

CHAPTER 13 THE PROPERTIES OF MIXTURES: SOLUTIONS AND COLLOIDS

- 13.1 A heterogeneous mixture has two or more phases, thus seawater has both dissolved and suspended particles. The composition of the seawater is different in various places where a sample may be obtained.
- 13.2 When a salt such as NaCl dissolves, ion-dipole forces cause the ions to separate, and many water molecules cluster around each of them in hydration shells. Ion-dipole forces hold the first shell. Additional shells are held by hydrogen bonding to inner shells.
- 13.3 In $\text{CH}_3(\text{CH}_2)_n\text{COOH}$, as n increases, the hydrophobic (CH) portion of the carboxylic acid increases and the hydrophilic part of the molecule stays the same, with a resulting decrease in water solubility.
- 13.4 **Sodium stearate** would be a more effective soap because the hydrocarbon chain in the stearate ion is longer than the chain in the acetate ion. A soap forms suspended particles called micelles with the polar end of the soap interacting with the water solvent molecules and the nonpolar ends forming a nonpolar environment inside the micelle. Oils dissolve in the nonpolar portion of the micelle. Thus, a better solvent for the oils in dirt is a more nonpolar substance. The long hydrocarbon chain in the stearate ion is better at dissolving oils in the micelle than the shorter hydrocarbon chain in the acetate ion.
- 13.5 Hexane and methanol, as gases, are free from any intermolecular forces and can simply intermix with each other. As liquids, hexane is a non-polar molecule, whereas methanol is a polar molecule. "Like dissolves like."
- 13.6 Hydrogen chloride (HCl) gas is actually reacting with the solvent (water) and thus shows a higher solubility than propane (C_3H_8) gas, which does not react.
- 13.7 a) A more concentrated solution will have more solute dissolved in the solvent. Potassium nitrate, KNO_3 , is an ionic compound and therefore soluble in a polar solvent like water. Potassium nitrate is not soluble in the nonpolar solvent CCl_4 . Because potassium nitrate dissolves to a greater extent in water, **KNO_3 in H_2O** will result in the more concentrated solution.
- 13.8 b) Stearic acid in CCl_4 . Stearic acid will not dissolve in water. It is non-polar while water is very polar. Stearic acid will dissolve in carbon tetrachloride, as both are non-polar.
- 13.9 To identify the strongest type of intermolecular force, check the formula of the solute and identify the forces that could occur. Then look at the formula for the solvent and determine if the forces identified for the solute would occur with the solvent. The strongest force is ion-dipole followed by dipole-dipole (including H bonds). Next in strength is ion-induced dipole force and then dipole-induced dipole force. The weakest intermolecular interactions are dispersion forces.
- a) **Ion-dipole forces** are the strongest intermolecular forces in the solution of the ionic substance cesium chloride in polar water.
- b) **Hydrogen bonding** (type of dipole-dipole force) is the strongest intermolecular force in the solution of polar propanone (or acetone) in polar water.
- c) **Dipole-induced dipole forces** are the strongest forces between the polar methanol and nonpolar carbon tetrachloride.
- 13.10 a) **metallic bonding**
b) **dipole-dipole**
c) **dipole-induced dipole**

- 13.11 a) **Hydrogen bonding** occurs between the H atom on water and the lone electron pair on the O atom in dimethyl ether (CH_3OCH_3). However, none of the hydrogen atoms on dimethyl ether participates in hydrogen bonding because the C–H bond does not have sufficient polarity.
 b) The dipole in water induces a dipole on the $\text{Ne}(g)$ atom, so **dipole-induced dipole** interactions are the strongest intermolecular forces in this solution.
 c) Nitrogen gas and butane are both nonpolar substances, so **dispersion forces** are the principal attractive forces.
- 13.12 a) **dispersion forces**
 b) **hydrogen bonding**
 c) **dispersion forces**
- 13.13 $\text{CH}_3\text{CH}_2\text{OCH}_2\text{CH}_3$ is polar with dipole-dipole interactions as the dominant intermolecular forces. Examine the solutes to determine which has intermolecular forces more similar to those for the diethyl ether. This solute is the one that would be more soluble.
 a) **HCl** would be more soluble since it is a covalent compound with dipole-dipole forces, whereas NaCl is an ionic solid. Dipole-dipole forces between HCl and diethyl ether are more similar to the dipole forces in diethyl ether than the ion-dipole forces between NaCl and diethyl ether.
 b) **CH_3CHO** (acetaldehyde) would be more soluble. The dominant interactions in H_2O are hydrogen bonding, a stronger type of dipole-dipole force. The dominant interactions in CH_3CHO are dipole-dipole. The solute-solvent interactions between CH_3CHO and diethyl ether are more similar to the solvent intermolecular forces than the forces between H_2O and diethyl ether.
 c) **$\text{CH}_3\text{CH}_2\text{MgBr}$** would be more soluble. $\text{CH}_3\text{CH}_2\text{MgBr}$ has a polar end ($-\text{MgBr}$) and a nonpolar end (CH_3CH_2-), whereas MgBr_2 is an ionic compound. The nonpolar end of $\text{CH}_3\text{CH}_2\text{MgBr}$ and diethyl ether would interact with dispersion forces, while the polar end of $\text{CH}_3\text{CH}_2\text{MgBr}$ and the dipole in diethyl ether would interact with dipole-dipole forces. Recall, that if the polarity continues to increase, the bond will eventually become ionic. There is a continuous sequence from nonpolar covalent to ionic.
- 13.14 a) **$\text{CH}_3\text{CH}_2-\text{O}-\text{CH}_3(g)$** , due to its smaller size (smaller molar mass).
 b) **CH_2Cl_2** , because it is more polar than CCl_4 .
 c) **Tetrahydropyran** is more water soluble due to hydrogen bonding between the oxygen atom and water molecules.
- 13.15 **No**, river water is a heterogeneous mixture, with its composition changing from one segment to another.
- 13.16 Gluconic acid is a very polar molecule because it has $-\text{OH}$ groups attached to every carbon. The abundance of $-\text{OH}$ bonds allows gluconic acid to participate in extensive H–bonding with water, hence its great solubility in water. On the other hand, caproic acid has a 5–carbon, nonpolar, hydrophobic (“water hating”) tail that does not easily dissolve in water. The dispersion forces in the nonpolar tail are more similar to the dispersion forces in hexane, hence its greater solubility in hexane.
- 13.17 There may be a disulfide linkage (a covalent disulfide bridge) between the ends of two cysteine side chains that bring together parts of the chain. There may be salt links between ions $-\text{COO}^-$ and $-\text{NH}_3^+$ groups. There may be hydrogen bonding between the $\text{C}=\text{O}$ of one peptide bond and the N–H of another.
- 13.18 The nitrogen bases hydrogen bond to their complimentary bases. The flat, N-containing bases stack above each other, which allow extensive interaction through dispersion forces. The exterior negatively charged sugar-phosphate chains form ion-dipole and hydrogen bonds to the aqueous surroundings, but this is of minor importance to the structure.
- 13.19 The more carbon and hydrogen atoms present, the more soluble the substance is in non-polar oil droplets. Therefore, sodium propanoate is not as effective a soap as sodium stearate with the longer hydrocarbon chain.
- 13.20 Dispersion forces are present between the nonpolar portions of the molecules within the bilayer. Polar groups are present to hydrogen bond or to form dipole-dipole interactions with the aqueous surroundings.

- 13.21 In soluble proteins, polar groups are found on the exterior and nonpolar groups on the interior. In proteins embedded in a membrane, the exterior of the protein that lies within the bilayer consists of nonpolar amino acid side chains, whereas the portion lying outside the bilayer has polar side chains.
- 13.22 While an individual hydrogen bond is not too strong, there are very large numbers of hydrogen bonds present in the wood. The strength of wood comes from the large number of hydrogen bonds, and to a lesser degree from the numerous dispersion interactions.
- 13.23 Amino acids with side chains that may be ionic are necessary. Two examples are **lysine** and **glutamic acid**.
- 13.24 The $\Delta H_{\text{solvent}}$ and ΔH_{mix} components of the heat of solution combined together represent the enthalpy change during solvation, the process of surrounding a solute particle with solvent particles. Solution in water is often called hydration.
- 13.25 For a general solvent, the energy changes needed to separate solvent into particles ($\Delta H_{\text{solvent}}$), and that needed to mix the solvent and solute particles (ΔH_{mix}) would be combined to obtain $\Delta H_{\text{solution}}$.
- 13.26 a) Charge density is the ratio of the ion's charge to its volume. An ion's charge and size affect its charge density.
 b) $- < + < 2- < 3+$
 c) The higher the charge density, the more negative is $\Delta H_{\text{hydration}}$. $\Delta H_{\text{hydration}}$ increases with increasing charge and decreases with increasing size.
- 13.27 The solution cycle for ionic compounds in water consists of two enthalpy terms: the lattice energy, and the combined heats of hydration of the cation and anion.

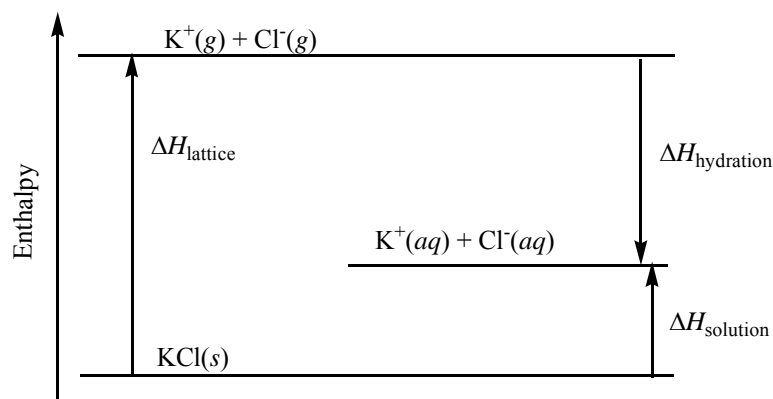
$$\Delta H_{\text{soln}} = \Delta H_{\text{lattice}} + \Delta H_{\text{hyd of ions}}$$

For a heat of solution to be zero (or very small)

$$\Delta H_{\text{lattice}} \approx \Delta H_{\text{hydration of ions}}, \text{ and they would have to have opposite signs.}$$

- 13.28 a) **Endothermic**
 b) The lattice energy term is much **larger** than the combined ionic heats of hydration.
 c) The increase in entropy outweighs the increase in enthalpy, so ammonium chloride dissolves..
- 13.29 This compound would be very soluble in water. A large exothermic value in $\Delta H_{\text{solution}}$ (enthalpy of solution) means that the solution has a much lower energy state than the isolated solute and solvent particles, so the system tends to the formation of the solution. Entropy that accompanies dissolution always favors solution formation. Entropy becomes important when explaining why solids with an endothermic $\Delta H_{\text{solution}}$ (and higher energy state) are still soluble in water.

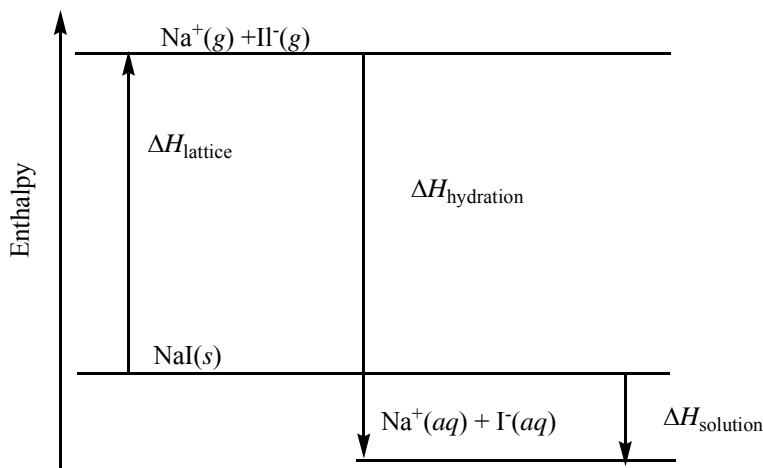
13.30



$$\Delta H_{\text{solution}} > 0 \text{ (endothermic)}$$

Lattice energy values are always positive as energy is required to separate the ions from each other. Hydration energy values are always negative as energy is released when intermolecular forces between ions and water form. Since the heat of solution for KCl is endothermic, the lattice energy must be greater than the hydration energy for an overall input of energy.

13.31



Lattice energy values are always positive as energy is required to separate the ions from each other. Hydration energy values are always negative as energy is released when intermolecular forces between ions and water form. Since the heat of solution for NaI is exothermic, the negative hydration energy must be greater than the positive lattice energy.

- 13.32 Charge density is the ratio of an ion's charge (regardless of sign) to its volume. An ion's volume is related to its radius. For ions whose charges have the same sign (+ or -), ion size decreases as a group in the periodic table is ascended and as you proceed from left to right in the periodic table. Charge density increases with increasing charge and increases with decreasing size.
- Both ions have a +1 charge, but the volume of Na^+ is smaller, so it has the greater charge density.
 - Sr^{2+} has a greater ionic charge and a smaller size (because it has a greater Z_{eff}), so it has the greater charge density.
 - Na^+ has a smaller ion volume than Cl^- , so it has the greater charge density.
 - O^{2-} has a greater ionic charge and similar ion volume, so it has the greater charge density.
 - OH^- has a smaller ion volume than SH^- (O is smaller than S), so it has the greater charge density.
 - Mg^{2+} has the higher charge density because it has a smaller ion volume.
 - Mg^{2+} has the higher charge density because it has both a smaller ion volume and greater charge.
 - CO_3^{2-} has the higher charge density because it has both a smaller ion volume and greater charge.
- 13.33
- I^- has a smaller charge density (larger ion volume) than Br^- .
 - Ca^{2+} is less than Sc^{3+} , due to its smaller ion charge.
 - Br^- is less than K^+ , due to its larger ion volume.
 - Cl^- is less than S^{2-} , due to its smaller ion charge.
 - Sc^{3+} is less than Al^{3+} , due to its larger ion volume.
 - ClO_4^- has the smaller charge density due to its smaller ion charge.
 - Fe^{2+} has the smaller charge density due to its smaller ion charge.
 - K^+ has the smaller charge density due to its smaller ion charge.

- 13.34 The ion with the greater charge density will have the larger $\Delta H_{\text{hydration}}$.
- Na^+ would have a larger $\Delta H_{\text{hydration}}$ than Cs^+ since its charge density is greater than that of Cs^+ .
 - Sr^{2+} would have a larger $\Delta H_{\text{hydration}}$ than Rb^+ .
 - Na^+ would have a larger $\Delta H_{\text{hydration}}$ than Cl^- .
 - O^{2-} would have a larger $\Delta H_{\text{hydration}}$ than F^- .
 - OH^- would have a larger $\Delta H_{\text{hydration}}$ than SH^- .
 - Mg^{2+} would have a larger $\Delta H_{\text{hydration}}$ than Ba^{2+} .
 - Mg^{2+} would have a larger $\Delta H_{\text{hydration}}$ than Na^+ .
 - CO_3^{2-} would have a larger $\Delta H_{\text{hydration}}$ than NO_3^- .
- 13.35 a) I^- b) Ca^{2+} c) Br^- d) Cl^- e) Sc^{3+} f) ClO_4^- g) Fe^{2+} h) K^+
- 13.36 a) The two ions in potassium bromate are K^+ and BrO_3^- . The heat of solution for ionic compounds is $\Delta H_{\text{soln}} = \Delta H_{\text{lattice}} + \Delta H_{\text{hydr of the ions}}$. Therefore, the combined heats of hydration for the ions is $(\Delta H_{\text{soln}} - \Delta H_{\text{lattice}})$ or $41.1 \text{ kJ/mol} - 745 \text{ kJ/mol} = -703.9 = \mathbf{-704 \text{ kJ/mol}}$.
- b) K^+ ion contributes more to the heat of hydration because it has a smaller size and, therefore, a greater charge density.
- 13.37 a) $\Delta H_{\text{hydration of ions}} = \Delta H_{\text{solution}} - \Delta H_{\text{lattice}}$
 $\Delta H_{\text{hydration of ions}} = 17.3 \text{ kJ/mol} - 763 \text{ kJ/mol}$
 $\Delta H_{\text{hyd}} = -745.7 = \mathbf{-746 \text{ kJ/mol}}$
- b) It is the Na^+ due to its smaller size (larger charge density).
- 13.38 Entropy increases as the possible states for a system increases.
- Entropy **increases** as the gasoline is burned. Gaseous products at a higher temperature form.
 - Entropy **decreases** as the gold is separated from the ore. Pure gold has only the arrangement of gold atoms next to gold atoms, while the ore mixture has a greater number of possible arrangements among the components of the mixture.
 - Entropy **increases** as a solute dissolves in the solvent.
- 13.39 a) Entropy **increases**
 b) Entropy **decreases**
 c) Entropy **increases**
- 13.40 $\Delta H_{\text{solution}} = \Delta H_{\text{lattice}} + \Delta H_{\text{hydration of ions}}$
 $\Delta H_{\text{soln}} = 822 \text{ kJ/mol} + (-799 \text{ kJ/mol})$
 $\Delta H_{\text{soln}} = \mathbf{23 \text{ kJ/mol}}$
- 13.41 Add a pinch of the solid solute to each solution. A saturated solution contains the maximum amount of dissolved solute at a particular temperature. When additional solute is added to this solution, it will remain undissolved. An unsaturated solution contains less than the maximum amount of dissolved solute and so will dissolve added solute. A supersaturated solution is unstable and addition of a "seed" crystal of solute causes the excess solute to crystallize immediately, leaving behind a saturated solution.
- 13.42 $\text{KMnO}_4(s) + \text{H}_2\text{O}(l) + \text{heat} \rightarrow \text{KMnO}_4(aq)$
 Prepare a mixture of more than 6.4 g KMnO_4 / 100 g H_2O and heat it until the solid completely dissolves. Then carefully cool it, without disturbing it or shaking it, back to 20°C . If no crystals form, you would then have a supersaturated solution.
- 13.43 An increase in temperature produces an increase in kinetic energy; the solute molecules overcome the weak intermolecular forces, which results in a decrease in solubility of any gas in water. In nearly all cases, gases dissolve exothermically ($\Delta H_{\text{soln}} < 0$).

- 13.44 a) Increasing pressure for a gas **increases** the solubility of the gas according to Henry's law.
 b) Increasing the volume of a gas causes a decrease in its pressure (Boyle's Law), which **decreases** the solubility of the gas.

- 13.45 a) **increase** b) **stay the same**

- 13.46 a) Solubility for a gas is calculated from Henry's law: $S_{\text{gas}} = k_{\text{H}} \times P_{\text{gas}}$. S_{gas} is expressed as mol/L, so convert moles of O_2 to mass of O_2 using the molar mass.

$$S_{\text{gas}} = k_{\text{H}} \times P_{\text{gas}}$$

$$S_{\text{gas}} = \left(1.28 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{atm}}\right)(1.00 \text{ atm}) = 1.28 \times 10^{-3} \text{ mol/L}$$

$$\left(\frac{1.28 \times 10^{-3} \text{ mol O}_2}{\text{L}}\right)\left(\frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2}\right)(2.50 \text{ L}) = 0.1024 = \mathbf{0.102 \text{ g O}_2}$$

- b) The amount of gas that will dissolve in a given volume decreases proportionately with the partial pressure of the gas, so

$$S_{\text{gas}} = \left(1.28 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{atm}}\right)(0.209 \text{ atm}) = 2.6752 \times 10^{-4} \text{ mol/L}$$

$$\left(\frac{2.6752 \times 10^{-4} \text{ mol O}_2}{\text{L}}\right)\left(\frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2}\right)(2.50 \text{ L}) = 0.0214016 = \mathbf{0.0214 \text{ g O}_2}$$

- 13.47 Solubility = $\left(1.5 \times 10^{-3} \frac{\text{mol}}{\text{L}\cdot\text{atm}}\right)(1.0 \text{ atm})\left(\frac{0.93\%}{100\%}\right) = 1.395 \times 10^{-5} = \mathbf{1.4 \times 10^{-5} \text{ mol/L}}$

- 13.48 The solution is **saturated**.

- 13.49 Solubility for a gas is calculated from Henry's law: $S_{\text{gas}} = k_{\text{H}} \times P_{\text{gas}}$.
 $S_{\text{gas}} = (3.7 \times 10^{-2} \text{ mol/L}\cdot\text{atm})(5.5 \text{ atm}) = 0.2035 = \mathbf{0.20 \text{ mol/L}}$

- 13.50 Solubility of gases increases with increasing partial pressure of the gas, and the goal of these devices is to increase the amount of oxygen dissolving in the bloodstream.

- 13.51 Molarity is defined as the number of moles of solute dissolved in one liter of solution. Molality is defined as the number of moles of solute dissolved in 1000 g (1 kg) of solvent. Molal solutions are prepared by measuring masses of solute and solvent, which are additive and not changed by temperature, so the concentration in molality does not change with temperature.

- 13.52 Refer to the Table 13.5 for the different methods of expressing concentration.

a) **Molarity** and **parts-by-volume** (% w/v or % v/v) include the volume of the solution.

b) **Parts-by-mass** (% w/w) include the mass of solution directly. (Others may involve the mass indirectly.)

c) **Molality** includes the mass of the solvent.

- 13.53 **No**, 21 g solute/kg of solvent would be 21 g solute / 1.021 kg solution.

- 13.54 Converting between molarity and molality involves conversion between volume of solution and mass of solution. Both of these quantities are given so interconversion is possible. To convert to mole fraction requires that the mass of solvent be converted to moles of solvent. Since the identity of the solvent is not given, conversion to mole fraction is not possible if the molar mass is not known.

- 13.55 % w/w, mole fraction, and molality are weight-to-weight relationships that are not affected by changes in temperature. % w/v and molarity are affected by changes in temperature, because the volume is temperature dependant.
- 13.56 Convert the masses to moles and the volumes to liters and use the definition of molarity: $M = \frac{\text{mol of solute}}{V(\text{L}) \text{ of solution}}$
- a) Molarity = $\left(\frac{32.3 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}}{100. \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol C}_{12}\text{H}_{22}\text{O}_{11}}{342.30 \text{ g C}_{12}\text{H}_{22}\text{O}_{11}}\right) = 0.943617 = \mathbf{0.944 M C_{12}H_{22}O_{11}}$
- b) Molarity = $\left(\frac{5.80 \text{ g LiNO}_3}{505 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol LiNO}_3}{68.95 \text{ g LiNO}_3}\right) = 0.166572 = \mathbf{0.167 M LiNO_3}$
- 13.57 a) Molarity = $\left(\frac{0.82 \text{ g C}_2\text{H}_5\text{OH}}{10.5 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.07 \text{ g C}_2\text{H}_5\text{OH}}\right) = 1.69514 = \mathbf{1.7 M C_2H_5OH}$
- b) Molarity = $\left(\frac{1.27 \text{ g NH}_3}{33.5 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right) = 2.2261 = \mathbf{2.23 M NH_3}$
- 13.58 Dilution calculations can be done using $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$
- a) $M_{\text{conc}} = 0.240 \text{ M NaOH}$ $V_{\text{conc}} = 78.0 \text{ mL}$ $M_{\text{dil}} = ?$ $V_{\text{dil}} = 0.250 \text{ L}$
 $M_{\text{dil}} = M_{\text{conc}} V_{\text{conc}} / V_{\text{dil}} = \frac{(0.240 \text{ M})(78.0 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)}{(0.250 \text{ L})} = 0.07488 = \mathbf{0.0749 M}$
- b) $M_{\text{conc}} = 1.2 \text{ M HNO}_3$ $V_{\text{conc}} = 38.5 \text{ mL}$ $M_{\text{dil}} = ?$ $V_{\text{dil}} = 0.130 \text{ L}$
 $M_{\text{dil}} = M_{\text{conc}} V_{\text{conc}} / V_{\text{dil}} = \frac{(1.2 \text{ M})(38.5 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)}{(0.130 \text{ L})} = 0.355385 = \mathbf{0.36 M}$
- 13.59 Dilution calculations can be done using $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$
- a) $M_{\text{conc}} = 6.25 \text{ M HCl}$ $V_{\text{conc}} = 25.5 \text{ mL}$ $M_{\text{dil}} = ?$ $V_{\text{dil}} = 0.500 \text{ L}$
 $M_{\text{dil}} = M_{\text{conc}} V_{\text{conc}} / V_{\text{dil}} = \frac{(6.25 \text{ M})(25.5 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)}{(0.500 \text{ L})} = 0.31875 = \mathbf{0.319 M}$
- b) $M_{\text{conc}} = 2.00 \times 10^{-2} \text{ M KI}$ $V_{\text{conc}} = 8.25 \text{ mL}$ $M_{\text{dil}} = ?$ $V_{\text{dil}} = 12.0 \text{ mL}$
 $M_{\text{dil}} = M_{\text{conc}} V_{\text{conc}} / V_{\text{dil}} = \frac{(2.00 \times 10^{-2} \text{ M})(8.25 \text{ mL})}{(12.0 \text{ mL})} = 0.01375 = \mathbf{0.0138 M}$
- 13.60 a) Find the number of moles KH_2PO_4 needed to make 365 mL of this solution. Convert moles to mass using the molar mass of KH_2PO_4 (Molar mass = 136.09 g/mol)
- $$\text{Mass KH}_2\text{PO}_4 = (365 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{8.55 \times 10^{-2} \text{ mol KH}_2\text{PO}_4}{\text{L}}\right)\left(\frac{136.09 \text{ g KH}_2\text{PO}_4}{1 \text{ mol KH}_2\text{PO}_4}\right)$$
- $$= 4.24703 = 4.25 \text{ g KH}_2\text{PO}_4$$
- Add **4.25 g KH_2PO_4** to enough water to make 365 mL of aqueous solution.
- b) Use the relationship $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$ to find the volume of 1.25 M NaOH needed.
 $M_{\text{conc}} = 1.25 \text{ M NaOH}$ $V_{\text{conc}} = ?$ $M_{\text{dil}} = 0.335 \text{ M NaOH}$ $V_{\text{dil}} = 465 \text{ mL}$
 $V_{\text{conc}} = M_{\text{dil}} V_{\text{dil}} / M_{\text{conc}} = \frac{(0.335 \text{ M})(465 \text{ mL})}{(1.25 \text{ M})} = 124.62 = \mathbf{125 \text{ mL}}$
 Add **125 mL** of 1.25 M NaOH to enough water to make 465 mL of solution.

- 13.61 a) Find the number of moles NaCl needed to make 2.5 L of this solution. Convert moles to mass using the molar mass of NaCl (Molar mass = 58.44 g/mol)

$$\text{Mass NaCl} = (2.5 \text{ L}) \left(\frac{0.65 \text{ mol NaCl}}{\text{L}} \right) \left(\frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} \right) = 94.965 = 95 \text{ g NaCl}$$

Add **95 g NaCl** to enough water to make 2.5 L of aqueous solution.

- b) Use the relationship $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$ to find the volume of 2.1 M urea needed.

$$M_{\text{conc}} = 2.1 \text{ M urea} \quad V_{\text{conc}} = ? \quad M_{\text{dil}} = 0.3 \text{ M urea} \quad V_{\text{dil}} = 15.5 \text{ L}$$

$$V_{\text{conc}} = M_{\text{dil}} V_{\text{dil}} / M_{\text{conc}} = (0.3 \text{ M}) (15.5 \text{ L}) / (2.1 \text{ M}) = 2.21429 = 2 \text{ L}$$

Add **2 L** of 2.1 M urea to enough water to make 15.5 L of solution.

Note because of the uncertainty in the concentration of the dilute urea (0.3 M), only one significant figure is justified in the answer.

- 13.62 a) To find the mass of KBr needed, find the moles of KBr in 1.40 L of a 0.288 M solution and convert to grams using molar mass of KBr.

$$\text{Mass KBr} = (1.40 \text{ L}) \left(\frac{0.288 \text{ mol KBr}}{\text{L}} \right) \left(\frac{119.00 \text{ g KBr}}{1 \text{ mol KBr}} \right) = 47.9808 = 48.0 \text{ g KBr}$$

To make the solution, weigh **48.0 g KBr** and then dilute to 1.40 L with distilled water.

- b) To find the volume of the concentrated solution that will be diluted to 255 mL, use $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$ and solve for V_{conc} .

$$M_{\text{conc}} = 0.264 \text{ M LiNO}_3 \quad V_{\text{conc}} = ? \quad M_{\text{dil}} = 0.0856 \text{ M LiNO}_3 \quad V_{\text{dil}} = 255 \text{ mL}$$

$$V_{\text{conc}} = M_{\text{dil}} V_{\text{dil}} / M_{\text{conc}} = (0.0856 \text{ M}) (255 \text{ mL}) / (0.264 \text{ M}) = 82.68182 = 82.7 \text{ mL}$$

To make the 0.0856 M solution, measure **82.7 mL** of the 0.264 M solution and add distilled water to make a total of 255 mL.

- 13.63 a) To find the mass of $\text{Cr}(\text{NO}_3)_3$ needed, find the moles of $\text{Cr}(\text{NO}_3)_3$ in 57.5 mL of a $1.53 \times 10^{-3} \text{ M}$ solution and convert to grams using molar mass of $\text{Cr}(\text{NO}_3)_3$.

$$\begin{aligned} \text{Mass Cr}(\text{NO}_3)_3 &= (57.5 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{1.53 \times 10^{-3} \text{ mol Cr}(\text{NO}_3)_3}{\text{L}} \right) \left(\frac{238.03 \text{ g Cr}(\text{NO}_3)_3}{1 \text{ mol Cr}(\text{NO}_3)_3} \right) \\ &= 0.020941 = 0.0209 \text{ g Cr}(\text{NO}_3)_3 \end{aligned}$$

To make the solution, weigh **0.0209 g $\text{Cr}(\text{NO}_3)_3$** and then dilute to 57.5 mL with distilled water.

- b) To find the volume of the concentrated solution that will be diluted to $5.8 \times 10^3 \text{ m}^3$ use $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$ and solve for V_{conc} .

$$M_{\text{conc}} = 2.50 \text{ M NH}_4\text{NO}_3 \quad V_{\text{conc}} = ? \quad M_{\text{dil}} = 1.45 \text{ M NH}_4\text{NO}_3 \quad V_{\text{dil}} = 5.8 \times 10^3 \text{ m}^3$$

$$V_{\text{conc}} = M_{\text{dil}} V_{\text{dil}} / M_{\text{conc}} = (1.45 \text{ M}) (5.8 \times 10^3 \text{ m}^3) / (2.50 \text{ M}) = 3.364 \times 10^3 = 3.4 \times 10^3 \text{ m}^3$$

To make the 1.45 M solution, measure **$3.4 \times 10^3 \text{ m}^3$** of the 2.50 M solution and add distilled water to make $5.8 \times 10^3 \text{ m}^3$.

- 13.64 Molality, $m = \frac{\text{moles of solute}}{\text{kg of solvent}}$

$$\text{a) } m \text{ glycine} = \frac{85.4 \text{ g Glycine} \left(\frac{1 \text{ mol Glycine}}{75.07 \text{ g Glycine}} \right)}{(1.270 \text{ kg})} = 0.895752 = \mathbf{0.896 \text{ m glycine}}$$

$$\text{b) } m \text{ glycerol} = \frac{8.59 \text{ g Glycerol} \left(\frac{1 \text{ mol Glycerol}}{92.09 \text{ g Glycerol}} \right)}{(77.0 \text{ g})} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 1.2114 = \mathbf{1.21 \text{ m glycerol}}$$

13.65 Molality = moles solute/kg solvent.

$$\text{a) } m \text{ HCl} = \frac{174 \text{ g HCl} \left(\frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \right)}{(757 \text{ g})} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 6.3043 = \mathbf{6.30 \text{ } m \text{ HCl}}$$

$$\text{b) } m \text{ naphthalene} = \frac{16.5 \text{ g Naphthalene} \left(\frac{1 \text{ mol Naphthalene}}{128.16 \text{ g Naphthalene}} \right)}{(53.3 \text{ g})} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \\ = 2.41548 = \mathbf{2.42 \text{ } m \text{ naphthalene}}$$

13.66 Molality, $m = \frac{\text{moles of solute}}{\text{kg of solvent}}$ Use density to convert volume to mass.

$$m \text{ benzene} = \frac{(44.0 \text{ mL C}_6\text{H}_6) \left(\frac{0.877 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol C}_6\text{H}_6}{78.11 \text{ g C}_6\text{H}_6} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)}{(167 \text{ mL C}_6\text{H}_{14}) \left(\frac{0.660 \text{ g}}{\text{mL}} \right)} = 4.48214 = \mathbf{4.48 \text{ } m \text{ C}_6\text{H}_6}$$

13.67 Molality = moles solute/kg of solvent.

$$m \text{ CCl}_4 = \frac{(2.66 \text{ mL CCl}_4) \left(\frac{1.59 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mol CCl}_4}{153.81 \text{ g CCl}_4} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)}{(76.5 \text{ mL CH}_2\text{Cl}_2) \left(\frac{1.33 \text{ g}}{\text{mL}} \right)} = 0.2702596 = \mathbf{0.270 \text{ } m \text{ CCl}_4}$$

13.68 a) The total mass of the solution is $3.10 \times 10^2 \text{ g}$, so $\text{mass}_{\text{solute}} + \text{mass}_{\text{solvent}} = 3.10 \times 10^2 \text{ g}$.

$$\text{Mass of C}_2\text{H}_6\text{O}_2 \text{ in } 1000 \text{ g (1 kg) of H}_2\text{O} = \left(\frac{0.125 \text{ mol C}_2\text{H}_6\text{O}_2}{1 \text{ kg H}_2\text{O}} \right) \left(\frac{62.07 \text{ g C}_2\text{H}_6\text{O}_2}{1 \text{ mol C}_2\text{H}_6\text{O}_2} \right) \\ = 7.75875 \text{ g C}_2\text{H}_6\text{O}_2 \text{ in } 1000 \text{ g H}_2\text{O}$$

$$\text{Grams of this solution} = 1000 \text{ g H}_2\text{O} + 7.75875 \text{ g C}_2\text{H}_6\text{O}_2 = 1007.75875 \text{ g}$$

$$\text{Mass C}_2\text{H}_6\text{O}_2 = \left(\frac{7.75875 \text{ g C}_2\text{H}_6\text{O}_2}{1007.75875 \text{ g solution}} \right) (3.10 \times 10^2 \text{ g solution}) = 2.386695 \text{ g C}_2\text{H}_6\text{O}_2$$

$$\text{Mass}_{\text{solvent}} = 3.10 \times 10^2 \text{ g} - \text{mass}_{\text{solute}} = 3.10 \times 10^2 \text{ g} - 2.386695 \text{ g C}_2\text{H}_6\text{O}_2 = 307.613305 = 308 \text{ g H}_2\text{O}$$

Therefore, **add 2.39 g C₂H₆O₂ to 308 g of H₂O** to make a 0.125 *m* solution.

b) This is a disguised dilution problem. First, determine the amount of solute in your target solution:

$$\left(\frac{2.20\%}{100\%} \right) (1.20 \text{ kg}) = 0.0264 \text{ kg HNO}_3 \text{ (solute)}$$

Then determine the amount of the concentrated acid solution needed to get 0.0264 kg solute:

$$\left(\frac{52.0\%}{100\%} \right) (\text{mass needed}) = 0.0264 \text{ kg}$$

$$\text{Mass solute needed} = 0.050769 = 0.0508 \text{ kg}$$

$$\text{Mass solvent} = \text{Mass solution} - \text{Mass solute} = 1.20 \text{ kg} - 0.050769 \text{ kg} = 1.149231 = 1.15 \text{ kg}$$

Add 0.0508 kg of the 52.0% (w/w) HNO₃ to 1.15 kg H₂O to make 1.20 kg of 2.20% (w/w) HNO₃.

- 13.69 a) The total weight of the solution is 1.50 kg, so

$$\text{mass}_{\text{solute}} + \text{mass}_{\text{solvent}} = 1.50 \text{ kg}$$

$$\text{g C}_2\text{H}_6\text{O}_2 / 1000\text{g H}_2\text{O in } 0.0355\text{m} = \left(\frac{0.0355 \text{ mol C}_2\text{H}_5\text{OH}}{1 \text{ kg H}_2\text{O}} \right) \left(\frac{46.07 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \right)$$

$$= 1.635485 \text{ g C}_2\text{H}_5\text{OH} / 1000 \text{ g H}_2\text{O (unrounded)}$$

$$\text{Grams of this solution} = 1000 \text{ g H}_2\text{O} + 1.635485 \text{ g C}_2\text{H}_5\text{OH} = 1001.635485 \text{ g (unrounded)}$$

$$\text{Mass C}_2\text{H}_6\text{O}_2 = [1.635485 \text{ g C}_2\text{H}_5\text{OH} / 1001.635485 \text{ g solution}] [1.50 \text{ kg solution} (10^3 \text{ g} / 1 \text{ kg})]$$

$$= 2.449222 = 2.45 \text{ g C}_2\text{H}_5\text{OH}$$

$$\text{Mass}_{\text{solvent}} = 1500 \text{ g} - \text{mass}_{\text{solute}} = 1500 \text{ g} - 2.449222 \text{ g C}_2\text{H}_6\text{O}_2 = 1497.551 = 1498 \text{ g H}_2\text{O}$$

Therefore, **add 2.45 g C₂H₅OH to 1498 g of H₂O** to make a 0.0355 *m* solution.

- b) This is a disguised dilution problem. First, determine the amount of solute in your target solution:

$$(445 \text{ g}) (13.0\% / 100\%) = 57.85 \text{ g HCl (solute) (unrounded)}$$

Then determine the amount of the concentrated acid solution needed to get 57.85 g solute:

$$(\text{Mass needed}) (34.1\% / 100\%) = 57.85 \text{ g}$$

$$\text{Mass solute needed} = 169.6481 = 170. \text{ g}$$

$$\text{Mass solvent} = \text{Mass solution} - \text{Mass solute} = 445 \text{ g} - 169.6481 \text{ g} = 275.35191 = 275 \text{ g}$$

Add 170. g of the 34.1% (w/w) HCl to 275 g H₂O

- 13.70 a) Mole fraction is moles of isopropanol per total moles.

$$X_{\text{isopropanol}} = \frac{0.35 \text{ mol Isopropanol}}{(0.35 + 0.85) \text{ mol}} = 0.2916667 = \mathbf{0.29}$$
 (Notice that mole fractions have no units.)

- b) Mass percent = $\frac{\text{mass of solute}}{\text{mass of solution}} (100)$. From the mole amounts, find the masses of isopropanol and water:

$$\text{Mass isopropanol} = (0.35 \text{ mol C}_3\text{H}_7\text{OH}) \left(\frac{60.09 \text{ g C}_3\text{H}_7\text{OH}}{1 \text{ mol C}_3\text{H}_7\text{OH}} \right) = 21.0315 \text{ g isopropanol}$$

$$\text{Mass water} = (0.85 \text{ mol H}_2\text{O}) \left(\frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) = 15.317 \text{ g water (unrounded)}$$

$$\text{Percent isopropanol} = \frac{(21.0315 \text{ g Isopropanol})}{(21.0315 + 15.317) \text{ g}} \times 100\% = 57.860710 = \mathbf{58\%}$$
 isopropanol

- c) Molality of isopropanol is moles of isopropanol per kg of solvent.

$$\text{Molality isopropanol} = \frac{0.35 \text{ mol Isopropanol}}{15.317 \text{ g Water}} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 22.85043 = \mathbf{23 m}$$
 isopropanol

- 13.71 a) Mole fraction is moles of NaCl per total moles.

$$X_{\text{NaCl}} = \frac{0.100 \text{ mol NaCl}}{(0.100 + 8.60) \text{ mol}} = 0.01149425 = \mathbf{0.0115}$$
 (Notice that mole fractions have no units.)

- b) Mass percent is the mass of NaCl per 100 g of solution.

$$\text{Mass NaCl} = (0.100 \text{ mol NaCl}) (58.44 \text{ g/mol}) = 5.844 \text{ g NaCl (unrounded)}$$

$$\text{Mass water} = (8.60 \text{ mol water}) (18.02 \text{ g/mol}) = 154.972 \text{ g water (unrounded)}$$

$$\text{Percent NaCl} = \frac{(5.844 \text{ g NaCl})}{(5.844 + 154.972) \text{ g}} \times 100\% = 3.63396677 = \mathbf{3.63\%}$$
 NaCl

- c) Molality of NaCl is moles of NaCl per kg of solvent.

$$\text{Molality NaCl} = \frac{0.100 \text{ mol NaCl}}{154.972 \text{ g Water}} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 0.645277856 = \mathbf{0.645 m}$$
 NaCl

13.72 The density of water is 1.00 g/mL. The mass of water is:

$$\text{Mass of water} = (0.500 \text{ L}) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1.00 \text{ g}}{1 \text{ mL}} \right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 0.500 \text{ kg}$$

$$0.400 \text{ m CsBr} = \frac{\text{moles CsBr}}{0.500 \text{ kg H}_2\text{O}}$$

$$\text{Moles CsBr} = 0.2000 \text{ mol (unrounded)}$$

$$\text{Mass of CsBr} = (0.200 \text{ mol CsBr}) \left(\frac{212.8 \text{ g CsBr}}{1 \text{ mol CsBr}} \right) = 42.56 = \mathbf{42.6 \text{ g CsBr}}$$

$$\text{Mass H}_2\text{O} = \text{Mass of solution} - \text{Mass of CsBr} =$$

$$(0.500 \text{ kg}) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) - 42.56 \text{ g CsBr} = 457.44 \text{ g H}_2\text{O}$$

$$\text{Moles H}_2\text{O} = (457.44 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 25.38513 \text{ mol H}_2\text{O}$$

$$X_{\text{CsBr}} = \frac{\text{mol CsBr}}{\text{mol CsBr} + \text{mol H}_2\text{O}} = \frac{0.2000 \text{ mol CsBr}}{(0.2000 + 25.38513) \text{ mol}} = 7.817 \times 10^{-3} = \mathbf{7.82 \times 10^{-3}}$$

$$\text{Percent CsBr} = \frac{\text{mass of CsBr}}{\text{mass of solution}} (100) = \frac{(42.56 \text{ g CsBr})}{(42.56 + 457.44) \text{ g}} \times 100\% = 8.512 = \mathbf{8.51\% \text{ CsBr}}$$

13.73 The density of water is 1.00 g/mL. The mass of water is:

$$\text{Mass of water} = (0.400 \text{ L}) (1 \text{ mL}/10^{-3} \text{ L}) (1.00 \text{ g/mL}) = 4.00 \times 10^2 \text{ g H}_2\text{O}$$

$$\text{Moles H}_2\text{O} = (400. \text{ g H}_2\text{O}) (1 \text{ mol H}_2\text{O}/18.02 \text{ g H}_2\text{O}) = 22.197558 \text{ mol H}_2\text{O}$$

$$\text{Moles KI} = (0.30 \text{ g KI}) (1 \text{ mol KI}/166.0 \text{ g KI}) = 1.80723 \times 10^{-3} \text{ mol KI}$$

$$X_{\text{KI}} = \frac{1.80723 \times 10^{-3} \text{ mol KI}}{(1.80723 \times 10^{-3} + 22.197558) \text{ mol}} = 8.14091 \times 10^{-5} = \mathbf{8.1 \times 10^{-5}}$$

$$\text{Percent KI} = \frac{(0.30 \text{ g KI})}{(0.30 + 400.) \text{ g}} \times 100\% = 0.07494 = \mathbf{0.075\% \text{ KI}}$$

13.74 The information given is 8.00 mass % NH₃ solution with a density of 0.9651 g/mL.

For convenience, choose exactly 100.00 grams of solution.

Determine some fundamental quantities:

$$\text{Mass of NH}_3 = (100 \text{ g solution}) \left(\frac{8.00\% \text{ NH}_3}{100\% \text{ solution}} \right) = 8.00 \text{ g NH}_3$$

$$\text{Mass H}_2\text{O} = \text{mass of solution} - \text{mass NH}_3 = (100.00 - 8.00) \text{ g} = 92.00 \text{ g H}_2\text{O}$$

$$\text{Moles NH}_3 = (8.00 \text{ g NH}_3) \left(\frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} \right) = 0.469759 \text{ mol NH}_3 \text{ (unrounded)}$$

$$\text{Moles H}_2\text{O} = (92.00 \text{ g H}_2\text{O}) \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 5.1054 \text{ mol H}_2\text{O (unrounded)}$$

$$\text{Volume solution} = (100.00 \text{ g solution}) \left(\frac{1 \text{ mL solution}}{0.9651 \text{ g solution}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 0.103616 \text{ L (unrounded)}$$

Using the above fundamental quantities and the definitions of the various units:

$$\text{Molality} = \frac{\text{Moles solute}}{\text{kg solvent}} = \left(\frac{0.469759 \text{ mol NH}_3}{92.00 \text{ g H}_2\text{O}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 5.106076 = \mathbf{5.11 \text{ m NH}_3}$$

$$\text{Molarity} = \frac{\text{Moles solute}}{\text{L solution}} = \left(\frac{0.469759 \text{ mol NH}_3}{0.103616 \text{ L}} \right) = 4.53365 = \mathbf{4.53 \text{ M NH}_3}$$

$$\text{Mole fraction} = X = \frac{\text{Moles substance}}{\text{total moles}} = \frac{0.469759 \text{ mol NH}_3}{(0.469759 + 5.1054) \text{ mol}} = 0.084259 = \mathbf{0.0843}$$

- 13.75 The information given is 28.8 mass % FeCl₃ solution with a density of 1.280 g/mL. For convenience, choose exactly 100.00 grams of solution.

Determine some fundamental quantities:

$$\text{Mass of FeCl}_3 = (100.00 \text{ grams solution}) (28.8\% \text{ FeCl}_3 / 100\%) = 28.8 \text{ g FeCl}_3$$

$$\text{Mass H}_2\text{O} = \text{mass of solution} - \text{mass FeCl}_3 = (100.00 - 28.8) \text{ g} = 71.20 \text{ g H}_2\text{O}$$

$$\text{Moles FeCl}_3 = (28.80 \text{ g FeCl}_3) (1 \text{ mol FeCl}_3 / 162.20 \text{ g FeCl}_3) = 0.1775586 \text{ mol FeCl}_3 \text{ (unrounded)}$$

$$\text{Moles H}_2\text{O} = (71.20 \text{ g H}_2\text{O}) (1 \text{ mol H}_2\text{O} / 18.02 \text{ g H}_2\text{O}) = 3.951165 \text{ mol H}_2\text{O} \text{ (unrounded)}$$

$$\text{Volume solution} = (100.00 \text{ g solution}) (1 \text{ mL} / 1.280 \text{ g}) (10^{-3} \text{ L} / 1 \text{ mL}) = 0.078125 \text{ L (unrounded)}$$

Using the above fundamental quantities and the definitions of the various units:

$$\text{Molality} = \text{Moles solute} / \text{kg solvent} = \left(\frac{0.1775586 \text{ mol FeCl}_3}{71.20 \text{ g H}_2\text{O}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 2.49380 = \mathbf{2.49 \text{ m FeCl}_3}$$

$$\text{Molarity} = \text{Moles solute} / \text{L solution} = \frac{0.1775586 \text{ mol FeCl}_3}{0.078125 \text{ L}} = 2.272750 = \mathbf{2.27 \text{ M FeCl}_3}$$

$$\text{Mole fraction} = X = \text{moles substance} / \text{total moles} = \frac{0.1775586 \text{ mol FeCl}_3}{(0.1775586 + 3.951165) \text{ mol}} = 0.043005688 = \mathbf{0.0430}$$

13.76 $\text{ppm} = \left(\frac{\text{mass solute}}{\text{mass solution}} \right) \times 10^6$

$$\text{The mass of 100.0 L of waste water solution is } (100.0 \text{ L solution}) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1.001 \text{ g}}{1 \text{ mL}} \right) = 1.001 \times 10^5 \text{ g.}$$

$$\text{ppm Ca}^{2+} = \left(\frac{0.25 \text{ g Ca}^{2+}}{1.001 \times 10^5 \text{ g solution}} \right) \times 10^6 = 2.49750 = \mathbf{2.5 \text{ ppm Ca}^{2+}}$$

$$\text{ppm Mg}^{2+} = \left(\frac{0.056 \text{ g Mg}^{2+}}{1.001 \times 10^5 \text{ g solution}} \right) \times 10^6 = 0.5594406 = \mathbf{0.56 \text{ ppm Mg}^{2+}}$$

- 13.77 The information given is ethylene glycol has a density of 1.114 g/mL and a molar mass of 62.07 g/mol. Water has a density of 1.00 g/mL. The solution has a density of 1.070 g/mL.

For convenience, choose exactly 1.0000 Liters as the equal volumes mixed. Ethylene glycol will be designated EG.

Determine some fundamental quantities:

$$\text{Mass of EG} = (1.0000 \text{ L EG}) (1 \text{ mL} / 10^{-3} \text{ L}) (1.114 \text{ g EG/mL}) = 1114 \text{ g EG}$$

$$\text{Mass of H}_2\text{O} = (1.0000 \text{ L H}_2\text{O}) (1 \text{ mL} / 10^{-3} \text{ L}) (1.00 \text{ g H}_2\text{O/mL}) = 1.00 \times 10^3 \text{ g H}_2\text{O}$$

$$\text{Moles EG} = (1114 \text{ g EG}) (1 \text{ mol EG} / 62.07 \text{ g EG}) = 17.94747865 \text{ mol EG (unrounded)}$$

$$\text{Moles H}_2\text{O} = (1.00 \times 10^3 \text{ g H}_2\text{O}) (1 \text{ mol H}_2\text{O} / 18.02 \text{ g H}_2\text{O}) = 55.49389567 \text{ mol H}_2\text{O (unrounded)}$$

$$\begin{aligned} \text{Volume solution} &= (1114 \text{ g EG} + 1.00 \times 10^3 \text{ g H}_2\text{O}) (\text{mL} / 1.070 \text{ g}) (10^{-3} \text{ L} / 1 \text{ mL}) \\ &= 1.97570 \text{ L (unrounded)} \end{aligned}$$

Using the above fundamental quantities and the definitions of the various units:

a) Volume percent = $(1.0000 \text{ L EG} / 1.97570 \text{ L}) 100\% = 50.61497 = \mathbf{50.61\% \text{ v/v}}$

b) Mass percent = $[(1114 \text{ g EG}) / (1114 + 1.00 \times 10^3 \text{ g})] 100\% = 52.6963 = \mathbf{52.7\% \text{ w/w}}$

$$\text{c) Molarity} = \text{Moles solute} / \text{L solution} = \frac{17.94747865 \text{ mol EG}}{1.97570 \text{ L}} = 9.08411 = \mathbf{9.08 \text{ M ethylene glycol}}$$

$$\text{d) Molality} = \text{Moles solute} / \text{kg solvent} = \frac{17.94747865 \text{ mol EG}}{1.00 \times 10^3 \text{ g H}_2\text{O}} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)$$

$$= 17.94747865 = \mathbf{17.9 \text{ m ethylene glycol}}$$

$$\text{e) Mole fraction} = X_{\text{EG}} = \text{moles substance} / \text{total moles} = \frac{17.94747865 \text{ mol EG}}{(17.94747865 + 55.49389567) \text{ mol}}$$

$$= 0.244378 = \mathbf{0.244}$$

- 13.78 Colligative properties of a solution are affected by the number of particles of solute in solution. The density of a solution would be affected by the chemical formula of the solute.
- 13.79 A nonvolatile nonelectrolyte is a covalently bonded molecule that does not dissociate into ions or evaporate when dissolved in a solvent. In this case, the colligative concentration is equal to the molar concentration, simplifying calculations.
- 13.80 The “strong” in “strong electrolyte” refers to the ability of an electrolyte solution to conduct a large current. This conductivity occurs because solutes that are strong electrolytes dissociate completely into ions when dissolved in water.
- 13.81 Raoult’s Law states that the vapor pressure of solvent above the solution equals the mole fraction of the solvent times the vapor pressure of the pure solvent. Raoult’s Law is not valid for a solution of a volatile solute in solution. Both solute and solvent would evaporate based upon their respective vapor pressures.
- 13.82 The boiling point temperature is higher and the freezing point temperature is lower for the solution compared to the solvent because the addition of a solute lowers the freezing point and raises the boiling point of a liquid.
- 13.83 **Yes**, the vapor at the top of the fractionating column is richer in content of the more volatile component.
- 13.84 The boiling point of a 0.01 *m* KF solution is higher than that of 0.01 *m* glucose. KF dissociates into ions in water (K^+ and F^-) while the glucose does not, so the KF produces more particles.
- 13.85 A dilute solution of an electrolyte behaves more ideally than a concentrated one. With increasing concentration, the effective concentration deviates from the molar concentration because of ionic attractions.. Thus, the more dilute **0.050 *m* NaF** solution has a boiling point closer to its predicted value.
- 13.86 Univalent ions behave more ideally than divalent ions. Ionic strength (which affects “activity” concentration) is greater for divalent ions. Thus, **0.01 *m* NaBr** has a freezing point that is closer to its predicted value.
- 13.87 Cyclohexane, with a freezing point depression constant of 20.1°C/m, would make calculation of molar mass of a substance easier, since ΔT_f would be greater.
- 13.88 Strong electrolytes are substances that produce a large number of ions when dissolved in water; strong acids and bases and soluble salts are strong electrolytes. Weak electrolytes produce few ions when dissolved in water; weak acids and bases are weak electrolytes. Nonelectrolytes produce no ions when dissolved in water. Molecular compounds other than acids and bases are nonelectrolytes.
- a) **Strong electrolyte** When hydrogen chloride is bubbled through water, it dissolves and dissociates completely into H^+ (or H_3O^+) ions and Cl^- ions. HCl is a strong acid.
- b) **Strong electrolyte** Potassium nitrate is a soluble salt.
- c) **Nonelectrolyte** Glucose solid dissolves in water to form individual $\text{C}_6\text{H}_{12}\text{O}_6$ molecules, but these units are not ionic and therefore do not conduct electricity. Glucose is a molecular compound.
- d) **Weak electrolyte** Ammonia gas dissolves in water, but is a weak base that forms few NH_4^+ and OH^- ions.

- 13.89 a) NaMnO₄ **strong electrolyte**
 b) CH₃COOH **weak electrolyte**
 c) CH₃OH **nonelectrolyte**
 d) Ca(C₂H₃O₂)₂ **strong electrolyte**

13.90 To count solute particles in a solution of an ionic compound, count the number of ions per mole and multiply by the number of moles in solution. For a covalent compound, the number of particles equals the number of molecules.

$$a) \left(\frac{0.3 \text{ mol KBr}}{\text{L}} \right) \left(\frac{2 \text{ mol particles}}{1 \text{ mol KBr}} \right) (1 \text{ L}) = \mathbf{0.6 \text{ mol of particles}}$$

Each KBr forms one K⁺ ion and one Br⁻ ion, 2 particles for each KBr.

$$b) \left(\frac{0.065 \text{ mol HNO}_3}{\text{L}} \right) \left(\frac{2 \text{ mol particles}}{1 \text{ mol HNO}_3} \right) (1 \text{ L}) = \mathbf{0.13 \text{ mol of particles}}$$

HNO₃ is a strong acid that forms H⁺(H₃O⁺) ions and NO₃⁻ ions in aqueous solution.

$$c) \left(\frac{10^{-4} \text{ mol KHSO}_4}{\text{L}} \right) \left(\frac{2 \text{ mol particles}}{1 \text{ mol KHSO}_4} \right) (1 \text{ L}) = \mathbf{2 \times 10^{-4} \text{ mol of particles}}$$

Each KHSO₄ forms 1 K⁺ ion and 1 HSO₄⁻ ion in aqueous solution, 2 particles for each KHSO₄.

$$d) \left(\frac{0.06 \text{ mol C}_2\text{H}_5\text{OH}}{\text{L}} \right) \left(\frac{1 \text{ mol particles}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \right) (1 \text{ L}) = \mathbf{0.06 \text{ mol of particles}}$$

Ethanol is not an ionic compound so each molecule dissolves as one particle. The number of moles of particles is the same as the number of moles of molecules, **0.06 mol** in 1 L.

- 13.91 a) (0.02 mol CuSO₄/L) (2 mol particles/mol CuSO₄) (10⁻³ L / 1 mL) (1 mL) = **4 x 10⁻⁵ mol of particles**
 b) (0.004 mol Ba(OH)₂/L) (3 mol particles/mol Ba(OH)₂) (10⁻³ L / 1 mL) (1 mL)
 = 1.2 x 10⁻⁵ = **1 x 10⁻⁵ mol of particles**
 c) (0.08 mol C₅H₅N/L) (1 mol particles/mol C₅H₅N) (10⁻³ L / 1 mL) (1 mL) = **8 x 10⁻⁵ mol of particles**
 d) (0.05 mol (NH₄)₂CO₃/L) (3 mol particles/mol (NH₄)₂CO₃) (10⁻³ L / 1 mL) (1 mL)
 = 1.5 x 10⁻⁴ = **2 x 10⁻⁴ mol of particles**

13.92 The magnitude of freezing point depression is directly proportional to molality.

$$a) \text{ Molality of CH}_3\text{OH} = \frac{(11.0 \text{ g CH}_3\text{OH})}{(100. \text{ g H}_2\text{O})} \left(\frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 3.4332085 = 3.43 \text{ m CH}_3\text{OH}$$

$$\text{Molality of CH}_3\text{CH}_2\text{OH} = \frac{(22.0 \text{ g CH}_3\text{CH}_2\text{OH})}{(200. \text{ g H}_2\text{O})} \left(\frac{1 \text{ mol CH}_3\text{CH}_2\text{OH}}{46.07 \text{ g CH}_3\text{CH}_2\text{OH}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)$$

$$= 2.387671 = 2.39 \text{ m CH}_3\text{CH}_2\text{OH}$$

The molality of methanol, CH₃OH, in water is 3.43 *m* whereas the molality of ethanol, CH₃CH₂OH, in water is 2.39 *m*. Thus, **CH₃OH/H₂O solution** has the lower freezing point.

$$b) \text{ Molality of H}_2\text{O} = \frac{(20.0 \text{ g H}_2\text{O})}{(1.00 \text{ kg CH}_3\text{OH})} \left(\frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 1.10988 = 1.11 \text{ m H}_2\text{O}$$

$$\text{Molality of CH}_3\text{CH}_2\text{OH} = \frac{(20.0 \text{ g CH}_3\text{CH}_2\text{OH})}{(1.00 \text{ kg CH}_3\text{OH})} \left(\frac{1 \text{ mol CH}_3\text{CH}_2\text{OH}}{46.07 \text{ g CH}_3\text{CH}_2\text{OH}} \right) = 0.434122 = 0.434 \text{ m CH}_3\text{CH}_2\text{OH}$$

The molality of H₂O in CH₃OH is 1.11 *m*, whereas CH₃CH₂OH in CH₃OH is 0.434 *m*. Therefore, **H₂O/CH₃OH solution** has the lower freezing point.

13.93 The magnitude of boiling point elevation is directly proportional to molality.

$$\text{a) Molality of } C_3H_8O_3 = \frac{(38.0 \text{ g } C_3H_8O_3)}{(250. \text{ g Ethanol})} \left(\frac{1 \text{ mol } C_3H_8O_3}{92.09 \text{ g } C_3H_8O_3} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 1.650559 = 1.65 \text{ m } C_3H_8O_3$$

$$\text{Molality of } C_2H_6O_2 = \frac{(38.0 \text{ g } C_2H_6O_2)}{(250. \text{ g Ethanol})} \left(\frac{1 \text{ mol } C_2H_6O_2}{62.07 \text{ g } C_2H_6O_2} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 2.44885 = 2.45 \text{ m } C_2H_6O_2$$

The molality of $C_2H_6O_2$, in ethanol is 2.45 *m* whereas the molality of $C_3H_8O_3$, in ethanol is 1.65 *m*. Thus, **$C_2H_6O_2$ /ethanol solution** has the higher boiling point.

$$\text{b) Molality of } C_2H_6O_2 = \frac{(15 \text{ g } C_2H_6O_2)}{(0.50 \text{ kg } H_2O)} \left(\frac{1 \text{ mol } C_2H_6O_2}{62.07 \text{ g } C_2H_6O_2} \right) = 0.4833253 = 0.48 \text{ m } C_2H_6O_2$$

$$\text{Molality of NaCl} = \frac{(15 \text{ g NaCl})}{(0.50 \text{ kg } H_2O)} \left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \right) = 0.513347 = 0.51 \text{ m NaCl}$$

Since the NaCl is a strong electrolyte, the molality of particles would be:

$$(2 \text{ particles/NaCl}) (0.513347 \text{ mol NaCl/kg}) = 1.026694 = 1.0 \text{ m particles}$$

The molality of $C_2H_6O_2$ in H_2O is 0.48 *m*, whereas NaCl in H_2O is 1.0 *m*. Therefore, **NaCl/ H_2O solution** has the higher boiling point.

13.94 To rank the solutions in order of increasing osmotic pressure, boiling point, freezing point, and vapor pressure, convert the molality of each solute to molality of particles in the solution. The higher the molality of particles, the higher the osmotic pressure, the higher the boiling point, the lower the freezing point, and the lower the vapor pressure at a given temperature.

$$\text{(I) } (0.100 \text{ m NaNO}_3) \left(\frac{2 \text{ mol particles}}{1 \text{ mol NaNO}_3} \right) = 0.200 \text{ m ions}$$

$NaNO_3$ consists of Na^+ ions and NO_3^- ions, 2 particles for each $NaNO_3$.

$$\text{(II) } (0.100 \text{ m glucose}) \left(\frac{1 \text{ mol particles}}{1 \text{ mol glucose}} \right) = 0.100 \text{ m molecules}$$

Ethanol is not an ionic compound so each molecule dissolves as one particle. The number of moles of particles is the same as the number of moles of molecules.

$$\text{(III) } (0.100 \text{ m CaCl}_2) \left(\frac{3 \text{ mol particles}}{1 \text{ mol CaCl}_2} \right) = 0.300 \text{ m ions}$$

$CaCl_2$ consists of Ca^{+2} ions and Cl^- ions, 3 particles for each $CaCl_2$.

- | | |
|----------------------------|--------------------------------|
| a) Osmotic pressure: | $\Pi_{II} < \Pi_I < \Pi_{III}$ |
| b) Boiling point: | $bp_{II} < bp_I < bp_{III}$ |
| c) Freezing point: | $fp_{III} < fp_I < fp_{II}$ |
| d) Vapor pressure at 50°C: | $vp_{III} < vp_I < vp_{II}$ |

13.95 I 0.04 *m* $(H_2N)_2CO$ x 1 mol particles / 1 mol $(H_2N)_2CO$ = 0.04 *m* molecules

II 0.01 *m* $AgNO_3$ x 2 mol particles / 1 mol $AgNO_3$ = 0.02 *m* ions

III 0.03 *m* $CuSO_4$ x 2 mol particles / 1 mol $CuSO_4$ = 0.06 *m* ions

- | | |
|-----------------------------|--------------------------------|
| a) Osmotic pressure: | $\Pi_{II} < \Pi_I < \Pi_{III}$ |
| b) Boiling point: | $bp_{II} < bp_I < bp_{III}$ |
| c) Freezing point: | $fp_{III} < fp_I < fp_{II}$ |
| d) Vapor pressure at 298 K: | $vp_{III} < vp_I < vp_{II}$ |

13.96 The mol fraction of solvent affects the vapor pressure according to the equation: $P_{\text{solvent}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}}$

$$\text{Moles } C_3H_8O_3 = (34.0 \text{ g } C_3H_8O_3) \left(\frac{1 \text{ mol } C_3H_8O_3}{92.09 \text{ g } C_3H_8O_3} \right) = 0.369204 \text{ mol } C_3H_8O_3 \text{ (unrounded)}$$

$$\text{Moles } H_2O = (500.0 \text{ g } H_2O) \left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \right) = 27.7469 \text{ mol } H_2O \text{ (unrounded)}$$

$$X_{\text{solvent}} = \frac{\text{mol } H_2O}{\text{mol } H_2O + \text{mol glycerol}} = \frac{27.7469 \text{ mol } H_2O}{27.7469 \text{ mol } H_2O + 0.369204 \text{ mol glycerol}} = 0.9868686$$

$$P_{\text{solvent}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}} = (0.9868686) (23.76 \text{ torr}) = 23.447998 = \mathbf{23.4 \text{ torr}}$$

13.97 The mole fraction of solvent affects the vapor pressure according to the equation: $P_{\text{solvent}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}}$

$$X_{\text{solvent}} = (5.4 \text{ mol toluene}) / [(0.39) + (5.4)] \text{ mol} = 0.93264 \text{ (unrounded)}$$

$$P_{\text{solvent}} = X_{\text{solvent}} P^{\circ}_{\text{solvent}} = (0.93264) (41 \text{ torr}) = 38.2382 = \mathbf{38 \text{ torr}}$$

13.98 The change in freezing point is calculated from $\Delta T_f = iK_f m$, where K_f is $1.86^{\circ}\text{C}/m$ for aqueous solutions, i is the van't Hoff factor, and m is the molality of particles in solution. Since urea is a covalent compound and does not ionize in water, $i = 1$. Once ΔT_f is calculated, the freezing point is determined by subtracting it from the freezing point of pure water (0.00°C).

$$\Delta T_f = iK_f m = (1) (1.86^{\circ}\text{C}/m) (0.251 \text{ m}) = 0.46686^{\circ}\text{C} \text{ (unrounded)}$$

$$\text{The freezing point is } 0.00^{\circ}\text{C} - 0.46686^{\circ}\text{C} = -0.46686 = \mathbf{-0.467^{\circ}\text{C}}.$$

13.99 $\Delta T_b = iK_b m = (1) (0.512^{\circ}\text{C}/m) (0.200 \text{ m}) = 0.1024^{\circ}\text{C}$ (unrounded)

$$\text{The boiling point is } 100.00^{\circ}\text{C} + 0.1024^{\circ}\text{C} = 100.1024 = \mathbf{100.10^{\circ}\text{C}}.$$

13.100 The boiling point of a solution is increased relative to the pure solvent by the relationship $\Delta T_b = iK_b m$. Vanillin is a nonelectrolyte so $i = 1$. The molality must be calculated, and K_b is given ($1.22^{\circ}\text{C}/m$).

$$\text{Molality of Vanillin} = \frac{\text{moles of vanillin}}{\text{kg of solvent (ethanol)}} = \frac{(6.4 \text{ g Vanillin}) \left(\frac{1 \text{ mol Vanillin}}{152.14 \text{ g Vanillin}} \right)}{(50.0 \text{ g Ethanol})} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)$$

$$= 0.841330354 \text{ m Vanillin (unrounded)}$$

$$\Delta T_b = iK_b m = (1) (1.22^{\circ}\text{C}/m) (0.841330354 \text{ m}) = 1.02642^{\circ}\text{C} \text{ (unrounded)}$$

$$\text{The boiling point is } 78.5^{\circ}\text{C} + 1.02642^{\circ}\text{C} = 79.52642 = \mathbf{79.5^{\circ}\text{C}}.$$

13.101 Moles $C_{10}H_8 = (5.00 \text{ g } C_{10}H_8) (1 \text{ mol } C_{10}H_8 / 128.16 \text{ g } C_{10}H_8) = 0.0390137 \text{ mol } C_{10}H_8$ (unrounded)

$C_{10}H_8$ is a nonelectrolyte so $i = 1$

$$\text{Mass} = (444 \text{ g benzene}) (1 \text{ kg} / 10^3 \text{ g}) = 0.444 \text{ kg benzene}$$

$$\text{Molality} = (0.0390137 \text{ mol } C_{10}H_8) / (0.444 \text{ kg}) = 0.08786869 \text{ m (unrounded)}$$

$$\Delta T_f = i K_f m = (1) (4.90^{\circ}\text{C}/m) (0.08786869 \text{ m}) = 0.43056^{\circ}\text{C} \text{ (unrounded)}$$

$$\text{Freezing point} = (5.5 - 0.43056)^{\circ}\text{C} = 5.06944 = \mathbf{5.1^{\circ}\text{C}}$$

13.102 The molality of the solution can be determined from the relationship $\Delta T_f = iK_f m$ with the value $1.86^{\circ}\text{C}/m$ inserted for K_f and $i = 1$ for the nonelectrolyte ethylene glycol (ethylene glycol is a covalent compound that will form one particle per molecule when dissolved). Convert the freezing point of the solution to $^{\circ}\text{C}$ and find ΔT_f :

$$^{\circ}\text{C} = (5/9) (^{\circ}\text{F} - 32.0) = (5/9) ((-12.0)^{\circ}\text{F} - 32.0) = -24.44444^{\circ}\text{C} \text{ (unrounded)}$$

$$\Delta T_f = (0.00 - (-24.44444))^{\circ}\text{C} = 24.44444^{\circ}\text{C}$$

$$m = \frac{\Delta T_f}{K_f} = \frac{24.44444^{\circ}\text{C}}{1.86^{\circ}\text{C}/m} = 13.14217 \text{ m (unrounded)}$$

Ethylene glycol will be abbreviated as EG.

Multiply the molality by the given mass of solvent to find the mass of ethylene glycol that must be in solution.

$$\begin{aligned} \text{Mass ethylene glycol} &= \left(\frac{13.14217 \text{ mol EG}}{1 \text{ kg H}_2\text{O}} \right) (14.5 \text{ kg H}_2\text{O}) \left(\frac{62.07 \text{ g EG}}{1 \text{ mol EG}} \right) \\ &= 1.18282 \times 10^4 = \mathbf{1.18 \times 10^4 \text{ g ethylene glycol}} \end{aligned}$$

To prevent the solution from freezing, dissolve a minimum of 1.13×10^4 g ethylene glycol in 14.5 kg water.

- 13.103 The molality of the solution can be determined from the relationship $\Delta T_f = i K_f m$ with the value $1.86^\circ\text{C}/m$ inserted for K_f , $i = 1$ for the nonelectrolyte glycerol, and the given ΔT_f of -15°C .

$$m = \Delta T_f / K_f = (15^\circ\text{C}) / (1.86^\circ\text{C}/m) = 8.06452 \text{ m}$$

Glycerol will be abbreviated as GLY.

$$\begin{aligned} \text{Mass glycerol} &= \left(\frac{8.06452 \text{ mol GLY}}{1 \text{ kg H}_2\text{O}} \right) (11.0 \text{ mg H}_2\text{O}) \left(\frac{10^{-3} \text{ g}}{1 \text{ mg}} \right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left(\frac{92.09 \text{ g GLY}}{1 \text{ mol GLY}} \right) \\ &= 0.0081693 = \mathbf{0.0082 \text{ g glycerol}} \end{aligned}$$

To prevent the solution from freezing, dissolve a minimum of 0.0082 g glycerol in 11.0 mg water.

- 13.104 Convert the mass percent to molality and use $\Delta T = iK_b m$ to find the van't Hoff factor.

a) Assume exactly 100 grams of solution.

$$\text{Mass NaCl} = (100 \text{ g solution}) \left(\frac{1.00\% \text{ NaCl}}{100\% \text{ solution}} \right) = 1.00 \text{ g NaCl}$$

$$\text{Moles NaCl} = (1.00 \text{ g NaCl}) \left(\frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \right) = 0.0171116 \text{ mol NaCl}$$

$$\text{Mass H}_2\text{O} = 100.00 \text{ g solution} - 1.00 \text{ g NaCl} = 99.00 \text{ g H}_2\text{O}$$

$$\text{Molality NaCl} = \frac{\text{mole NaCl}}{\text{kg H}_2\text{O}} = \frac{0.0171116 \text{ mol NaCl}}{99.00 \text{ g H}_2\text{O}} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 0.172844 = \mathbf{0.173 \text{ m NaCl}}$$

$$\text{Calculate } \Delta T = (0.000 - (-0.593))^\circ\text{C} = 0.593^\circ\text{C}$$

$$\Delta T_f = iK_f m$$

$$i = \frac{\Delta T_f}{K_f m} = \frac{0.593^\circ\text{C}}{(1.86^\circ\text{C}/m)(0.172844 \text{ m})} = 1.844537 = \mathbf{1.84}$$

The value of i should be close to 2 because NaCl dissociates into 2 particles when dissolving in water.

b) For acetic acid, CH_3COOH :

Assume exactly 100 grams of solution.

$$\text{Mass CH}_3\text{COOH} = (100 \text{ g solution}) \left(\frac{0.500\% \text{ CH}_3\text{COOH}}{100\% \text{ solution}} \right) = 0.500 \text{ g CH}_3\text{COOH}$$

$$\text{Moles CH}_3\text{COOH} = (0.500 \text{ g CH}_3\text{COOH}) \left(\frac{1 \text{ mol CH}_3\text{COOH}}{60.05 \text{ g CH}_3\text{COOH}} \right) = 0.0083264 \text{ mol CH}_3\text{COOH}$$

$$\text{Mass H}_2\text{O} = 100.00 \text{ g solution} - 0.500 \text{ g CH}_3\text{COOH} = 99.500 \text{ g H}_2\text{O}$$

$$\text{Molality CH}_3\text{COOH} = \frac{\text{mole CH}_3\text{COOH}}{\text{kg H}_2\text{O}} = \frac{0.0083264 \text{ mol CH}_3\text{COOH}}{99.500 \text{ g H}_2\text{O}} \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)$$

$$= 0.083682 = \mathbf{0.0837 \text{ m CH}_3\text{COOH}}$$

$$\text{Calculate } \Delta T = (0.000 - (-0.159))^\circ\text{C} = 0.159^\circ\text{C}$$

$$\Delta T_f = iK_f m$$

$$i = \frac{\Delta T_f}{K_f m} = \frac{0.159^\circ\text{C}}{(1.86^\circ\text{C}/m)(0.083682 \text{ m})} = 1.02153 = \mathbf{1.02}$$

Acetic acid is a weak acid and dissociates to a small extent in solution, hence a van't Hoff factor that is close to 1.

- 13.105 Convert the mass % to molality and use $\Delta T = iK_b m$ to find the van't Hoff factor.
 a) Assume exactly 100 grams of solution. Thus, the solution contains 0.500 grams of KCl in 99.500 grams of water.

$$\text{Molality KCl} = \left(\frac{0.500 \text{ g KCl}}{99.500 \text{ g H}_2\text{O}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \right) = 0.067406 \text{ m KCl (unrounded)}$$

$$\text{Calculate } \Delta T = (0.000 - (-0.234))^\circ\text{C} = 0.234^\circ\text{C}$$

$$\Delta T_f = i K_f m$$

$$i = \Delta T_f / K_f m = (0.234^\circ\text{C}) / [(1.86^\circ\text{C/m})(0.067406 \text{ m})] = 1.866398 = \mathbf{1.87}$$

The value of i should be close to 2 because KCl dissociates into 2 particles when dissolving in water.

b) For sulfuric acid, H_2SO_4 :

Assume exactly 100 grams of solution. Thus, the solution contains 1.00 grams of H_2SO_4 in 99.00 grams of water.

$$\text{Molality H}_2\text{SO}_4 = \left(\frac{1.00 \text{ g H}_2\text{SO}_4}{99.00 \text{ g H}_2\text{O}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol H}_2\text{SO}_4}{98.09 \text{ g H}_2\text{SO}_4} \right) = 0.10297696 \text{ m H}_2\text{SO}_4 \text{ (unrounded)}$$

$$\text{Calculate } \Delta T = (0.000 - (-0.423))^\circ\text{C} = 0.423^\circ\text{C}$$

$$\Delta T_f = i K_f m$$

$$i = \Delta T_f / K_f m = (0.423^\circ\text{C}) / [(1.86^\circ\text{C/m})(0.10297696 \text{ m})] = 2.2084 = \mathbf{2.21}$$

Sulfuric acid is a strong acid and dissociates to give a hydrogen ion and a hydrogen sulfate ion. The hydrogen sulfate ion may further dissociate to another hydrogen ion and a sulfate ion. If ionization in both steps were complete, the value of the van't Hoff factor would be 3.

- 13.106 Osmotic pressure is calculated from the molarity of particles, gas constant and temperature. Convert the mass of sucrose to moles using the molar mass, and then to molarity. Sucrose is a nonelectrolyte so $i = 1$.

$$T = (273 + 20.)\text{K} = 293 \text{ K}$$

$$\text{Molarity} = \frac{\text{moles of sucrose}}{\text{volume of solution}} = \frac{(3.55 \text{ g sucrose}) \left(\frac{1 \text{ mol sucrose}}{342.30 \text{ g sucrose}} \right)}{1.0 \text{ L}} = 1.0371 \times 10^{-2} \text{ M sucrose}$$

$$\Pi = i MRT = (1)(1.0371 \times 10^{-2} \text{ mol/L})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(293 \text{ K}) = 0.2494778 = \mathbf{0.249 \text{ atm}}$$

A pressure greater than 0.249 atm must be applied to obtain pure water from a 3.55 g/L solution.

- 13.107 Use the osmotic pressure equation ($\Pi = i MRT$) to find the molarity of the solution (assuming $i = 1$).

$$M = \Pi / i RT = \frac{0.272 \text{ atm}}{(1) \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) ((273 + 25) \text{ K})} = 0.01111756 \text{ M (unrounded)}$$

$$\text{Moles} = (0.01111756 \text{ mol/L})(100.0 \text{ mL})(10^{-3} \text{ L} / 1 \text{ mL}) = 0.001111756 \text{ mol (unrounded)}$$

$$\text{Molar mass} = (6.053 \text{ g}) / (0.001111756 \text{ mol}) = 5.4445 \times 10^3 = \mathbf{5.44 \times 10^3 \text{ g/mol}}$$

- 13.108 The pressure of each compound is proportional to its mole fraction according to Raoult's law: $P_A = X_A P_A^\circ$

$$X_{\text{CH}_2\text{Cl}_2} = \frac{\text{moles CH}_2\text{Cl}_2}{\text{moles CH}_2\text{Cl}_2 + \text{mol CCl}_4} = \frac{1.60 \text{ mol}}{1.60 + 1.10 \text{ mol}} = 0.592593$$

$$X_{\text{CCl}_4} = \frac{\text{moles CCl}_4}{\text{moles CH}_2\text{Cl}_2 + \text{mol CCl}_4} = \frac{1.10 \text{ mol}}{1.60 + 1.10 \text{ mol}} = 0.407407$$

$$P_A = X_A P_A^\circ$$

$$= (0.592593)(352 \text{ torr}) = 208.593 = \mathbf{209 \text{ torr CH}_2\text{Cl}_2}$$

$$= (0.407407)(118 \text{ torr}) = 48.0740 = \mathbf{48.1 \text{ torr CCl}_4}$$

- 13.109 The fluid inside a bacterial cell is **both a solution and a colloid**. It is a solution of ions and small molecules and a colloid of large molecules, proteins, and nucleic acids.

- 13.110 a) milk - **liquid / liquid** colloid.
 b) fog - **liquid / gas** colloid.
 c) shaving cream - **gas / liquid** colloid
- 13.111 Brownian motion is a characteristic movement in which the colloidal particles change speed and direction erratically by the motion of the dispersing molecules.
- 13.112 When light passes through a colloid, it is scattered randomly by the dispersed particles because their sizes are similar to the wavelengths of visible light. Viewed from the side, the scattered beam is visible and broader than one passing through a solution, a phenomenon known as the Tyndall effect.
- 13.113 Soap micelles have nonpolar “tails” pointed inward and anionic “heads” pointed outward. The charges on the “heads” on one micelle repel the “heads” on a neighboring micelle because the charges are the same. This repulsion between soap micelles keeps them from coagulating.
 Soap is more effective in **freshwater** than in seawater because the divalent cations in seawater combine with the anionic “head” to form an insoluble precipitate.

- 13.114 a) Total molarity of ions:

$$\text{Solution A: } M = \frac{(8 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)}{25 \text{ mL}} = \mathbf{3.2 \text{ M}}$$

$$\text{Solution B: } M = \frac{(10 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)}{50 \text{ mL}} = \mathbf{2.0 \text{ M}}$$

$$\text{Solution C: } M = \frac{(12 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)}{100 \text{ mL}} = \mathbf{1.2 \text{ M}}$$

- b) Molarity of compound:

$$\text{Solution A: } M = \frac{(8 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mole compound}}{2 \text{ moles of ions}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)}{25 \text{ mL}} = \mathbf{1.6 \text{ M}}$$

$$\text{Solution B: } M = \frac{(10 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mole compound}}{2 \text{ moles of ions}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)}{50 \text{ mL}} = \mathbf{1.0 \text{ M}}$$

$$\text{Solution C: } M = \frac{(12 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mole compound}}{3 \text{ moles of ions}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)}{100 \text{ mL}} = \mathbf{0.40 \text{ M}}$$

Solution A has the highest molarity.

- c) Molality of compound:

$$\text{Solution A: } m = \frac{(8 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mole compound}}{2 \text{ moles of ions}} \right) \left(\frac{1 \text{ mL}}{1.0 \text{ g}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)}{25 \text{ mL}} = \mathbf{1.6 \text{ m}}$$

$$\text{Solution B: } m = \frac{(10 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mole compound}}{2 \text{ moles of ions}} \right) \left(\frac{1 \text{ mL}}{1.0 \text{ g}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right)}{50 \text{ mL}} = \mathbf{1.0 \text{ m}}$$

$$\text{Solution C: } m = \frac{(12 \text{ spheres}) \left(\frac{0.010 \text{ mol ions}}{1 \text{ sphere}} \right) \left(\frac{1 \text{ mole compound}}{3 \text{ moles of ions}} \right)}{100 \text{ mL}} \left(\frac{1 \text{ mL}}{1.0 \text{ g}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = \mathbf{0.40 \text{ m}}$$

Solution C has the lowest molality.

d) Osmotic pressure: assume a temperature of 298 K

$$\text{Solution A: } \Pi = i MRT = (2) (1.6 \text{ mol/L}) (0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K}) (298 \text{ K}) = 78.29056 = \mathbf{78 \text{ atm}}$$

$$\text{Solution B: } \Pi = i MRT = (2) (1.0 \text{ mol/L}) (0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K}) (298 \text{ K}) = 48.9316 = \mathbf{49 \text{ atm}}$$

$$\text{Solution C: } \Pi = i MRT = (3) (0.40 \text{ mol/L}) (0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K}) (298 \text{ K}) = 29.35896 = \mathbf{29 \text{ atm}}$$

Solution A has the highest osmotic pressure.

13.115 Assume exactly 100 grams of solution.

Thus, the solution contains:

$$(100 \text{ g}) (10\% \text{ glucose} / 100\%) = 10. \text{ g glucose}$$

$$(10. \text{ g glucose}) (1 \text{ mol glucose} / 180.16 \text{ g glucose}) = 0.055506216 \text{ mol glucose (unrounded)}$$

and 90. g of water

$$(90. \text{ g H}_2\text{O}) (1 \text{ kg} / 10^3 \text{ g}) = 0.090 \text{ kg}$$

The volume of the solution is:

$$(100 \text{ g}) (1 \text{ mL} / 1.039 \text{ g}) (10^{-3} \text{ L} / 1 \text{ mL}) = 0.096246 \text{ L (unrounded)}$$

$$\text{Molarity glucose} = (0.055506216 \text{ mol glucose}) / (0.096246 \text{ L}) = 0.57671 \text{ M glucose (unrounded)}$$

$$\text{Molality glucose} = (0.055506216 \text{ mol glucose}) / (0.090 \text{ kg}) = 0.6167357 \text{ m glucose (unrounded)}$$

Glucose is a nonelectrolyte so $i = 1$.

$$T = (273 + 20) = 293 \text{ K}$$

$$\Delta T_f = iK_f m = (1) (1.86^\circ\text{C/m}) (0.6167357 \text{ m}) = 1.1471^\circ\text{C}$$

$$\text{Freezing point} = (0.00 - 1.1471) = -1.1471 = \mathbf{-1.1^\circ\text{C}}$$

$$\Delta T_b = iK_b m = (1) (0.512^\circ\text{C/m}) (0.6167357 \text{ m}) = 0.3157687^\circ\text{C}$$

$$\text{Boiling point} = (100.00 + 0.3157687) = 100.3157687 = \mathbf{100.32^\circ\text{C}}$$

$$\Pi = iMRT = (1) (0.57671 \text{ mol/L}) (0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K}) (293 \text{ K}) = 13.8729 = \mathbf{14 \text{ atm}}$$

13.116 The density of the mixture will be the weighted average of the constituents. Thus, density of mixture = contribution from copper + contribution from zinc. The percent zinc plus the percent copper must total 100%. Zinc atoms are heavier than copper atoms so a factor equal to the ratio of their atomic weights (65.41 / 63.55) must be applied to the zinc contribution.

$$\begin{aligned} \text{a) Density of alloy} &= (90.0\% \text{ Cu} / 100\%) (8.95 \text{ g/cm}^3) + (10.0\% \text{ Zn} / 100\%) (65.41 / 63.55) (8.95 \text{ g/cm}^3) \\ &= 8.9762 = \mathbf{8.98 \text{ g/cm}^3} \end{aligned}$$

$$\begin{aligned} \text{b) Density of alloy} &= (62.0\% \text{ Cu} / 100\%) (8.95 \text{ g/cm}^3) + (38.0\% \text{ Zn} / 100\%) (65.41 / 63.55) (8.95 \text{ g/cm}^3) \\ &= 9.04954 = \mathbf{9.05 \text{ g/cm}^3} \end{aligned}$$

13.117 To find the volume of seawater needed, substitute the given information into the equation that describes the ppb concentration, account for extraction efficiency, and convert mass to volume using the density of seawater.

$$1.1 \times 10^{-2} \text{ ppb} = \frac{\text{mass Gold}}{\text{mass seawater}} \times 10^9$$

$$1.1 \times 10^{-2} \text{ ppb} = \frac{31.1 \text{ g Au}}{\text{mass seawater}} \times 10^9$$

$$\text{Mass seawater} = \left[\frac{31.1 \text{ g}}{1.1 \times 10^{-2}} \times 10^9 \right] = 2.827273 \times 10^{12} \text{ g (with 100\% efficiency)}$$

$$\text{Mass of seawater} = \left(2.827273 \times 10^{12} \text{ g} \right) \left(\frac{100\%}{81.5\%} \right) = 3.46905 \times 10^{12} \text{ g seawater (unrounded) (81.5\% efficiency)}$$

$$\text{Volume seawater} = \left(3.46905 \times 10^{12} \text{ g} \right) \left(\frac{1 \text{ mL}}{1.025 \text{ g}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 3.384439 \times 10^9 = \mathbf{3.4 \times 10^9 \text{ L}}$$

- 13.118 Xe is a much larger atom than He, so it is much more polarizable. This would increase the dipole-induced dipole forces when Xe is placed in water, increasing the solubility relative to He.
- 13.119 **C.** The principal factor in the solubility of ionic compounds in water is ion-dipole forces. Virtually all of the ionic compound's ions would become separated and surrounded by water molecules (the number depending on the sizes of the ions) interacting with the ions via H-bonding or other forces.
- 13.120 a) Solution A has a van't Hoff factor of 3, Solutions B and C have a van't Hoff factor of 2 and Solution D's van't Hoff factor is 1. Since Solution A has the largest van't Hoff factor, **Solution A** would have the highest boiling point.
 b) **Solution A** also has the lowest freezing point since it has the largest van't Hoff factor.
 c) **No**, the solution with the highest osmotic pressure cannot be determined. Osmotic pressure is determined by the molarity, not the molality, of the solution. Since we do not know the identity of the solutes and the density of the solutions, the 0.50 m value cannot be converted to molarity.
- 13.121 Convert mass of O₂ dissolved to moles of O₂. Use the density to convert the 1 kg mass of solution to volume in L.

0.0°C

$$\left(\frac{14.5 \text{ mg O}_2}{1 \text{ kg H}_2\text{O}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)\left(\frac{0.99987 \text{ g}}{\text{mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

$$= 4.53066 \times 10^{-4} = \mathbf{4.53 \times 10^{-4} \text{ M O}_2}$$

20.0°C

$$\left(\frac{9.07 \text{ mg O}_2}{1 \text{ kg H}_2\text{O}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)\left(\frac{0.99823 \text{ g}}{\text{mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

$$= 2.829358 \times 10^{-4} = \mathbf{2.83 \times 10^{-4} \text{ M O}_2}$$

40.0°C

$$\left(\frac{6.44 \text{ mg O}_2}{1 \text{ kg H}_2\text{O}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)\left(\frac{0.99224 \text{ g}}{\text{mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

$$= 1.996883 \times 10^{-4} = \mathbf{2.00 \times 10^{-4} \text{ M O}_2}$$

- 13.122 Pyridine has non-polar aromatic qualities like organic solvents but also has the potential to associate with water by hydrogen bonding through its lone pair of electrons (localized on the nitrogen atom).
- 13.123 a) First, find the molality from the freezing point depression and then use the molality, given mass of solute and volume of water to calculate the molar mass of the solute compound. Assume the solute is a nonelectrolyte ($i = 1$).

$$\Delta T_f = iK_f m = (0.000 - (-0.201)) = 0.201^\circ\text{C}$$

$$m = \frac{\Delta T_f}{K_f i} = \frac{0.201^\circ\text{C}}{(1.86^\circ\text{C}/m)(1)} = 0.1080645 \text{ m (unrounded)}$$

$$m = \frac{\text{moles solute}}{\text{kg solvent}} \quad (25.0 \text{ mL})\left(\frac{1.00 \text{ g}}{1 \text{ mL}}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 0.0250 \text{ kg water}$$

$$\text{moles solute} = (m)(\text{kg solvent}) = (0.1080656 \text{ m})(0.0250 \text{ kg}) = 0.0027016 \text{ mol}$$

$$\text{molar mass} = \frac{0.243 \text{ g}}{0.0027016 \text{ mol}} = 89.946698 = \mathbf{89.9 \text{ g/mol}}$$

b) Assume that 100.00 g of the compound gives 53.31 g carbon, 11.18 g hydrogen and $100.00 - 53.31 - 11.18 = 35.51$ g oxygen.

$$\text{Moles C} = (53.31 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 4.43880 \text{ mol C (unrounded)}; \quad \frac{4.43880}{2.219375} = 2$$

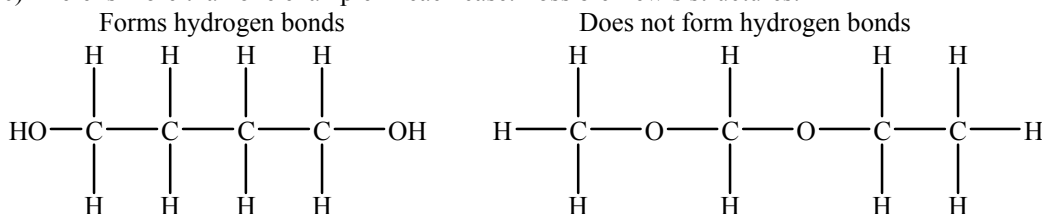
$$\text{Moles H} = (11.18 \text{ g H}) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 11.09127 \text{ mol H (unrounded)}; \quad \frac{11.09127}{2.219375} = 5$$

$$\text{Moles O} = (35.51 \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 2.219375 \text{ mol O (unrounded)}; \quad \frac{2.219375}{2.219375} = 1$$

Dividing the values by the lowest amount of moles (2.219375) gives an **empirical formula of $\text{C}_2\text{H}_5\text{O}$** with molar mass 45.06 g/mol.

Since the molar mass of the compound, 89.9 g/mol from part (a), is twice the molar mass of the empirical formula, the **molecular formula is $2(\text{C}_2\text{H}_5\text{O})$ or $\text{C}_4\text{H}_{10}\text{O}_2$** .

c) There is more than one example in each case. Possible Lewis structures:



13.124 Mass CO = $(4.0 \times 10^{-6} \text{ mol/L}) (11 \text{ L/min}) (60 \text{ min/1 h}) (8.0 \text{ h}) (28.01 \text{ g CO/mol}) = 0.59157 = \mathbf{0.59 \text{ g CO}}$

13.125 **No**, both are the same because masses are additive.

13.126 a) $X_{\text{N}_2} = \frac{\text{moles of N}_2}{\text{total moles}}$

Mixture A: $X_{\text{N}_2} = \frac{3 \text{ moles N}_2}{3 + 2 + 3 \text{ moles}} = 0.38$

Mixture B: $X_{\text{N}_2} = \frac{4 \text{ moles N}_2}{4 + 4 + 2 \text{ moles}} = 0.40$

Mixture C: $X_{\text{N}_2} = \frac{4 \text{ moles N}_2}{4 + 3 + 5 \text{ moles}} = 0.33$

Mixture C has the smallest mole fraction of N_2 .

b) Mixture A: $X_{\text{Ne}} = \frac{2 \text{ moles Ne}}{3 + 2 + 3 \text{ moles}} = 0.25$

Mixture B: $X_{\text{Ne}} = \frac{4 \text{ moles Ne}}{4 + 4 + 2 \text{ moles}} = 0.40$

Mixture C: $X_{\text{Ne}} = \frac{3 \text{ moles Ne}}{4 + 3 + 5 \text{ moles}} = 0.25$

Mixtures A and C have the same mole fraction of Ne.

c) Mixture A: $X_{\text{Cl}_2} = \frac{3 \text{ moles Cl}_2}{3 + 2 + 3 \text{ moles}} = 0.38$

Mixture B: $X_{\text{Cl}_2} = \frac{2 \text{ moles Cl}_2}{4 + 4 + 2 \text{ moles}} = 0.20$

Mixture C: $X_{\text{Cl}_2} = \frac{5 \text{ moles Cl}_2}{4 + 3 + 5 \text{ moles}} = 0.42$

Mixture B < Mixture A < Mixture C

3.127 a) $\left(\frac{4.50 \times 10^{-5} \text{ mol F}^-}{\text{L}} \right) (5000. \text{ L}) \left(\frac{1 \text{ mol NaF}}{1 \text{ mol F}^-} \right) \left(\frac{41.99 \text{ g NaF}}{1 \text{ mol NaF}} \right) = 9.44775 = \mathbf{9.45 \text{ g NaF}}$

b) $\left(\frac{4.50 \times 10^{-5} \text{ mol F}^-}{\text{L}} \right) (2.0 \text{ L}) \left(\frac{19.00 \text{ g F}^-}{1 \text{ mol F}^-} \right) = \mathbf{0.0017 \text{ g F}^-}$

- 13.128 a) To shorten the settling time, lime (CaO) and cake alum (Al₂(SO₄)₃) are added to form a fluffy, gel-like precipitate of Al(OH)₃.
 b) Water that contains large amounts of divalent cations (such as Ca²⁺, Mg²⁺, and Fe²⁺) is called hard water. During cleaning, these ions combine with the fatty-acid anions in soaps to produce insoluble deposits.
 c) In reverse osmosis, a higher pressure is applied to the solution, forcing the water back through the membrane and leaving the ions behind.
 d) Chlorine may give the water an unpleasant odor, and can form toxic chlorinated hydrocarbons.
 e) The high concentration of NaCl displaces the divalent and polyvalent ions from the ion-exchange resin.
- 13.129 a) The solution in U-tube **B** is the most concentrated since it has the highest osmotic pressure.
 b) Solution **C** has the smallest number of dissolved ions and thus the smallest osmotic pressure.
- 13.130 Calculate the individual partial pressures from $P = X P^\circ$. Assign the "equal masses" as exactly 1 g.
 Liquid:

$$X(\text{pinene}) = \frac{\left(\frac{1 \text{ g pinene}}{136.23 \text{ g pinene/mol}}\right)}{\left(\frac{1 \text{ g pinene}}{136.23 \text{ g pinene/mol}}\right) + \left(\frac{1 \text{ g terpineol}}{154.24 \text{ g terpineol/mol}}\right)} = 0.53100 \text{ (unrounded)}$$

$$P(\text{pinene}) = (0.53100)(100.3 \text{ torr}) = 53.2593 \text{ torr (unrounded)}$$

$$X(\text{terpineol}) = \frac{\left(\frac{1 \text{ g terpineol}}{154.24 \text{ g terpineol/mol}}\right)}{\left(\frac{1 \text{ g pinene}}{136.23 \text{ g pinene/mol}}\right) + \left(\frac{1 \text{ g terpineol}}{154.24 \text{ g terpineol/mol}}\right)} = 0.4689985 \text{ (unrounded)}$$

$$P(\text{terpineol}) = (0.4689985)(9.8 \text{ torr}) = 4.5961855 \text{ torr (unrounded)}$$

Vapor:

$$X(\text{pinene}) = \frac{53.2593 \text{ torr}}{(53.2593 + 4.5961855) \text{ torr}} = 0.9205575 = \mathbf{0.921}$$

$$X(\text{terpineol}) = \frac{4.5961855 \text{ torr}}{(53.2593 + 4.5961855) \text{ torr}} = 0.0794425 = \mathbf{0.079}$$

- 13.131 a) Use the boiling point elevation of 0.45°C to calculate the molality of the solution. Then, with molality, the mass of solute, and volume of water calculate the molar mass.

$$\Delta T = iK_b m \quad i = 1 \text{ (nonelectrolyte)}$$

$$\Delta T = (100.45 - 100.00)^\circ\text{C} = 0.45^\circ\text{C}$$

$$m = \frac{\Delta T_b}{K_b i} = \frac{0.45^\circ\text{C}}{(0.512^\circ\text{C}/m)(1)} = 0.878906 \text{ } m = 0.878906 \text{ mol/kg (unrounded)}$$

$$m = \frac{\text{moles solute}}{\text{kg solvent}} \quad (25.0 \text{ mL}) \left(\frac{0.997 \text{ g}}{1 \text{ mL}}\right) \left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 0.0249250 \text{ kg water}$$

$$\text{moles solute} = (m)(\text{kg solvent}) = (0.878906 \text{ } m)(0.0249250 \text{ kg}) = 0.0219067 \text{ mol}$$

$$\text{molar mass} = \frac{1.50 \text{ g}}{0.0219067 \text{ mol}} = 68.4722 = \mathbf{68 \text{ g/mol}}$$

b) The molality calculated would be the moles of ions per kg of solvent. If the compound consists of three ions the molality of the compound would be 1/3 of 0.878906 *m* and the calculated molar mass would be three times greater: 3 x 68.4722 = 205.417 = **2.1 x 10² g/mol**.

c) The molar mass of CaN₂O₆ is 164.10 g/mol. This molar mass is less than the 2.1 x 10² g/mol calculated when the compound is assumed to be a strong electrolyte and is greater than the 68 g/mol calculated when the compound is assumed to be a nonelectrolyte. Thus, the compound is an electrolyte, since it dissociates into ions in solution. However, the ions do not dissociate completely in solution.

d) Use the molar mass of CaN_2O_6 to calculate the molality of the compound. Then calculate i in the boiling point elevation formula.

$$m = \left(\frac{1.50 \text{ g CaN}_2\text{O}_6}{25.0 \text{ mL}} \right) \left(\frac{\text{mL}}{0.997 \text{ g}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol}}{164.10 \text{ g CaN}_2\text{O}_6} \right) = 0.3667309 \text{ m (unrounded)}$$

$$\Delta T = iK_b m$$

$$i = \frac{\Delta T_b}{K_b m} = \frac{(0.45^\circ\text{C})}{(0.512^\circ/\text{m})(0.3667309 \text{ m})} = 2.396597 = \mathbf{2.4}$$

$$13.132 \quad \frac{\text{mol C}_2\text{H}_5\text{OH(g)}}{\text{mol CH}_3\text{OH(g)}} = \frac{\text{mol C}_2\text{H}_5\text{OH(l)}}{\text{mol CH}_3\text{OH(l)}} \left(\frac{60.5 \text{ torr}}{126.0 \text{ torr}} \right) = \frac{\text{mol C}_2\text{H}_5\text{OH(l)}}{\text{mol CH}_3\text{OH(l)}} (0.4801587) \text{ (unrounded)}$$

A 97:1 mass ratio gives 97 grams of $\text{C}_2\text{H}_5\text{OH}$ for every 1 gram of CH_3OH . (This limits the significant figures.)

$$\frac{97 \text{ g C}_2\text{H}_5\text{OH(g)} \left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.07 \text{ g C}_2\text{H}_5\text{OH}} \right)}{1 \text{ g CH}_3\text{OH(g)} \left(\frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} \right)} = \frac{2.10549 \text{ mol C}_2\text{H}_5\text{OH(g)}}{0.03121 \text{ mol CH}_3\text{OH(g)}}$$

$$\frac{\text{mol C}_2\text{H}_5\text{OH(l)}}{\text{mol CH}_3\text{OH(l)}} = \frac{(2.10549/0.03121)}{0.4801587} = 140.4994$$

$$\frac{(140.4994 \text{ mol C}_2\text{H}_5\text{OH}) \left(\frac{46.07 \text{ g C}_2\text{H}_5\text{OH}}{1 \text{ mol C}_2\text{H}_5\text{OH}} \right)}{(1 \text{ mol CH}_3\text{OH}) \left(\frac{32.04 \text{ g CH}_3\text{OH}}{1 \text{ mol CH}_3\text{OH}} \right)} = 202.0227 = \mathbf{2 \times 10^2}$$

13.133 Convert from ppb to pph (part per hundred = mass percent)

$$\left(\frac{100. \text{ ppb}}{10^9} \right) \left(\frac{100 \text{ pph}}{1} \right) = \mathbf{1.00 \times 10^{-5}\%}$$

Determine the molarity of CH_3Cl in 1.00 L corresponding to 100. ppb. (Assume the density of the solution is the same as for pure water, 1.00 g/mL.)

$$\left(\frac{100. \text{ g CH}_3\text{Cl}}{10^9 \text{ g Solution}} \right) \left(\frac{1 \text{ mol CH}_3\text{Cl}}{50.48 \text{ g CH}_3\text{Cl}} \right) \left(\frac{1.00 \text{ g Solution}}{\text{mL}} \right) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)$$

$$= 1.98098 \times 10^{-6} = \mathbf{1.98 \times 10^{-6} \text{ M CH}_3\text{Cl}}$$

If the density is 1.00 g/mL then 1.00 L of solution would weigh 1.00 kg. The mass of CH_3Cl is insignificant compared to 1.00 kg, thus the mass of the solution may be taken as the mass of the solvent. This makes the molarity equal to the molality, in other words: $\mathbf{1.98 \times 10^{-6} \text{ m CH}_3\text{Cl}}$

Still using 1.00 L of solution:

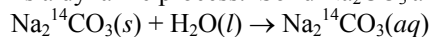
$$\text{Moles CH}_3\text{Cl} = (1.98098 \times 10^{-6} \text{ mol/L}) (1.00 \text{ L}) = 1.98098 \times 10^{-6} \text{ mol CH}_3\text{Cl (unrounded)}$$

$$\text{Moles H}_2\text{O} = (1.00 \text{ kg}) (10^3 \text{ g/1 kg}) (1 \text{ mol H}_2\text{O} / 18.02 \text{ g H}_2\text{O}) = 55.49389567 \text{ mol H}_2\text{O (unrounded)}$$

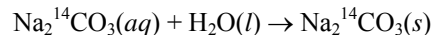
$$X_{\text{chloroform}} = (1.98098 \times 10^{-6} \text{ mol CH}_3\text{Cl}) / [(1.98098 \times 10^{-6} + 55.49389567) \text{ mol}]$$

$$= 3.569726 \times 10^{-8} = \mathbf{3.57 \times 10^{-8}}$$

13.134 a) **Yes**, equilibrium is a dynamic process. Solid $\text{Na}_2^{14}\text{CO}_3$ and solid $\text{Na}_2^{14}\text{CO}_3$ both dissolve in the equilibrium process.



b) Radioactivity would be found in all of the solid as some of the Na^{14} that dissolves will also precipitate back out of solution.



- 13.135 a) From the osmotic pressure, the molarity of the solution can be found. The ratio of mass per volume to moles per volume gives the molar mass of the compound.

$$\Pi = MRT$$

$$M = \frac{\Pi}{RT} = \frac{(0.340 \text{ torr})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)\left((273 + 25)\text{K}\right)} \left(\frac{1 \text{ atm}}{760 \text{ torr}}\right) = 1.828546 \times 10^{-5} M \text{ (unrounded)}$$

$$(M)(V) = \text{moles}$$

$$\text{Moles} = (1.828546 \times 10^{-5} M)(30.0 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 5.48564 \times 10^{-7} \text{ mol}$$

$$\text{Molar mass} = \frac{(10.0 \text{ mg}) \left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)}{5.48564 \times 10^{-7} \text{ mol}} = 1.82294 \times 10^4 = \mathbf{1.82 \times 10^4 \text{ g/mol}}$$

- b) To find the freezing point depression, the molarity of the solution must be converted to molality. Then use $\Delta T_f = iK_f m$. ($i = 1$)

Mass solvent = mass of solution – mass of solute

$$\text{Mass solvent} = \left[(30.0 \text{ mL}) \left(\frac{0.997 \text{ g}}{1 \text{ mL}}\right) - (10.0 \text{ mg}) \left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right) \right] \left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right) = 0.0299 \text{ kg}$$

$$\text{Moles solute} = \left(\frac{1.828546 \times 10^{-5} \text{ mol}}{\text{L}}\right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) (30.0 \text{ mL}) = 5.485638 \times 10^{-7} \text{ mol (unrounded)}$$

$$\text{Molality} = \frac{5.485638 \times 10^{-7} \text{ mol}}{0.0299 \text{ kg}} = 1.83466 \times 10^{-5} m \text{ (unrounded)}$$

$$\Delta T_f = iK_f m = (1)(1.86^\circ\text{C}/m)(1.83466 \times 10^{-5} m) = 3.412 \times 10^{-5} = \mathbf{3.41 \times 10^{-5}^\circ\text{C}}$$

(So the solution would freeze at $0 - (3.41 \times 10^{-5}^\circ\text{C}) = -3.41 \times 10^{-5}^\circ\text{C}$.)

- 13.136 Henry's law expresses the relationship between gas pressure and the gas solubility (S_{gas}) in a given solvent. Use Henry's law to solve for pressure (assume that the constant (k_H) is given at 21°C), use the ideal gas law to find moles per unit volume, and convert moles/L to ng/L.

$$S_{\text{gas}} = k_H P_{\text{gas}}$$

$$P_{\text{gas}} = S_{\text{gas}} / k_H = \left(\frac{0.65 \text{ mg/L}}{0.033 \text{ mol/L}\cdot\text{atm}}\right) \left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right) \left(\frac{1 \text{ mol C}_2\text{H}_2\text{Cl}_2}{96.94 \text{ g C}_2\text{H}_2\text{Cl}_2}\right)$$

$$= 2.3446799 \times 10^{-4} \text{ atm (unrounded)}$$

$$PV = nRT$$

$$n / V = P / RT = \frac{(2.3446799 \times 10^{-4} \text{ atm})}{\left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)\left((273 + 21)\text{K}\right)} = 9.7130 \times 10^{-6} \text{ mol/L (unrounded)}$$

$$\left(\frac{9.7130 \times 10^{-6} \text{ mol C}_2\text{H}_2\text{Cl}_2}{\text{L}}\right) \left(\frac{96.94 \text{ g C}_2\text{H}_2\text{Cl}_2}{1 \text{ mol C}_2\text{H}_2\text{Cl}_2}\right) \left(\frac{1 \text{ ng}}{10^{-9} \text{ g}}\right) = 9.41578 \times 10^5 = \mathbf{9.4 \times 10^5 \text{ ng/L}}$$

- 13.137 $A_{\text{circle}} = \pi r^2 = \pi(38.6 \text{ cm} / 2)^2 = 1.17021 \times 10^3 \text{ cm}^2$ (unrounded)

$$\text{Area of one molecule} = \frac{1.17021 \times 10^3 \text{ cm}^2}{2.50 \text{ mg}} \left(\frac{1 \text{ mg}}{10^{-3} \text{ g}}\right) \left(\frac{283 \text{ g}}{\text{mol}}\right) \left(\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}}\right)$$

$$= 2.19973 \times 10^{-16} = \mathbf{2.20 \times 10^{-16} \text{ cm}^2/\text{molecule}}$$

- 13.138 a) Looking at the data for CaCl_2 , K_2CO_3 , and Na_2SO_4 , the average conductivity is 7.0 ± 0.7 units for the 5.00×10^3 ppm solutions and 14 ± 1.7 units for the 10.00×10^3 ppm solutions. This represents a relative error of about 10% if you assume that the identity of the solute is immaterial. If your application can tolerate an error of this magnitude, then this method would be acceptable.
- b) This would be an unreliable estimate of the concentration for those substances which are non-electrolytes, or weak electrolytes, as their conductivity would be much reduced in comparison to their true concentration.
- c) Concentration (ppm) = $(14.0 / 16.0) (10.00 \times 10^3 \text{ ppm}) = 8.75 \times 10^3 \text{ ppm}$
 Assume the mass of CaCl_2 present is negligible relative to the mass of the solution.

$$\text{Molality CaCl}_2 = \left(\frac{8.75 \times 10^3 \text{ g CaCl}_2}{10^6 \text{ g Solution}} \right) \left(\frac{1 \text{ mol CaCl}_2}{110.98 \text{ g CaCl}_2} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) = 0.0788430 = \mathbf{0.0788 \text{ m CaCl}_2}$$

$$\text{Moles CaCl}_2 = (8.75 \times 10^3 \text{ g CaCl}_2) (1 \text{ mol CaCl}_2 / 110.98 \text{ g CaCl}_2) = 78.84303 \text{ mol CaCl}_2 \text{ (unrounded)}$$

$$\text{Moles H}_2\text{O} = (1.00 \times 10^6 \text{ g H}_2\text{O}) (1 \text{ mol H}_2\text{O} / 18.02 \text{ g H}_2\text{O}) = 5.5493896 \times 10^4 \text{ mol H}_2\text{O} \text{ (unrounded)}$$

$$\text{Mole fraction CaCl}_2 = X = \frac{(78.84303 \text{ mol CaCl}_2)}{(78.84303 + 5.5493896 \times 10^4) \text{ mol}} = 1.4187 \times 10^{-3} = \mathbf{1.42 \times 10^{-3}}$$

- 13.139 The vapor pressure of H_2O above the pure water is greater than that above the sugar solution. This means that water molecules will leave the pure water and enter the sugar solution in order to make their vapor pressures closer to equal.

- 13.140 a) Assume a concentration of 1 mol/m³ for both ethanol and 2-butoxyethanol in the detergent solution. Then, from Henry's Law, the partial pressures of the two substances can be calculated:

$$S_{\text{gas}} = k_H \times P_{\text{gas}}$$

$$P_{\text{gas}} = S_{\text{gas}} / k_H$$

$$P_{\text{ethanol}} = \left(\frac{1 \text{ mol}}{\text{m}^3} \right) \left(\frac{5 \times 10^{-6} \text{ atm} \cdot \text{m}^3}{\text{mol}} \right) = 5 \times 10^{-6} \text{ atm}$$

$$P_{\text{2-butoxyethanol}} = \left(\frac{1 \text{ mol}}{\text{m}^3} \right) \left(\frac{1.6 \times 10^{-6} \text{ atm} \cdot \text{m}^3}{\text{mol}} \right) = 1.6 \times 10^{-6} \text{ atm}$$

$$\%_{\text{2-butoxyethanol}} = (5\%) \left(\frac{1.6 \times 10^{-6} \text{ atm 2-butoxyethanol}}{5 \times 10^{-6} \text{ atm ethanol}} \right) = 1.6\%$$

“Down-the-drain” factor is $0.016 = \mathbf{0.02}$

$$\text{b) } k_{\text{Hethanol}} = \left(\frac{5 \times 10^{-6} \text{ atm} \cdot \text{m}^3}{\text{mol}} \right) \left(\frac{1 \text{ L}}{10^{-3} \text{ m}^3} \right) = \mathbf{5 \times 10^{-3} \text{ atm} \cdot \text{L/mol}}$$

$$\text{c) } k_{\text{Hethanol}} = \left(\frac{0.64 \text{ Pa} \cdot \text{m}^3}{\text{mol}} \right) \left(\frac{1 \text{ atm}}{1.01325 \times 10^5 \text{ Pa}} \right) = \mathbf{6.3 \times 10^{-6} \text{ atm} \cdot \text{m}^3/\text{mol}}$$

Considering the single significant figure in the measured value of 5×10^{-6} , the agreement is good.

- 13.141 The fraction remaining in the water (f_w) is related to the volume of water (V_w), the volume of dichloromethane (V_d), and the distribution ratio for the solubility ($D = 8.35 / 1$).

$$f_w = V_w / (V_w + DV_d)$$

Mass remaining in water = f_w (original mass)

$$\text{a) Mass in water} = \frac{(100.0 \text{ mL})}{(100.0 + 8.35(60.0)) \text{ mL}} (10.0 \text{ mg}) = 1.66389 = \mathbf{1.66 \text{ mg}} \text{ remaining}$$

b) Perform a similar calculation to part (a), then take the result and repeat the procedure. Combine the results to get the total removed.

$$\text{Mass in water} = \frac{(100.0 \text{ mL})}{(100.0 + 8.35(30.0)) \text{ mL}} (10.0 \text{ mg}) = 2.853067 \text{ mg remaining after first extraction}$$

$$\begin{aligned} \text{Mass in water} &= \frac{(100.0 \text{ mL})}{(100.0 + 8.35(30.0)) \text{ mL}} (2.853067 \text{ mg}) \\ &= 0.813999 = \mathbf{0.814 \text{ mg}} \text{ remaining after second extraction} \end{aligned}$$

c) The two-step extraction extracts more of the caffeine.

13.142 Molality is defined as moles of solute per kg of solvent, so 0.150 *m* means 0.150 mol NaHCO₃ per kg of water. The total mass of the solution would be 1 kg + (0.150 mol x molar mass of NaHCO₃).

$$\begin{aligned} 0.150 \text{ m} &= (0.150 \text{ mol NaHCO}_3) / (1 \text{ kg solvent}) = \frac{(0.150 \text{ mol NaHCO}_3) \left(\frac{84.01 \text{ g NaHCO}_3}{1 \text{ mol NaHCO}_3} \right)}{1 \text{ kg}} \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) \\ &= 12.6015 \text{ g NaHCO}_3 / 1000 \text{ g solvent} \end{aligned}$$

$$\left(\frac{12.6015 \text{ g NaHCO}_3}{(1000 + 12.6015) \text{ g Solution}} \right) (250. \text{ g Solution}) = 3.111 \text{ g NaHCO}_3 \text{ (unrounded)}$$

$$\text{Grams H}_2\text{O} = 250. \text{ g} - 3.111 \text{ g} = 246.889 \text{ g H}_2\text{O}$$

To make 250. g of a 0.150 *m* solution of NaHCO₃, **weigh 3.11 g NaHCO₃ and dissolve in 247 g water.**

13.143 To determine the molecular formula, both the empirical formula and the molar mass are needed. First, determine the empirical formula assuming exactly 100 grams of sample, which makes the percentages equal to the mass of each element present:

$$\text{Moles C} = 32.3 \text{ g C} (1 \text{ mol C} / 12.01 \text{ g C}) = 2.6894 \text{ mol C (unrounded)}$$

$$\text{Moles H} = 3.97 \text{ g H} (1 \text{ mol H} / 1.008 \text{ g H}) = 3.93849 \text{ mol H (unrounded)}$$

$$\text{Moles O} = (100 - 32.3 - 3.97) \text{ g O} (1 \text{ mol O} / 16.00 \text{ g O}) = 3.9831 \text{ mol O (unrounded)}$$

Dividing each mole value by the smallest value (moles C) gives: C = 1, H = 1.5, and O = 1.5 leading to an empirical formula of: **C₂H₃O₃**.

The molar mass comes from the freezing point depression:

$$\Delta T_f = iK_f m \quad (\text{Assume the compound is a nonelectrolyte, } i = 1.)$$

$$m = \Delta T_f / iK_f = (1.26^\circ\text{C}) / [(1) (1.86^\circ\text{C}/m)] = 0.677419 \text{ m (unrounded)}$$

$$\text{Molar mass} = \left(\frac{\text{kg Solvent}}{0.677419 \text{ mol}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{0.981 \text{ g}}{11.23 \text{ g Solvent}} \right) = 128.953 \text{ g/mol (unrounded)}$$

The empirical formula mass is approximately 75 g/mol.

The ratio of the molar to the empirical formula mass normally gives the conversion factor to change the empirical formula to the molecular formula. In this case, 129 / 75 = 1.72; this is not near a whole number. (This result is low due to dissociation of the weak acid; the assumption of *i* = 1 is too low. If *i* = 1.2, then the molar mass would increase to about 154 g/mol.) The 1.72 value implies the molecular formula is twice the empirical formula, or **C₄H₆O₆**.

13.144 The range has to fall between the point where the number of moles of methanol is just greater than the number of moles of ethanol, to the point where the mass of methanol is just less than the mass of ethanol.

The first point is the point at which the mole fractions are just becoming unequal. The methanol mole fraction is greater than 0.5000.

Point 2: where the mass percents are just beginning to become unequal.

First, find where they are equal.

$$(1.000 \text{ g methanol} / 2.000 \text{ g solution}) = (1.000 \text{ g ethanol} / 2.000 \text{ g solution})$$

$$\text{Moles methanol} = (1.000 \text{ g ethanol}) (1 \text{ mol ethanol} / 32.04 \text{ g ethanol}) = 0.031210986 \text{ mol methanol (unrounded)}$$

$$\text{Moles ethanol} = (1.000 \text{ g ethanol}) (1 \text{ mol ethanol} / 46.07 \text{ g ethanol}) = 0.021706 \text{ mol ethanol (unrounded)}$$

$$\text{Mole fraction methanol} = (0.031210986 \text{ mol methanol}) / [(0.031210986) + (0.021706)] \text{ mol} \\ = 0.589810 \text{ (unrounded)}$$

$$\text{Range of mole fractions of methanol: } 0.5000 < X_{\text{methanol}} < 0.5897$$

- 13.145 a) The molar mass comes from the boiling point elevation:
The boiling point and elevation constant values come from Table 13.6.

$$\Delta T_b = (77.5 - 76.5) = 1.0^\circ\text{C}$$

$$\Delta T_b = iK_b m \quad (\text{Assume the compound is a nonelectrolyte, } i = 1.)$$

$$m = \Delta T_b / iK_b = (1.0^\circ\text{C}) / [(1) (5.03^\circ\text{C}/m)] = 0.198807 \text{ } m \text{ (unrounded)}$$

$$\text{Molar mass} = \left(\frac{\text{kg Solvent}}{0.198807 \text{ mol}} \right) \left(\frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left(\frac{5.0 \text{ g}}{100.0 \text{ g Solvent}} \right) = 251.5 = \mathbf{2.5 \times 10^2 \text{ g/mol}}$$

b) The molar mass, based on the formula, is 122.12 g/mol. The molar mass determined in part (a) is double the actual molar mass. This is because the acid dimerizes (forms pairs) in the solution. These pairs are held together by relatively strong hydrogen bonds, and give a “molecule” that is double the mass of a normal molecule.

- 13.146 Molarity is moles solute/L solution and molality is moles solute/kg solvent.
Multiplying molality by concentration of solvent in kg solvent per liter of solution gives molarity:

$$\frac{\text{mol of solute}}{\text{L of solution}} = \left(\frac{\text{mol of solute}}{\text{kg solvent}} \right) \left(\frac{\text{kg solvent}}{\text{L of solution}} \right)$$

$$M = m \left(\frac{\text{kg solvent}}{\text{L of solution}} \right)$$

For a very dilute solution, the assumption that mass of solvent \cong mass of solution is valid. This equation then becomes

$$M = m (\text{kg solvent/L solution}) = m \times d_{\text{solution}}$$

Thus, for very dilute solutions **molality x density = molarity**.

In an aqueous solution, the liters of solution have approximately the same value as the kg of solvent because the density of water is close to 1 kg/L, so $m = M$.

$$13.147 \text{ Moles} = (5.66 \text{ g } \text{NH}_4\text{NO}_3) \left(\frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{80.05 \text{ g } \text{NH}_4\text{NO}_3} \right) \left(\frac{1 \text{ mol } \text{NH}_4^+}{1 \text{ mol } \text{NH}_4\text{NO}_3} \right) = 7.07058 \times 10^{-2} \text{ mol } \text{NH}_4^+ \text{ (unrounded)}$$

$$\text{Moles} = (4.42 \text{ g } (\text{NH}_4)_3\text{PO}_4) \left(\frac{1 \text{ mol } (\text{NH}_4)_3\text{PO}_4}{149.10 \text{ g } (\text{NH}_4)_3\text{PO}_4} \right) \left(\frac{3 \text{ mol } \text{NH}_4^+}{1 \text{ mol } (\text{NH}_4)_3\text{PO}_4} \right) \\ = 8.89336 \times 10^{-2} \text{ mol } \text{NH}_4^+ \text{ (unrounded)}$$

$$\text{Moles} = (4.42 \text{ g } (\text{NH}_4)_3\text{PO}_4) \left(\frac{1 \text{ mol } (\text{NH}_4)_3\text{PO}_4}{149.10 \text{ g } (\text{NH}_4)_3\text{PO}_4} \right) \left(\frac{1 \text{ mol } \text{PO}_4^{3-}}{1 \text{ mol } (\text{NH}_4)_3\text{PO}_4} \right) \\ = 2.96445 \times 10^{-2} \text{ mol } \text{PO}_4^{3-} \text{ (unrounded)}$$

$$M \text{ NH}_4^+ = [(7.07058 \times 10^{-2}) + (8.89336 \times 10^{-2})] \text{ mol } \text{NH}_4^+ / 20.0 \text{ L} = 7.98197 \times 10^{-3} = \mathbf{7.98 \times 10^{-3} \text{ } M \text{ NH}_4^+}$$

$$M \text{ PO}_4^{3-} = (2.96445 \times 10^{-2} \text{ mol } \text{PO}_4^{3-}) / 20.0 \text{ L} = 1.482225 \times 10^{-3} = \mathbf{1.48 \times 10^{-3} \text{ } M \text{ PO}_4^{3-}}$$

- 13.148 a) $M \text{ SO}_2 = P \text{ } K_H = (2.0 \times 10^{-3} \text{ atm}) (1.23 \text{ mol/L}\cdot\text{atm}) = 2.46 \times 10^{-3} = \mathbf{2.5 \times 10^{-3} \text{ } M \text{ SO}_2}$

b) The base reacts with the sulfur dioxide to produce calcium sulfite. The reaction of sulfur dioxide makes “room” for more sulfur dioxide to dissolve.

- 13.149 a) Assume a 100 g sample of urea. This leads to the mass of each element being equal to the percent of that element.

$$\text{Moles C} = (20.1 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 1.6736 \text{ mol C (unrounded)}$$

$$\text{Moles H} = (6.7 \text{ g H}) \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 6.6468 \text{ mol H (unrounded)}$$

$$\text{Moles N} = (46.5 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 3.31906 \text{ mol N (unrounded)}$$

$$\text{Moles O} = ((100 - 20.1 - 6.7 - 46.5) \text{ g O}) \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 1.66875 \text{ mol O (unrounded)}$$

Dividing all by the smallest value (1.66875 mol O) gives: C = 1, H = 4, N = 2, O = 1. Thus, the empirical formula is **CH₄N₂O**. The empirical formula weight is 60.06 g/mol.

b) Use $\Pi = MRT$ to solve for the molarity of the urea solution. The solution molarity is related to the concentration expressed in % w/v by using the molar mass.

$$M = \Pi / RT = \frac{(2.04 \text{ atm})}{\left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) ((273 + 25) \text{ K})} = 0.0833817 \text{ M (unrounded)}$$

$$\text{Molar mass} = \frac{\left(\frac{5.0 \text{ g}}{\text{L}} \right)}{\left(\frac{0.0833817 \text{ mol}}{\text{L}} \right)} = 59.965 = \mathbf{60. \text{ g/mol}}$$

Because the molecular weight equals the empirical weight, the molecular formula is also **CH₄N₂O**.

13.150 a) Mass glucose = $(2.5 \text{ h}) \left(\frac{100. \text{ mL}}{\text{h}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.30 \text{ mol Glucose}}{1 \text{ L}} \right) \left(\frac{180.16 \text{ g Glucose}}{1 \text{ mol Glucose}} \right)$
 $= 13.512 = \mathbf{14 \text{ g glucose}}$

b) At low concentrations sodium chloride dissociates completely, forming twice as many dissolved particles per mole as glucose, so a sodium chloride solution would have to have a molarity that is one-half of glucose to be isotonic: 0.15 M

c) Mass NaCl = $(1.5 \text{ h}) \left(\frac{150. \text{ mL}}{\text{h}} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{0.15 \text{ mol NaCl}}{1 \text{ L}} \right) \left(\frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} \right)$
 $= 1.97235 = \mathbf{2.0 \text{ g NaCl}}$

- 13.151 a) There is a **positive deviation** since benzene molecules are held together only by weak dispersion forces while methanol molecules are held together by relatively strong hydrogen bonding. The two components will not interact with each other since the intermolecular forces are so different.
- b) There is a **positive deviation** since ethyl acetate will have weaker hydrogen bonding to water than water has with itself.
- c) Since hexane and heptane are very similar compounds with weak dispersion forces, they will obey Raoult's Law. The behavior will be **nearly ideal**.
- d) The behavior will be **nearly ideal** since the hydrogen bonding in methanol and water is very similar.
- e) There is a **negative deviation** because HCl exists as ions in solution and water is in the hydration shells around the H₃O⁺ and Cl⁻ ions.

13.152 a) Use Henry's Law:

$$S_{\text{gas}} = k_{\text{H}} \times P_{\text{gas}}$$

$$k_{\text{H}} = S_{\text{gas}} / P_{\text{gas}}$$

$$k_{\text{H}} = 1.64 \times 10^{-3} \text{ M}/0.2095 \text{ atm}$$

$$k_{\text{H}} = 7.8282 \times 10^{-3} = \mathbf{7.83 \times 10^{-3} \text{ mol/L}\cdot\text{atm}}$$

b) $S_{\text{gas}} = k_{\text{H}} \times P_{\text{gas}}$

$$S_{\text{gas}} = (7.8282 \times 10^{-3} \text{ mol/L}\cdot\text{atm})(0.005 \text{ atm})$$

$$S_{\text{gas}} = 3.9141 \times 10^{-5} = \mathbf{4 \times 10^{-5} \text{ M}}$$

c) Assume a 1.0 L sample. Acrylic acid is 14.6 mol/L or 14.6 mol in 1.0 L.

Oxygen is 4×10^{-5} mol/L or 4×10^{-5} mol in 1.0 L.

$$X_{\text{O}_2} = \frac{\text{moles of O}_2}{\text{moles of O}_2 + \text{moles of acrylic acid}} = \frac{4 \times 10^{-5} \text{ mol}}{(4 \times 10^{-5}) + 14.6 \text{ mol}} = 2.73972 \times 10^{-6} = \mathbf{3 \times 10^{-6}}$$

d) Mass of acrylic acid = $\left(\frac{14.6 \text{ mol acrylic acid}}{\text{L}}\right)\left(\frac{72.06 \text{ g acrylic acid}}{1 \text{ mol acrylic acid}}\right) = 1052.076 \text{ g/L}$

$$\text{Mass of oxygen} = \left(\frac{4 \times 10^{-5} \text{ mol O}_2}{\text{L}}\right)\left(\frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2}\right) = 0.00128 \text{ g/L}$$

$$\text{pm} = \frac{\text{mass of solute}}{\text{mass of solution}} (1 \times 10^6) = \frac{0.00128 \text{ g}}{0.00128 \text{ g} + 1052.176 \text{ g}} (1 \times 10^6) = 1.2165 = \mathbf{1 \text{ ppm}}$$

13.153 Mass percents:

Iodine in chloroform

$$[2.7 \text{ g I}_2 / (2.7 + 100.0)\text{g}] \times 100\% = 2.6290 = \mathbf{2.6\% \text{ I}_2}$$

Iodine in carbon tetrachloride

$$[2.5 \text{ g I}_2 / (2.5 + 100.0)\text{g}] \times 100\% = 2.4390 = \mathbf{2.4\% \text{ I}_2}$$

Iodine in carbon disulfide

$$[16 \text{ g I}_2 / (16 + 100.0)\text{g}] \times 100\% = 13.793 = \mathbf{14\% \text{ I}_2}$$

Mole fraction:

Iodine in chloroform

$$\text{Moles I}_2 = (2.7 \text{ g I}_2) (1 \text{ mol I}_2 / 253.8 \text{ g I}_2) = 0.010638 \text{ mol I}_2 \text{ (unrounded)}$$

$$\text{Moles solvent} = (100.0 \text{ g CHCl}_3) (1 \text{ mol CHCl}_3 / 119.37 \text{ g CHCl}_3)$$

$$= 0.8377314 \text{ mol CHCl}_3 \text{ (unrounded)}$$

$$\text{Mole fraction} = (0.010638 \text{ mol I}_2) / [(0.010638) + (0.8377314)] \text{ mol} = 0.01245966 = \mathbf{0.012}$$

Iodine in carbon tetrachloride

$$\text{Moles I}_2 = (2.5 \text{ g I}_2) (1 \text{ mol I}_2 / 253.8 \text{ g I}_2) = 0.009850 \text{ mol I}_2 \text{ (unrounded)}$$

$$\text{Moles solvent} = (100.0 \text{ g CCl}_4) (1 \text{ mol CCl}_4 / 153.81 \text{ g CCl}_4)$$

$$= 0.650152785 \text{ mol CCl}_4 \text{ (unrounded)}$$

$$\text{Mole fraction} = (0.009850 \text{ mol I}_2) / [(0.009850) + (0.650152785)] \text{ mol} = 0.014924 = \mathbf{0.015}$$

Iodine in carbon disulfide

$$\text{Moles I}_2 = (16 \text{ g I}_2) (1 \text{ mol I}_2 / 253.8 \text{ g I}_2) = 0.0630418 \text{ mol I}_2 \text{ (unrounded)}$$

$$\text{Moles solvent} = (100.0 \text{ g CS}_2) (1 \text{ mol CS}_2 / 76.15 \text{ g CS}_2) = 1.3131976 \text{ mol CS}_2 \text{ (unrounded)}$$

$$\text{Mole fraction} = (0.0630418 \text{ mol I}_2) / [(0.0630418) + (1.3131976)] \text{ mol} = 0.045807 = \mathbf{0.046}$$

Molality:

Moles of iodine were calculated in part (b). Kilograms of solvent = $100.0 \text{ g} (1 \text{ kg} / 10^3 \text{ g}) = 0.1000 \text{ kg}$ in all cases.

Iodine in chloroform

$$\text{Molality} = (0.010638 \text{ mol I}_2) / (0.1000 \text{ kg}) = 0.10638 = \mathbf{0.11 \text{ m I}_2}$$

Iodine in carbon tetrachloride

$$\text{Molality} = (0.009850 \text{ mol I}_2) / (0.1000 \text{ kg}) = 0.09850 = \mathbf{0.098 \text{ m I}_2}$$

Iodine in carbon disulfide

$$\text{Molality} = (0.0630418 \text{ mol I}_2) / (0.1000 \text{ kg}) = 0.630418 = \mathbf{0.63 \text{ m I}_2}$$

13.154 The lower the boiling point the greater the volatility. **acetic acid < water < benzene < ethanol < carbon tetrachloride < chloroform < carbon disulfide < diethyl ether**

13.155 Use the equation:

$$\ln \frac{P_2}{P_1} = -\frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

$$P_1 = 1.00 \text{ atm} \quad T_1 = (273 + 100) \text{ K} = 373 \text{ K}$$

$$P_2 = ? \quad T_2 = (273 + 200.) \text{ K} = 473 \text{ K}$$

$$\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$$

$$\ln \frac{P_2}{1.00 \text{ atm}} = -\frac{40.7 \text{ kJ/mol}}{8.314 \frac{\text{J}}{\text{mol}\cdot\text{K}}} \left(\frac{1}{473 \text{ K}} - \frac{1}{373 \text{ K}} \right) \left(\frac{10^3 \text{ J}}{1 \text{ kJ}} \right)$$

$$\ln \frac{P_2}{1.00 \text{ atm}} = 2.774689665 \text{ (unrounded)}$$

$$\frac{P_2}{1.00 \text{ atm}} = 16.03365$$

$$P_2 = 16.03365 = \mathbf{16.0 \text{ atm}}$$

13.156 a) $\Delta T_f = iK_f m$ Assume NaCl is a strong electrolyte with $i = 2$.

$$m = \Delta T_f / iK_f = (5.0^\circ\text{C}) / [(2)(1.86^\circ\text{C}/m)] = 1.344086 \text{ m NaCl (unrounded)}$$

$$\text{Mass} = \left(\frac{1.344086 \text{ mol NaCl}}{\text{kg}} \right) (5.5 \text{ kg}) \left(\frac{58.44 \text{ g NaCl}}{\text{mol NaCl}} \right) = 432.016 = \mathbf{4.3 \times 10^2 \text{ g NaCl}}$$

b) $\Delta T_f = iK_f m$ Assume CaCl_2 is a strong electrolyte with $i = 3$.

$$m = \Delta T_f / iK_f = (5.0^\circ\text{C}) / [(3)(1.86^\circ\text{C}/m)] = 0.896057 \text{ m CaCl}_2 \text{ (unrounded)}$$

$$\text{Mass} = \left(\frac{0.896057 \text{ mol CaCl}_2}{\text{kg}} \right) (5.5 \text{ kg}) \left(\frac{110.98 \text{ g CaCl}_2}{\text{mol CaCl}_2} \right) = 546.944 = \mathbf{5.5 \times 10^2 \text{ g CaCl}_2}$$

13.157 a) Find the molarity of C_6H_{14} so that the molarity of O_2 can be found from its mole fraction. Then Henry's Law can be used to find k_H .

$$\text{Molarity of } \text{C}_6\text{H}_{14} = (1 \text{ L}) \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right) \left(\frac{1.674 \text{ g C}_6\text{F}_{14}}{1 \text{ mL}} \right) \left(\frac{1 \text{ mol C}_6\text{F}_{14}}{338 \text{ g C}_6\text{F}_{14}} \right) = 4.9527 \text{ mol/L}$$

$$X_{\text{O}_2} = \frac{\text{mol of O}_2}{\text{mol of O}_2 + \text{mol of C}_6\text{F}_{14}}$$

$$4.28 \times 10^{-3} = \frac{x \text{ mol O}_2}{4.9527 \text{ mol}}$$

Moles of $\text{O}_2 = 0.021198$ (The moles of O_2 is small enough to ignore in the denominator.)

$$S_{\text{gas}} = k_H \times P_{\text{gas}}$$

$$k_H = S_{\text{gas}} / P_{\text{gas}}$$

$$k_H = 0.021198 \text{ mol/L}(101,325 \text{ Pa})(101,325 \text{ Pa}/1 \text{ atm}) = 0.021198 = \mathbf{0.0212 \text{ mol/L}\cdot\text{atm}}$$

$$\text{b) } S_{\text{gas}} = k_H \times P_{\text{gas}} \quad k_H = \frac{1}{756.7 \text{ L}\cdot\text{atm/mol}} = 1.322 \times 10^{-3} \text{ mol/L}\cdot\text{atm}$$

$$P_{\text{O}_2} = 0.2095 \times 1 \text{ atm} = 0.2095 \text{ atm} \quad \text{since O}_2 \text{ is } 20.95\% \text{ of air}$$

$$S_{\text{gas}} = (1.322 \times 10^{-3} \text{ mol/L}\cdot\text{atm})(0.2095 \text{ atm}) = 2.76959 \times 10^{-4} \text{ mol/L}$$

$$\text{ppm} = \left(\frac{2.76959 \times 10^{-4} \text{ mol}}{\text{L}} \right) \left(\frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 8.88627 \times 10^{-3} \text{ g/L} = (8.88627 \times 10^{-3} \text{ g}/1000 \text{ g})(1 \times 10^6)$$

$$= 8.895688 = \mathbf{8.90 \text{ ppm}}$$

c) k_H : $C_6F_{14} > C_6H_{14} > \text{ethanol} > \text{water}$

To dissolve oxygen in a solvent, the solvent molecules must be moved apart to make room for the gas. The stronger the intermolecular forces in the solvent, the more difficult it is to separate solvent particles and the lower the solubility of the gas. Both C_6F_{14} and C_6H_{14} have weak dispersion forces with C_6H_{14} having the weaker forces due to the electronegative fluorine atoms repelling each other. Both ethanol and water are held together by strong hydrogen bonds with those bonds being stronger in water as the boiling point indicates.

13.158 a) Molarity of $N_2 = (1.00 \text{ atm}) (78\% N_2 / 100\%) (7.0 \times 10^{-4} \text{ mol/L}\cdot\text{atm}) = 5.46 \times 10^{-4} = \mathbf{5.5 \times 10^{-4} M N_2}$

b) The additional pressure due to 50. ft of water must be added to 1.00 atm.

Water pressure: The value, 9.80665 m/s^2 , is the standard acceleration of gravity from the inside back cover of the book.

$$P_{\text{water}} = \left(\frac{1.00 \text{ g}}{\text{mL}} \right) \left(\frac{1 \text{ mL}}{1 \text{ cm}^3} \right) \left(\frac{1 \text{ cm}}{10^{-2} \text{ m}} \right)^2 \left(\frac{2.54 \text{ cm}}{1 \text{ in}} \right) \left(\frac{12 \text{ in}}{1 \text{ ft}} \right) (50.0 \text{ ft}) \left(\frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left(9.80665 \frac{\text{m}}{\text{s}^2} \right) \left(\frac{1 \text{ Pa}}{1 \text{ kg/m}\cdot\text{s}^2} \right) \left(\frac{1 \text{ atm}}{1.013 \times 10^5 \text{ Pa}} \right)$$

$$= 1.47535 \text{ atm (unrounded)}$$

This is the pressure due to the 50. ft of water, and it must be added to the atmospheric pressure pressing down on the surface of the water (1.00 atm). This gives an unrounded total pressure of 2.47535 atm.

Molarity of $N_2 = (2.47535 \text{ atm}) (78\% N_2 / 100\%) (7.0 \times 10^{-4} \text{ mol/L}\cdot\text{atm})$

$$= 1.35154 \times 10^{-3} = \mathbf{1.4 \times 10^{-3} M N_2}$$

c) Moles of N_2 per liter at the surface = $5.56 \times 10^{-4} \text{ mol } N_2$.

Moles of N_2 per liter at 50. ft = $1.35154 \times 10^{-3} \text{ mol } N_2$.

Moles N_2 released per liter = $(1.35154 \times 10^{-3} - 5.56 \times 10^{-4}) \text{ mol} = 7.9554 \times 10^{-4} \text{ mol (unrounded)}$

$$PV = nRT \text{ so } V = nRT / P = \frac{(7.9554 \times 10^{-4} \text{ mol}) \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) ((273 + 25) \text{K})}{(1.00 \text{ atm})} \left(\frac{1 \text{ mL}}{10^{-3} \text{ L}} \right)$$

$$= 19.4635 = \mathbf{19 \text{ mL } N_2}$$

13.159 a) **Yes**, the phases of water can still coexist at some temperature and can therefore establish equilibrium.

b) The triple point would occur at a lower pressure and **lower temperature** because the dissolved air solute lowers the vapor pressure of the solvent.

c) **Yes**, this is possible because the gas-solid phase boundary exists below the new triple point.

d) **No**, the presence of the solute lowers the vapor pressure of the liquid.

13.160 a) Assuming 100 g of water, the solubilities (in g) of the indicated salts at the indicated temperatures would be:

	KNO_3	$KClO_3$	KCl	NaCl
50°C	85	18	42	36
0°C	12	4	28	35
Difference	73	14	14	1
% recovery	86	78	33	3

(The "difference" is the number of grams of the salt, which could be recovered if a solution containing the amount of salt in the first line were cooled to 0°C. The "% recovery" is calculated by dividing the "difference" by the original amount, then multiplying by 100.)

The highest percent recovery would be found for KNO_3 (86%), and the lowest would be for NaCl (3%).

b) If you began with 100. g of the salts given above, then the "% recovery" line above gives the number of grams which could be recovered by the process described.

13.161 a) Concentration = $\left(\frac{28 \text{ mL}}{7.0 \text{ L}} \right) \left(\frac{40\%}{100\%} \right) \left(\frac{0.789 \text{ g}}{\text{mL}} \right) \left(\frac{22\%}{100\%} \right) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) = 2.77728 \times 10^{-4} = \mathbf{2.8 \times 10^{-4} g/mL}$

b) $\left(\frac{28 \text{ mL}}{2.77728 \times 10^{-4} \text{ g/mL}} \right) (8.0 \times 10^{-4} \text{ g/mL}) = 80.6545 = \mathbf{81 \text{ mL}}$

13.162 a) Moles $\text{CO}_2 = (355 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{3.3 \times 10^{-2} \text{ mol}}{\text{L}\cdot\text{atm}} \right) (4 \text{ atm}) = 0.04686 = \mathbf{0.05 \text{ mol CO}_2}$

b) If it is completely flat there is no CO_2 remaining or 0.00 moles CO_2 , however a small amount will remain in solution:

$$\text{Moles CO}_2 = (355 \text{ mL}) \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}} \right) \left(\frac{3.3 \times 10^{-2} \text{ mol}}{\text{L}\cdot\text{atm}} \right) (3 \times 10^{-4} \text{ atm}) = 3.5145 \times 10^{-6} = \mathbf{4 \times 10^{-6} \text{ mol CO}_2}$$

c) The difference in the moles will determine the number of moles entering the gas phase.

$$\text{PV} = n\text{RT so } V = n\text{RT} / P = \frac{\left[(0.04686 - 3.5145 \times 10^{-6}) \text{ mol} \right] \left(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) ((273 + 25) \text{ K})}{(1.00 \text{ atm})}$$

$$= 1.14638 = \mathbf{1 \text{ L CO}_2}$$

13.163 a) **Scene C** represents the system at the higher temperature of 298 K. At the higher temperature, the solubility of oxygen decreases so more oxygen leaves the solution to go into the vapor phase.

b) **Scene B** represents the system when the pressure of oxygen is increased by half. The increase in pressure would result in $4 + \frac{1}{2}(4) = 6$ moles of oxygen in the vapor phase. The increased pressure results in increased solubility of oxygen in the water. Of the 6 moles of oxygen in the vapor phase, 1 mole dissolves in the water to bring the dissolved moles to 3.