^{2-}} = 47.916 \text{ mmol Fe}^{2+}
\]

Or 47.916 \times 10^{-3} \text{ mol Fe}^{2+}

We can determine the volume of solvent from the calculated molality of the iron(II) solution and the solution’s density.

First, we need to determine the mass of solvent containing 4.7916 \times 10^{-2} \text{ mol of iron salt}.

\[
\frac{1 \text{ kg solvent}}{1.0824 \text{ mol Fe(NO}_3)_2} \times 0.04792 \text{ mol} = 4.427 \times 10^{-2} \text{ kg solvent}
\]

Or 44.27 \text{ g of solvent}

Mass of Fe(NO\textsubscript{3})\textsubscript{2} = 0.04792 \text{ mol} \times \frac{179.857 \text{ g}}{\text{mol Fe(NO}_3)_2} = 8.62 \text{ g}

Mass of solution = 8.62 \text{ g Fe(NO}_3)_2 + 44.27 \text{ g H}_2\text{O} = 52.89 \text{ g}

Volume of solution = 52.89 \text{ g} \times \frac{1 \text{ mL}}{1.032 \text{ g}} = 51.25 \text{ mL}

\[
\text{M(Fe(NO}_3)_2} = \frac{47.916 \text{ mmol Fe(NO}_3)_2}{51.25 \text{ mL}} = 0.93 \text{ M}
\]

The molarity of the KCl solution would be:

\[
\text{M(KCl)} = \frac{0.7216 \text{ mol KCl}}{\text{kg solvent}} \times 4.427 \times 10^{-2} \text{ kg solvent} \times \frac{1}{5.125 \times 10^{-2} \text{ L}} = 0.62 \text{ M}
\]

13.93 Osmotic pressure is given by \( \Pi = iMRT \)

Since this is a mixture, the molarity, \( M \), is the sum of the individual solutes in solution. BaSO\textsubscript{4} will precipitate out of solution. Therefore, we need to determine how much Ba\textsuperscript{2+} and SO\textsubscript{4}\textsuperscript{2-} will be removed from the solution.

Before reaction:

mmol Ba\textsuperscript{2+} = 25.00 \text{ mL} \times 0.200 \text{ M} = 5 \text{ mmol}

mmol NO\textsubscript{3}\textsuperscript{−} = 2 \times 5 \text{ mmol} = 10 \text{ mmol}
mmol SO$_4^{2-}$ = 14.00 mL x 0.250 M = 3.5 mmol
mmol K$^+$ = 2 x 3.5 mmol = 7 mmol
SO$_4^{2-}$ is the limiting reagent.

After reaction we have:
1.5 mmol of Ba$^{2+}$
10 mmol of NO$_3^-$
7 mmol of K$^+$

iM = \frac{1.5 \text{ mmol} + 10 \text{ mmol} + 7 \text{ mmol}}{39.00 \text{ mL}} = 0.474 \text{ M}

\Pi = 0.474 \text{ M} \times 0.0821 \text{ L atm} \cdot \text{mol}^{-1} \times 298 \text{ K} = 11.6 \text{ atm}

11.6 \text{ atm} \times 760 \text{ torr atm}^{-1} = 8.82 \times 10^3 \text{ torr}

13.94 There are two isomers of C$_2$H$_6$O, ethanol, and dimethyl ether.

Shown below are the abbreviated structural forms of the two isomers.

\begin{center}
\begin{tikzpicture}
  \node at (0,0) {OH};
  \node at (1,0) {O};
\end{tikzpicture}
\end{center}

ethanol \hspace{1cm} dimethyl ether

(a) Both are expected to be liquids at 25 °C. Ethanol possesses a dipole moment, 1.68 Debye, it can hydrogen bond with itself, and has significant London forces. Dimethyl ether has a dipole, 1.30 Debye, and possesses significant London forces.

(b) Both should be soluble in water since they are both polar and can interact with polar water molecules. Also, ethanol will hydrogen bond with water. Neither should be significantly soluble in non-polar pentane.

(c) Ethanol will form hydrogen bonds due the presence of the OH group and unshared electrons on the oxygen atom. Dimethyl ether cannot hydrogen bond as it does not have a hydrogen attached to the oxygen atom.

(d) Both are non-electrolytes since they do not significantly ionize in water. The hydrogen on the OH group is not an ionizable hydrogen.

13.95 The reaction is given below.

\[3\text{Sn}^{2+} + \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ \rightarrow 3\text{Sn}^{4+} + 2\text{Cr}^{3+} + 7\text{H}_2\text{O}\]

\[0.155 \text{ L} \times \frac{0.650 \text{ mol Sn}^{2+}}{\text{L}} = 0.1008 \text{ mol Sn}^{2+}\]

\[0.1008 \text{ mol Sn}^{2+} \times \frac{1 \text{ mol Cr}_2\text{O}_7^{2-}}{3 \text{ mol Sn}^{2+}} = 0.223 \times \frac{\text{mol Cr}_2\text{O}_7^{2-}}{\text{L}} \times V\]

\[V = 0.151 \text{ L} \text{ or } 151 \text{ mL} \text{ required}\]
To determine the osmotic pressure we first need to determine the molarity of the resulting solution.

\[
M(\text{Sn}^{2+}) = \frac{0.1008 \text{ mol Sn}^{2+}}{(0.155 \text{ L} + 0.151 \text{ L})} = 0.329 \text{ M}
\]

\[
M(\text{Cr}_2\text{O}_7^{2-}) = \frac{0.1008 \text{ mol Sn}^{2+} \times 1 \text{ mol Cr}_2\text{O}_7^{2-}}{3 \text{ mol Sn}^{2+}} = 0.110 \text{ M}
\]

Using these molarities and the number of particles created for each solution when the reaction is complete, we can determine the osmotic pressure. Be sure to include the spectator ions when determining the value of “i.”

\[
\Pi = iMRT
\]

\[
\Pi = \left[ 4 \times 0.110 \frac{\text{mol Cr}_2\text{O}_7^{2-}}{\text{L}} + 3 \times 0.329 \frac{\text{mol Sn}^{2+}}{\text{L}} \right] \times 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}
\]

\[
\Pi = 34.9 \text{ atm}
\]

In this solution to the problem a value of \(i = 4\) was used for the dichromate since 2 moles of \(\text{Cr}^{3+}\) are produced by the reaction. Alternately, you could determine the molar concentration of the \(\text{Cr}^{3+}\) and then use the molar concentration of the dichromate to determine the molar concentration of the \(\text{K}^+\) ions.

13.96 (a) The formula masses are \(\text{Na}_2\text{Cr}_2\text{O}_7\cdot2\text{H}_2\text{O}: 298 \text{ g/mol}, \text{C}_3\text{H}_8\text{O}: 60.1 \text{ g/mol}, \text{and C}_3\text{H}_6\text{O}: 58.1 \text{ g/mol.} \)

\[
g \text{Na}_2\text{Cr}_2\text{O}_7\cdot2\text{H}_2\text{O} = (21.4 \text{ g C}_3\text{H}_8\text{O}) \left( \frac{1 \text{ mol C}_3\text{H}_8\text{O}}{60.1 \text{ g C}_3\text{H}_8\text{O}} \right) \left( \frac{1 \text{ mol Na}_2\text{Cr}_2\text{O}_7\cdot2\text{H}_2\text{O}}{3 \text{ mol C}_3\text{H}_8\text{O}} \right)
\]

\[
x \left( \frac{298 \text{ g Na}_2\text{Cr}_2\text{O}_7\cdot2\text{H}_2\text{O}}{1 \text{ mol Na}_2\text{Cr}_2\text{O}_7\cdot2\text{H}_2\text{O}} \right) = 35.4 \text{ g Na}_2\text{Cr}_2\text{O}_7\cdot2\text{H}_2\text{O}
\]

(b) The theoretical yield is:

\[
g \text{C}_3\text{H}_6\text{O} = (21.4 \text{ g C}_3\text{H}_6\text{O}) \left( \frac{1 \text{ mol C}_3\text{H}_6\text{O}}{60.1 \text{ g C}_3\text{H}_6\text{O}} \right) \left( \frac{3 \text{ mol C}_3\text{H}_6\text{O}}{1 \text{ mol C}_3\text{H}_6\text{O}} \right) = 20.7 \text{ g C}_3\text{H}_6\text{O}
\]

The percent yield is therefore: \(12.4/20.7 \times 100\% = 59.9\%\)

(c) First, we determine the number of grams of C, H, and O that are found in the products, and then the \% by mass of C, H, and O that were present in the sample that was analyzed by combustion, i.e. the by-product:

\[
g \text{ C} = (22.368 \times 10^{-3} \text{ g CO}_2) \left( \frac{12.011 \text{ g C}}{44.010 \text{ g CO}_2} \right) = 6.1046 \times 10^{-3} \text{ g C}
\]

and the \% C is: \((6.1046 \times 10^{-3} \text{ g}/8.654 \times 10^{-3} \text{ g})\times 100\% = 70.54\%\)

\[
g \text{ H} = (10.655 \times 10^{-3} \text{ g H}_2\text{O}) \left( \frac{2.0159 \text{ g H}}{18.015 \text{ g H}_2\text{O}} \right) = 1.1923 \times 10^{-3} \text{ g H}
\]

and the \% H is: \((1.1923 \times 10^{-3} \text{ g H}/8.654 \times 10^{-3} \text{ g})\times 100\% = 13.78\%\)

For O, the mass is the total mass minus that of C and H in the sample that was analyzed:

\[
8.654 \times 10^{-3} \text{ g total} - (6.1046 \times 10^{-3} \text{ g C} + 1.1923 \times 10^{-3} \text{ g H}) = 1.357 \times 10^{-3} \text{ g O}
\]
and the % O is: \((1.357 \times 10^{-3} \text{ g})/(8.654 \times 10^{-3} \text{ g}) \times 100\% = 15.68\% \text{ O}\).

Alternatively, we could have determined the amount of oxygen by using the mass % values, realizing that the sum of the mass percent values should be 100.

Next, we convert these mass amounts for C, H, and O into mole amounts by dividing the amount of each element by the atomic mass of each element:

For C, \(6.1046 \times 10^{-3} \text{ g C} / 12.011 \text{ g/mol} = 0.50825 \times 10^{-3} \text{ mol C}\)

For H, \(1.1923 \times 10^{-3} \text{ g H} / 1.0079 \text{ g/mol} = 1.1829 \times 10^{-3} \text{ mol H}\)

For O, \(1.357 \times 10^{-3} \text{ g O} / 16.00 \text{ g/mol} = 0.08481 \times 10^{-3} \text{ mol O}\)

Lastly, these are converted to relative mole amounts by dividing each of the above mole amounts by the smallest of the three (We can ignore the \(10^{-3}\) term since it is common to all three components):

For C, \(0.50825 \text{ mol}/0.08481 \text{ mol} = 5.993\)

For H, \(1.1829 \text{ mol}/0.08481 \text{ mol} = 13.95\)

For O, \(0.08481 \text{ mol}/0.08481 \text{ mol} = 1.000\)

and the empirical formula is given by this ratio of relative mole amounts, namely \(C_6H_{14}O\).

(d) \(\Delta T = K_f m, \quad (5.45 \, ^\circ\text{C} - 4.87 \, ^\circ\text{C}) = (5.07 \, ^\circ\text{C}/\text{m}) \times m, \quad m = 0.11 \text{ molal, and there are 0.11 moles of solute dissolved in each kg of solvent. Thus, the number of moles of solute that have been used here is:}\)

\[0.11 \text{ mol/kg} \times 0.1150 \text{ kg} = 1.3 \times 10^{-2} \text{ mol solute.}\]

The formula mass is thus: \(1.338 \text{ g/0.013 mol} = 102 \text{ g/mol. Since the empirical formula has this same mass, we conclude that the molecular formula is the same as the empirical formula, i.e. } C_6H_{14}O.\)