CHAPTER 13 THE PROPERTIES OF SOLUTIONS

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 $CH_3(CH_2)_nCOOH$, 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 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.

13.5 b) Stearic acid in CCl_4. Stearic acid will not dissolve in water. It is nonpolar while water is very polar. Stearic acid will dissolve in carbon tetrachloride, as both are nonpolar.

13.6 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.

13.7 a) **metallic bonding**
b) **dipole-dipole**
c) **dipole-induced dipole**

13.8 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 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.

13.9 a) **dispersion forces**
b) **hydrogen bonding**
c) **dispersion forces**

13.10 $CH_3CH_2OCH_2CH_3$ 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) **CH_3CH_2MgBr** would be more soluble. CH_3CH_2MgBr has a polar end (–MgBr) and a nonpolar end (CH_3CH_2–), whereas $MgBr_2$ is an ionic compound. The nonpolar end of CH_3CH_2MgBr and diethyl ether would interact with dispersion forces, while the polar end of CH_3CH_2MgBr 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.

13.11 a) **CH_3CH_2–O–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.

13.12 Gluconic acid is a very polar molecule because it has –OH groups attached to every carbon. The abundance of –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.

13.13 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.

13.14 a) Charge density is the ratio of the ion's charge to its volume.

b) $- < + < 2- < 3+$

c) The higher the charge density, the more negative is $\Delta H_{\text{hydration}}$. $\Delta H_{\text{hydration}}$ increases with charge and decreases with increasing volume.

13.15 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.

$$\Delta H_{\text{solution}} = \Delta H_{\text{lattice}} + \Delta H_{\text{hydration of ions}}$$

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

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

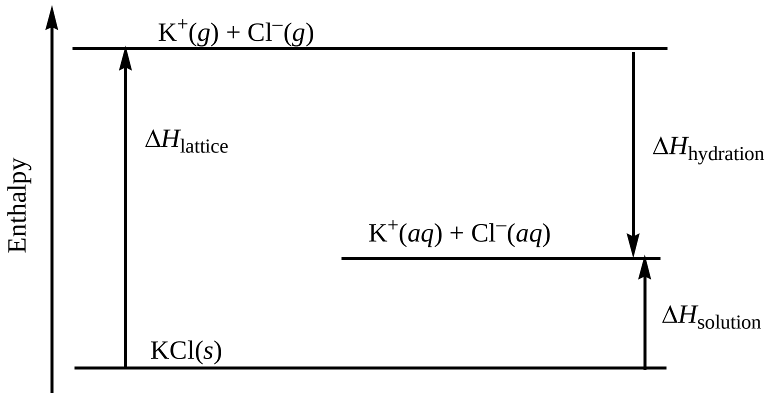
13.16 a) **Endothermic**

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.

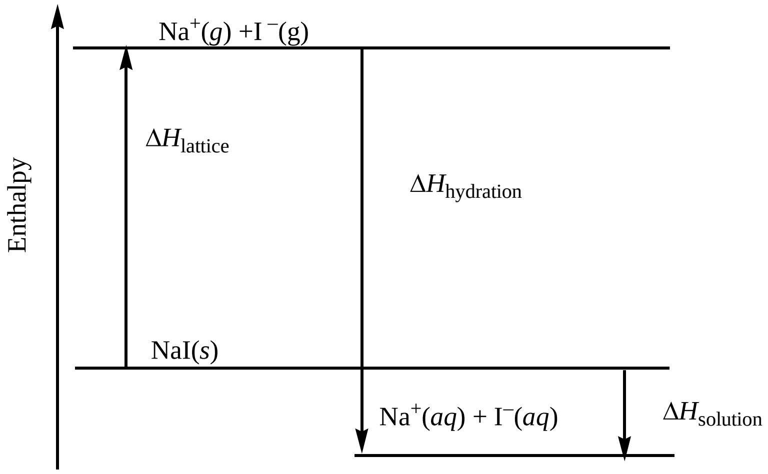
13.17 This compound would be very soluble in water. A large exothermic value in $\Delta H_{\text{solution}}$ (enthalpy of solution) means that the solution has a much lower energy state than the isolated solute and solvent particles, so the system tends to the formation of the solution. Entropy that 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.

13.18



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.

13.19



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.

13.20 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.

a) Both ions have a +1 charge, but the volume of $\mathbf{Na^+}$ is smaller, so it has the greater charge density.

b) $\mathbf{Sr^{2+}}$ has a greater ionic charge and a smaller size (because it has a greater Z_{eff}), so it has the greater charge density.

c) $\mathbf{Na^+}$ has a smaller ion volume than Cl^-, so it has the greater charge density.

d) $\mathbf{O^{2-}}$ has a greater ionic charge and similar ion volume, so it has the greater charge density.

e) $\mathbf{OH^-}$ has a smaller ion volume than SH^-, so it has the greater charge density.

13.21 a) **I^-** has a smaller charge density (larger ion volume) than Br^-.
b) **Ca^{2+}** is less than Sc^{3+}, due to its smaller ion charge.
c) **Br^-** is less than K^+, due to its larger ion volume.
d) **Cl^-** is less than S^{2-}, due to its smaller ion charge.
e) **Sc^{3+}** is less than Al^{3+}, due to its larger ion volume.

13.22 The ion with the greater charge density will have the larger $\Delta H_{\text{hydration}}$.
a) **Na^+** would have a larger $\Delta H_{\text{hydration}}$ than Cs^+ since its charge density is greater than that of Cs^+.
b) **Sr^{2+}** would have a larger $\Delta H_{\text{hydration}}$ than Rb^+.
c) **Na^+** would have a larger $\Delta H_{\text{hydration}}$ than Cl^-.
d) **O^{2-}** would have a larger $\Delta H_{\text{hydration}}$ than F^-.
e) **OH^-** would have a larger $\Delta H_{\text{hydration}}$ than SH^-.

13.23 a) **I^-** b) **Ca^{2+}** c) **Br^-** d) **Cl^-** e) **Sc^{3+}**

13.24 a) The two ions in potassium bromate are K^+ and BrO_3^-. The heat of solution for ionic compounds is $\Delta H_{\text{solution}} = \Delta H_{\text{lattice}} + \Delta H_{\text{hydration of the ions}}$. Therefore, the combined heats of hydration for the ions is ($\Delta H_{\text{solution}} - \Delta H_{\text{lattice}}$) or 41.1 kJ/mol – 745 kJ/mol = –703.9 = **–704 kJ/mol**.
b) **K^+** ion contributes more to the heat of hydration because it has a smaller size and, therefore, a greater charge density.

13.25 a) $\Delta H_{\text{hydration of ions}} = \Delta H_{\text{solution}} - \Delta H_{\text{lattice}}$
$\Delta H_{\text{hydration of ions}}$ = 17.3 kJ/mol –763 kJ/mol
$\Delta H_{\text{hydation of ions}}$ = –745.7 = **– 746 kJ/mol**
b) It is the Na^+ due to its smaller size (larger charge density).

13.26 Entropy increases as the possible states for a system increases.
a) Entropy **increases** as the gasoline is burned. Gaseous products at a higher temperature form.
b) 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.
c) Entropy **increases** as a solute dissolves in the solvent.

13.27 a) Entropy **increases**
b) Entropy **decreases**
c) Entropy **increases**

13.28 $\Delta H_{\text{solution}} = \Delta H_{\text{lattice}} + \Delta H_{\text{hydration of ions}}$
$\Delta H_{\text{solution}}$ = 822 kJ/mol – 799 kJ/mol
$\Delta H_{\text{solution}}$ = **23 kJ/mol**

13.29 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.

13.30 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 ($\Delta H_{\text{soln}} < 0$).

13.31 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.

13.32 a) **increase** b) **stay the same**

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

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

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

$$\left(\frac{1.28 \text{ x } 10^{-3} \text{ mol } O_2}{\text{L}}\right)\left(\frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2}\right)(2.50 \text{ L}) = 0.1024 = \mathbf{0.102 \text{ g } O_2}$$

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

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

$$\left(\frac{2.6752 \text{ x } 10^{-4} \text{ mol } O_2}{\text{L}}\right)\left(\frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2}\right)(2.50 \text{ L}) = 0.0214016 = \mathbf{0.0214 \text{ g } O_2}$$

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

13.35 The solution is **saturated**.

13.36 Solubility for a gas is calculated from Henry's law: $S_{gas} = k_H \text{ x } P_{gas}$.

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

$$S_{gas} = \left(3.7 \text{ x } 10^{-2} \frac{\text{mol}}{\text{L} \bullet \text{atm}}\right)(5.5 \text{ atm}) = 0.2035 = \mathbf{0.20 \text{ mol/L}}$$

13.37 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 does not change with temperature.

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

13.39 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. Why is the identity of the solute not necessary for conversion?

13.40 % 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.

13.41 Convert the masses to moles and the volumes to liters and use the definition of molarity: $M = \frac{\text{mol of solute}}{V\text{(L) of solution}}$

a) Molarity = $\left(\frac{32.3 \text{ g } C_{12}H_{22}O_{11}}{100. \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.30 \text{ g } C_{12}H_{22}O_{11}}\right) = 0.943617 = \mathbf{0.944}\ \boldsymbol{M}\ \mathbf{C_{12}H_{22}O_{11}}$

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 = **0.167 M $LiNO_3$**

13.42 a) Molarity = $\left(\frac{0.82 \text{ g } C_2H_5OH}{10.5 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right)\left(\frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g } C_2H_5OH}\right)$ = 1.69514 = **1.7 M C_2H_5OH**

b) Molarity = $\left(\frac{1.22 \text{ g NH}_3}{33.5 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3}\text{ L}}\right)\left(\frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3}\right)$ = 2.138456 = **2.14 M NH_3**

13.43 Dilution calculations can be done using $M_{\text{conc}}\text{V}_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$

a) M_{conc} = 0.240 M NaOH V_{conc} = 78.0 mL M_{dil} = ? V_{dil} = 0.250 L

$$M_{\text{dil}} = M_{\text{conc}}\, V_{\text{conc}}/V_{\text{dil}} = \frac{(0.240\ M)(78.0 \text{ mL})}{(0.250 \text{ L})}\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right) = 0.07488 = \mathbf{0.0749}\ \boldsymbol{M}$$

b) M_{conc} = 1.2 M HNO_3 V_{conc} = 38.5 mL M_{dil} = ? V_{dil} = 0.130 L

$$M_{\text{dil}} = M_{\text{conc}}\, V_{\text{conc}}/V_{\text{dil}} = \frac{(1.2\ M)(38.5 \text{ mL})}{(0.130 \text{ L})}\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right) = 0.355385 = \mathbf{0.36}\ \boldsymbol{M}$$

13.44 Dilution calculations can be done using $M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$

a) M_{conc} = 6.25 M HCl V_{conc} = 25.5 mL M_{dil} = ? V_{dil} = 0.500 L

$$M_{\text{dil}} = M_{\text{conc}}\, V_{\text{conc}}/V_{\text{dil}} = \frac{(6.25\ M)(25.5 \text{ mL})}{(0.500 \text{ L})}\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right) = 0.31875 = \mathbf{0.319}\ \boldsymbol{M}$$

b) M_{conc} = 2.00 x 10^{-2} M KI V_{conc} = 8.25 mL M_{dil} = ? V_{dil} = 12.0 mL

$$M_{\text{dil}} = M_{\text{conc}}\, V_{\text{conc}}/V_{\text{dil}} = \frac{(2.00 \text{ x } 10^{-2}\ M)(8.25 \text{ mL})}{(12.0 \text{ mL})} = 0.01375 = \mathbf{0.0138}\ \boldsymbol{M}$$

13.45 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)

$$\text{Mass } KH_2PO_4 = (365 \text{ mL})\left(\frac{10^{-3}\text{ L}}{1 \text{ mL}}\right)\left(\frac{8.55 \text{ x } 10^{-2}\text{ mol } KH_2PO_4}{\text{L}}\right)\left(\frac{136.09 \text{ g } KH_2PO_4}{1 \text{ mol } KH_2PO_4}\right)$$

= 4.24703 = 4.25 g KH_2PO_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_{conc} = 1.25 M NaOH V_{conc} = ? M_{dil} = 0.335 M NaOH V_{dil} = 465 mL

$$V_{\text{conc}} = M_{\text{dil}}\, V_{\text{dil}}/M_{\text{conc}} = \frac{(0.335\ M)(465 \text{ mL})}{(1.25\ M)} = 124.62 = \mathbf{125 \text{ mL}}$$

Add **125 mL** of 1.25 M NaOH to enough water to make 465 mL of solution.

13.46 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_{conc} = 2.1 M urea V_{conc} = ? M_{dil} = 0.3 M urea V_{dil} = 15.5 L

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

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

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

a) *m* 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 = **0.896 *m* glycine**

b) *m* 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 = **1.21 *m* glycerol**

13.48 Molality = moles solute/kg solvent.

a) *m* HCl = $\frac{164\text{ g HCl}\left(\frac{1\text{ mol HCl}}{36.46\text{ g HCl}}\right)}{(753\text{ g})}\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)$ = 5.9735 = **5.97 *m* HCl**

b) *m* 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 = **2.42 *m* naphthalene**

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

m benzene = $\frac{(44.0\text{ mL }C_6H_6)\left(\frac{0.877\text{ g}}{\text{mL}}\right)\left(\frac{1\text{ mol }C_6H_6}{78.11\text{ g }C_6H_6}\right)}{(167\text{ mL }C_6H_{14})\left(\frac{0.660\text{ g}}{\text{mL}}\right)}\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)$ = 4.48214 = **4.48 *m* C_6H_6**

13.50 Molality = moles solute/kg of solvent.

m CCl_4 = $\frac{(2.66\text{ mL }CCl_4)\left(\frac{1.59\text{ g}}{\text{mL}}\right)\left(\frac{1\text{ mol }CCl_4}{153.81\text{ g }CCl_4}\right)}{(76.5\text{ mL }CH_2Cl_2)\left(\frac{1.33\text{ g}}{\text{mL}}\right)}\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)$ = 0.2702596 = **0.270 *m* CCl_4**

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

Mass of $C_2H_6O_2$ in 1000 g (1 kg) of H_2O = $\left(\frac{0.125\text{ mol }C_2H_6O_2}{1\text{ kg }H_2O}\right)\left(\frac{62.07\text{ g }C_2H_6O_2}{1\text{ mol }C_2H_6O_2}\right)$

= 7.75875 g $C_2H_6O_2$ in 1000 g H_2O

Grams of this solution = 1000 g H_2O + 7.75875 g $C_2H_6O_2$ = 1007.75875 g

Mass $C_2H_6O_2$ = $\left(\frac{7.75875\text{ g }C_2H_6O_2}{1007.75875\text{ g solution}}\right)(3.10\text{ x }10^2\text{g solution})$ = 2.386695 g $C_2H_6O_2$

$\text{Mass}_{\text{solvent}}$ = 3.10 x 10^2 g – $\text{mass}_{\text{solute}}$ = 3.10 x 10^2 g – 2.386695 g $C_2H_6O_2$ = 307.613305 = 308 g H_2O

Therefore, **add 2.39 g $C_2H_6O_2$ to 308 g of H_2O** 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 kg HNO_3 (solute)

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

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

Mass solute needed = 0.050769 = 0.0508 kg
Mass solvent = Mass solution – Mass solute = 1.20 kg – 0.050769 kg = 1.149231 = 1.15 kg
Add 0.0508 kg of the 52.0% (w/w) HNO_3 to 1.15 kg H_2O to make 1.20 kg of 2.20% (w/w) HNO_3.

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

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

g $C_2H_6O_2$/1000 g H_2O in 0.0355 m = $\left(\frac{0.0355 \text{ mol } C_2H_5OH}{1 \text{ kg } H_2O}\right)\left(\frac{46.07 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH}\right)$

= 1.635485 g C_2H_5OH/1000 g H_2O (unrounded)

Grams of this solution = 1000 g H_2O + 1.635485 g C_2H_5OH = 1001.635485 g (unrounded)

Mass $C_2H_6O_2$ = [1.635485 g C_2H_5OOH/1001.635485 g solution] [1.50 kg solution (10^3 g/1 kg)]
= 2.449222 = 2.45 g C_2H_5OH

$\text{Mass}_{\text{solvent}}$ = 1500 g – $\text{mass}_{\text{solute}}$ = 1500 g – 2.449222 g $C_2H_6O_2$ = 1497.551 = 1498 g H_2O
Therefore, **add 2.45 g C_2H_5OH to 1498 g of H_2O** 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 g) (13.0%/100%) = 57.85 g HCl (solute) (unrounded)
Then determine the amount of the concentrated acid solution needed to get 57.85 g solute:
(Mass needed) (34.1%/100%) = 57.85 g
Mass solute needed = 169.6481 = 170. g
Mass solvent = Mass solution – Mass solute = 445 g – 169.6481 g = 275.35191 = 275 g
Add 170. g of the 34.1% (w/w) HCl to 275 g H_2O

13.53 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 = **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:

Mass isopropanol = $(0.35 \text{ mol } C_3H_7OH)\left(\frac{60.09 \text{ g } C_3H_7OH}{1 \text{ mol } C_3H_7OH}\right)$ = 21.0315 g isopropanol

Mass water = $(0.85 \text{ mol } H_2O)\left(\frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$ = 15.317 g water (unrounded)

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

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

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 = **23 *m*** isopropanol

13.54 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 = **0.0115** (Notice that mole fractions have no units.)

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

Mass NaCl = (0.100 mol NaCl) (58.44 g/mol) = 5.844 g NaCl (unrounded)
Mass water = (8.60 mol water) (18.02 g/mol) = 154.972 g water (unrounded)

Percent NaCl = $\frac{(5.844 \text{ g NaCl})}{(5.844 + 154.972) \text{g}}$ x100% = 3.63396677 = **3.63%** NaCl

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

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

13.55 The information given is 8.00 mass % NH_3 solution with a density of 0.9651 g/mL. For convenience, choose exactly 100.00 grams of solution. Determine some fundamental quantities:

$$\text{Mass of } NH_3 = (100 \text{ g solution})\left(\frac{8.00\% \text{ } NH_3}{100\% \text{ solution}}\right) = 8.00 \text{ g } NH_3$$

$$\text{Mass } H_2O = \text{mass of solution} - \text{mass } NH_3 = (100.00 - 8.00) \text{ g} = 92.00 \text{ g } H_2O$$

$$\text{Moles } NH_3 = (8.00 \text{ g } NH_3)\left(\frac{1 \text{ mol } NH_3}{17.03 \text{ g } NH_3}\right) = 0.469759 \text{ mol } NH_3 \text{ (unrounded)}$$

$$\text{Moles } H_2O = (92.00 \text{ g } H_2O)\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 5.1054 \text{ mol } H_2O \text{ (unrounded)}$$

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

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

$$\text{Molality} = \frac{\text{Moles solute}}{\text{kg solvent}} = \left(\frac{0.469759 \text{ mol } NH_3}{92.00 \text{ g } H_2O}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 5.106076 = \mathbf{5.11}\ \boldsymbol{m}\ \mathbf{NH_3}$$

$$\text{Molarity} = \frac{\text{Moles solute}}{\text{L solution}} = \left(\frac{0.469759 \text{ mol } NH_3}{0.103616 \text{ L}}\right) = 4.53365 = \mathbf{4.53}\ \boldsymbol{M}\ \mathbf{NH_3}$$

$$\text{Mole fraction} = X = \frac{\text{Moles substance}}{\text{total moles}} = \frac{0.469759 \text{ mol } NH_3}{(0.469759 + 5.1054)\text{mol}} = 0.084259 = \mathbf{0.0843}$$

13.56 The information given is 28.8 mass % $FeCl_3$ solution with a density of 1.280 g/mL.

For convenience, choose exactly 100.00 grams of solution.

Determine some fundamental quantities:

Mass of $FeCl_3$ = (100.00 grams solution) (28.8% $FeCl_3$/100%) = 28.8 g $FeCl_3$

Mass H_2O = mass of solution – mass $FeCl_3$ = (100.00 – 28.8) g = 71.20 g H_2O

Moles $FeCl_3$ = (28.80 g $FeCl_3$) (1 mol $FeCl_3$/162.20 g $FeCl_3$) = 0.1775586 mol $FeCl_3$ (unrounded)

Moles H_2O = (71.20 g H_2O) (1 mol H_2O/18.02 g H_2O) = 3.951165 mol H_2O (unrounded)

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

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

$$\text{Molality} = \text{Moles solute/kg solvent} = \left(\frac{0.1775586 \text{ mol } FeCl_3}{71.20 \text{ g } H_2O}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 2.49380 = \mathbf{2.49}\ \boldsymbol{m}\ \mathbf{FeCl_3}$$

$$\text{Molarity} = \text{Moles solute/L solution} = \frac{0.1775586 \text{ mol } FeCl_3}{0.078125 \text{ L}} = 2.272750 = \mathbf{2.27}\ \boldsymbol{M}\ \mathbf{FeCl_3}$$

$$\text{Mole fraction} = X = \text{moles substance/total moles} = \frac{0.1775586 \text{ mol } FeCl_3}{(0.1775586 + 3.951165)\text{mol}} = 0.043005688 = \mathbf{0.0430}$$

13.57 $\text{ppm} = \left(\dfrac{\text{mass solute}}{\text{mass solution}}\right) \text{ x } 10^6$

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

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

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

13.58 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:

Mass of EG = (1.0000 L EG) (1mL/10^{-3} L) (1.114 g EG/mL) = 1114 g EG
Mass of H_2O = (1.0000 L H_2O) (1mL/10^{-3} L) (1.00 g H_2O/mL) = 1.00 x 10^3 g H_2O
Moles EG = (1114 g EG) (1 mol EG/62.07 g EG) = 17.94747865 mol EG (unrounded)
Moles H_2O = (1.00 x 10^3 g H_2O) (1 mol H_2O/18.02 g H_2O) = 55.49389567 mol H_2O (unrounded)
Volume solution = (1114 g EG + 1.00 x 10^3 g H_2O) (mL/1.070 g) (10^{-3} L/1 mL)
= 1.97570 L (unrounded)

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

a) Volume percent = (1.0000 L EG/1.97570 L) 100% = 50.61497 = **50.61% v/v**

b) Mass percent = [(1114 g EG)/(1114 + 1.00 x 10^3) g] 100% = 52.6963 = **52.7% w/w**

c) Molarity = Moles solute / L solution = $\dfrac{17.94747865\text{ mol EC}}{1.97570\text{ L}}$ = 9.08411 = **9.08 *M* ethylene glycol**

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

= 17.94747865 = **17.9 *m* ethylene glycol**

e) Mole fraction = *X* = moles substance / total moles = $\dfrac{17.94747865\text{ mol EG}}{\left(17.94747865 + 55.49389567\right)\text{mol}}$

= 0.244378 = **0.244**

13.59 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.

13.60 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.

13.61 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.

13.62 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.

13.63 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.

13.64 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.

b) **Strong electrolyte** Potassium nitrate is a soluble salt.

c) **Nonelectrolyte** Glucose solid dissolves in water to form individual $C_6H_{12}O_6$ molecules, but these units are not ionic and therefore do not conduct electricity.

d) **Weak electrolyte** Ammonia gas dissolves in water, but is a weak base that forms few NH_4^+ and OH^- ions.

13.65

a) $NaMnO_4$	**strong electrolyte**
b) CH_3COOH	**weak electrolyte**
c) CH_3OH	**nonelectrolyte**
d) $Ca(C_2H_3O_2)_2$	**strong electrolyte**

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

Each KBr forms 1 K^+ ion and 1 Br^- ion, 2 particles for each KBr.

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

HNO_3 is a strong acid that forms $H^+(H_3O^+)$ ions and NO_3^- ions in aqueous solution.

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

Each $KHSO_4$ forms 1 K^+ ion and 1 HSO_4^- ion in aqueous solution, 2 particles for each $KHSO_4$.

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

13.67 a) (0.01 mol $CuSO_4$/L) (2 mol particles/mol $CuSO_4$) (10^{-3} L/1 mL) (1 mL) = **2×10^{-5} mol of particles**

b) (0.005 mol $Ba(OH)_2$/L) (3 mol particles/mol $Ba(OH)_2$) (10^{-3} L/1 mL) (1 mL)
= 1.5×10^{-5} = **2×10^{-5} mol of particles**

c) (0.06 mol C_5H_5N/L) (1 mol particles/mol C_5H_5N) (10^{-3} L/1 mL) (1 mL) = **6×10^{-5} mol of particles**

d) (0.05 mol $(NH_4)_2CO_3$/L) (3 mol particles/mol $(NH_4)_2CO_3$) (10^{-3} L/1 mL) (1 mL)
= 1.5×10^{-4} = **2×10^{-4} mol of particles**

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

a) Molality of CH_3OH = $\frac{(11.0\text{ g CH}_3\text{OH})}{(100.\text{ g H}_2\text{O})}\left(\frac{1\text{ mol CH}_3\text{OH}}{32.04\text{ g CH}_3\text{OH}}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)$ = 3.4332085 = 3.43 *m* CH_3OH

Molality of CH_3CH_2OH = $\frac{(22.0\text{ g CH}_3\text{CH}_2\text{OH})}{(200.\text{ g H}_2\text{O})}\left(\frac{1\text{ mol CH}_3\text{CH}_2\text{OH}}{46.07\text{ g CH}_3\text{CH}_2\text{OH}}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)$

= 2.387671 = 2.39 *m* CH_3CH_2OH

The molality of methanol, CH_3OH, in water is 3.43 *m* whereas the molality of ethanol, CH_3CH_2OH, in water is 2.39 m. Thus, **CH_3OH/H_2O solution** has the lower freezing point.

b) Molality of H_2O = $\frac{(20.0\text{ g }H_2O)}{(1.00\text{ kg }CH_3OH)}\left(\frac{1\text{ mol }H_2O}{18.02\text{ g }H_2O}\right)$ = 1.10988 = 1.11 m H_2O

Molality of CH_3CH_2OH = $\frac{(20.0\text{ g }CH_3CH_2OH)}{(1.00\text{ kg }CH_3OH)}\left(\frac{1\text{ mol }CH_3CH_2OH}{46.07\text{ g }CH_3CH_2OH}\right)$ = 0.434122 = 0.434 m CH_3CH_2OH

The molality of H_2O in CH_3OH is 1.11 *m*, whereas CH_3CH_2OH in CH_3OH is 0.434 *m*. Therefore, **H_2O/CH_3OH solution** has the lower freezing point.

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

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 m $C_3H_8O_3$

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 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.

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 m $C_2H_6O_2$

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 m NaCl

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

(2 particles/NaCl) (0.513347 mol NaCl/kg) = 1.026694 = 1.0 *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.

13.70 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 and boiling point, and the lower the freezing point and vapor pressure at a given temperature.

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

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

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

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

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

a) Osmotic pressure:	$\mathbf{\Pi_{II} < \Pi_{I} < \Pi_{III}}$
b) Boiling point:	$\mathbf{bp_{II} < bp_{I} < bp_{III}}$
c) Freezing point:	$\mathbf{fp_{III} < fp_{I} < fp_{II}}$
d) Vapor pressure at 50°C:	$\mathbf{vp_{III} < vp_{I} < vp_{II}}$

13.71 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:	$\mathbf{\Pi_{II} < \Pi_{I} < \Pi_{III}}$
b) Boiling point:	$\mathbf{bp_{II} < bp_{I} < bp_{III}}$
c) Freezing point:	$\mathbf{fp_{III} < fp_{I} < fp_{II}}$
d) Vapor pressure at 298 K:	$\mathbf{vp_{III} < vp_{I} < vp_{II}}$

13.72 The mol fraction of solvent affects the vapor pressure according to the equation: $P_{solvent} = X_{solvent}P°_{solvent}$

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

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

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

$P_{solvent} = X_{solvent}P°_{solvent}$ = (0.9868686) (23.76 torr) = 23.447998 = **23.4 torr**

13.73 The mole fraction of solvent affects the vapor pressure according to the equation: $P_{solvent} = X_{solvent}P°_{solvent}$

$X_{solvent}$ = (5.4 mol toluene)/[(0.39) + (5.4)] mol = 0.93264 (unrounded)
$P_{solvent} = X_{solvent}P°_{solvent}$ = (0.93264) (41 torr) = 38.2382 = **38 torr**

13.74 The change in freezing point is calculated from $\Delta T_f = iK_fm$, where K_f is 1.86°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°C/*m*) (0.251 *m*) = 0.46686°C (unrounded)

The freezing point is 0.00°C – 0.46686°C = –0.46686 = **–0.467°C**.

13.75 $\Delta T_b = iK_b\, m$ = (1) (0.512°C/m) (0.200 m) = 0.1024°C (unrounded)
The boiling point is 100.00°C + 0.1024°C = 100.1024 = **100.10°C**

13.76 The boiling point of a solution is increased relative to the pure solvent by the relationship $\Delta T_b = iK_bm$. Vanillin is a nonelectrolyte so $i = 1$. The molality must be calculated, and K_b is given (1.22°C/*m*).

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

= 0.841330354 *m* Vanillin (unrounded)

$\Delta T_b = iK_b\, m$ = (1) (1.22°C/*m*) (0.841330354 *m*) = 1.02642°C (unrounded)

The boiling point is 78.5°C + 1.02642°C = 79.52642 = **79.5°C**.

13.77 Moles $C_{10}H_8$ = (5.00 g $C_{10}H_8$) (1 mol $C_{10}H_8$/128.16 g $C_{10}H_8$) = 0.0390137 mol $C_{10}H_8$ (unrounded)
$C_{10}H_8$ is a nonelectrolyte so $i = 1$
Mass = (444 g benzene) (1 kg/10^3 g) = 0.444 kg benzene
Molality = (0.0390137 mol $C_{10}H_8$)/(0.444 kg) = 0.08786869 *m* (unrounded)
$\Delta T_f = i\, K_f\, m$ = (1) (4.90°C/m) (0.08786869 m) = 0.43056°C (unrounded)
Freezing point = (5.5 – 0.43056)°C = 5.06944 = **5.1°C**

13.78 The molality of the solution can be determined from the relationship $\Delta T_f = iK_f m$ with the value 1.86°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 °C and find ΔT_f:

$°C = (5/9)\ (°F - 32.0) = (5/9)\ ((-12.0)°F - 32.0) = -24.44444°C$ (unrounded)

$\Delta T_f = (0.00 - (-24.44444))°C = 24.44444°C$

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

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.

$$\text{Mass ethylene glycol} = \left(\frac{13.14217\ \text{mol EG}}{1\ \text{kg H}_2\text{O}}\right)\left(14.5\ \text{kg H}_2\text{O}\right)\left(\frac{62.07\ \text{g EG}}{1\ \text{mol EG}}\right)$$

$= 1.18282 \times 10^4 =$ **1.18×10^4 g ethylene glycol**

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

13.79 The molality of the solution can be determined from the relationship $\Delta T_f = i\ K_f\ m$ with the value 1.86°C/m inserted for K_f, $i = 1$ for the nonelectrolyte glycerol, and the given ΔT_f of –15°C.

$m = \Delta T_f / K_f = (15°C) / (1.86°C/m) = 8.06452\ m$

Glycerol will be abbreviated as GLY.

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

$= 0.0081693 =$ **0.0082 g glycerol**

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

13.80 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.

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

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

Mass H_2O = 100.00 g solution – 1.00 g NaCl = 99.00 g H_2O

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

Calculate $\Delta T = (0.000 - (-0.593))°C = 0.593°C$

$\Delta T_f = iK_f m$

$$i = \frac{\Delta T_f}{K_f m} = \frac{0.593°C}{\left(1.86°C/m\right)\left(0.172844\ m\right)} = 1.844537 = \mathbf{1.84}$$

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.

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

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

Mass H_2O = 100.00 g solution – 0.500 g CH_3COOH = 99.500 g H_2O

$$\text{Molality } CH_3COOH = \frac{\text{mole } CH_3COOH}{\text{kg } H_2O} = \frac{0.0083264 \text{ mol } CH_3COOH}{99.500 \text{ g } H_2O}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)$$

$$= 0.083682 = \mathbf{0.0837}\ \boldsymbol{m}\ \mathbf{CH_3COOH}$$

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

$\Delta T_f = \text{i}K_f\, m$

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

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.

13.81 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.

$$\text{Molality KCl} = \left(\frac{0.500 \text{ g KCl}}{99.500 \text{ g } H_2O}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}}\right) = 0.067406\ m \text{ KCl (unrounded)}$$

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

$\Delta T_f = i\, K_f\, m$

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

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.

$$\text{Molality } H_2SO_4 = \left(\frac{1.00 \text{ g } H_2SO_4}{99.00 \text{ g } H_2O}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol } H_2SO_4}{98.09 \text{ g } H_2SO_4}\right) = 0.10297696\ m\ H_2SO_4 \text{ (unrounded)}$$

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

$\Delta T_f = i\, K_f\, m$

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

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.

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

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

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

Molar mass = (6.053 g)/(0.001111756 mol) = 5.4445 x 10^3 = **5.44 x 10^3 g/mol**

13.83 The pressure of each compound is proportional to its mole fraction according to Raoult's law: $P_A = X_A\, P°_A$

$$X_{CH_2Cl_2} = \frac{\text{moles } CH_2Cl_2}{\text{moles } CH_2Cl_2 + \text{mol } CCl_4} = \frac{1.60 \text{ mol}}{1.60 + 1.10 \text{ mol}} = 0.592593$$

$$X_{CCl_4} = \frac{\text{moles } CCl_4}{\text{moles } CH_2Cl_2 + \text{mol } CCl_4} = \frac{1.10 \text{ mol}}{1.60 + 1.10 \text{ mol}} = 0.407407$$

$P_A = X_A\, P°_A$

= (0.592593) (352 torr) = 208.593 = **209 torr CH_2Cl_2**

= (0.407407) (118 torr) = 48.0740 = **48.1 torr CCl_4**

13.84 **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.

13.85 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.

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

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

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

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

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

13.86 Xe is a much larger atom than He, so it is much more polarizible. This would increase the dipole-induced dipole forces when Xe is placed in water, increasing the solubility relative to He.

13.87 a) Total molarity of ions:

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

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

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

b) Molarity of compound:

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

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

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

Solution A has the highest molarity.

c) Molality of compound:

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

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

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

Solution C has the lowest molality.
d) Osmotic pressure: assume a temperature of 298 K
Solution A: $\Pi = i\,MRT$ = (2) (1.6 mol/L) (0.0821 L•atm/mol•K) (298 K) = 78.29056 = **78 atm**
Solution B: $\Pi = i\,MRT$ = (2) (1.0 mol/L) (0.0821 L•atm/mol•K) (298 K) = 48.9316 = **49 atm**
Solution C: $\Pi = i\,MRT$ = (3) (0.40 mol/L) (0.0821 L•atm/mol•K) (298 K) = 29.35896 = **29 atm**
Solution A has the highest osmotic pressure.

13.88 Convert mass of O_2 dissolved to moles of O_2. Use the density to convert the 1 kg mass of solution to volume in L.

0.0°C

$$\left(\frac{14.5\text{ mg }O_2}{1\text{ kg }H_2O}\right)\left(\frac{10^{-3}\text{ g}}{1\text{ mg}}\right)\left(\frac{1\text{ mol }O_2}{32.00\text{ g }O_2}\right)\left(\frac{1\text{ kg}}{10^3\text{ g}}\right)\left(\frac{0.99987\text{ g}}{\text{mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

$= 4.53066 \times 10^{-4} = \mathbf{4.53 \times 10^{-4}}\ \boldsymbol{M}\ \mathbf{O_2}$

20.0°C

$$\left(\frac{9.07\text{ mg }O_2}{1\text{ kg }H_2O}\right)\left(\frac{10^{-3}\text{ g}}{1\text{ mg}}\right)\left(\frac{1\text{ mol }O_2}{32.00\text{ g }O_2}\right)\left(\frac{1\text{ kg}}{10^3\text{ g}}\right)\left(\frac{0.99823\text{ g}}{\text{mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

$= 2.829358 \times 10^{-4} = \mathbf{2.83 \times 10^{-4}}\ \boldsymbol{M}\ \mathbf{O_2}$

40.0°C

$$\left(\frac{6.44\text{ mg}O_2}{1\text{ kg }H_2O}\right)\left(\frac{10^{-3}\text{ g}}{1\text{ mg}}\right)\left(\frac{1\text{ mol }O_2}{32.00\text{ g }O_2}\right)\left(\frac{1\text{ kg}}{10^3\text{ g}}\right)\left(\frac{0.99224\text{ g}}{\text{mL}}\right)\left(\frac{1\text{ mL}}{10^{-3}\text{ L}}\right)$$

$= 1.996883 \times 10^{-4} = \mathbf{2.00 \times 10^{-4}}\ \boldsymbol{M}\ \mathbf{O_2}$

13.89 Price NaCl (\$/ion) = (\$0.22/kg NaCl) (1 kg/10^3 g) (58.44 g NaCl/1 mol NaCl) (1 mol NaCl/2 ions)
= \$ 6.4284×10^{-3}/ion (unrounded)
Price $CaCl_2$ (\$/kg) = (\$ 6.4284×10^{-3}/ion) (3 mol ions/1 mol $CaCl_2$) (1 mol $CaCl_2$/110.98 g $CaCl_2$) (10^3 g/kg)
= 0.17377 = **\$ 0.17/kg $CaCl_2$**

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

13.91 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).

$\Delta T_f = iK_f m = (0.000 - (-0.201)) = 0.201°C$

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

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

moles solute = (m)(kg solvent) = (0.1080656 m)(0.0250 kg) = 0.0027016 mol

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

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.

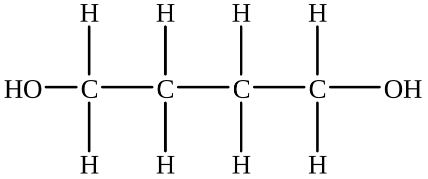
$$\text{Moles C} = (53.31 \text{ g C})\left(\frac{1 \text{ mol C}}{12.01 \text{ g C}}\right) = 4.43880 \text{ mol C (unrounded)}$$

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

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

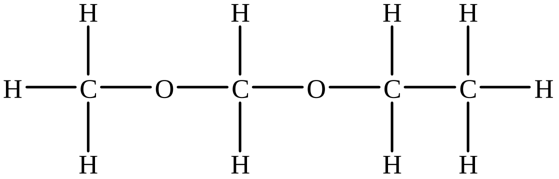
Dividing the values by the lowest amount of moles (2.219375) will give 2 mol C, 5 mol H, and 1 mol O for an **empirical formula C_2H_5O** with molar mass 45.06 g/mol.
Since the molar mass of the compound, 89.9 g/mol from part (a), is twice the molar mass of the empirical formula, the **molecular formula is 2(C_2H_5O) or $C_4H_{10}O_2$**.
c) There is more than one example in each case. Possible Lewis structures:

Forms hydrogen bonds



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

Does not form hydrogen bonds



$$\begin{array}{ccccccccccccccc} & & H & & & & H & & & & H & & H & & \\ & & | & & & & | & & & & | & & | & & \\ H & - & C & - & O & - & C & - & O & - & C & - & C & - & H \\ & & | & & & & | & & & & | & & | & & \\ & & H & & & & H & & & & H & & H & & \end{array}$$

13.92 Calculate the individual partial pressures from $P = X P°$. Assign the “equal masses” as exactly 1 g.
Liquid:

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

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

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

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

Vapor:

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

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

13.93 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.

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

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

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

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

moles solute = (m)(kg solvent) = (0.878906 m)(0.0249250 kg) = 0.0219067 mol

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

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 = **2.1×10^2 g/mol**.

c) The molar mass of CaN_2O_6 is 164.10 g/mol. This molar mass is less than the 2.1×10^2 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.

$$m = \left(\frac{1.50\text{ g }CaN_2O_6}{25.0\text{ mL}}\right)\left(\frac{\text{mL}}{0.997\text{ g}}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol}}{164.10\text{ g }CaN_2O_6}\right) = 0.3667309\ m\text{ (unrounded)}$$

$\Delta T = iK_bm$

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

13.94 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.

13.95 $$\frac{\text{mol }C_2H_5OH(g)}{\text{mol }CH_3OH(g)} = \frac{\text{mol }C_2H_5OH(l)}{\text{mol }CH_3OH(l)}\left(\frac{60.5\text{ torr}}{126.0\text{ torr}}\right) = \frac{\text{mol }C_2H_5OH(l)}{\text{mol }CH_3OH(l)}(0.4801587)\text{ (unrounded)}$$

A 97:1 mass ratio gives 97 grams of C_2H_5OH for every 1 gram of CH_3OH. (This limits the significant figures.)

$$\frac{97\text{ g }C_2H_5OH(g)\left(\frac{1\text{ mol }C_2H_5OH}{46.07\text{ g }C_2H_5OH}\right)}{1\text{ g }CH_3OH(g)\left(\frac{1\text{ mol }CH_3OH}{32.04\text{ g }CH_3OH}\right)} = \frac{2.10549\text{ mol }C_2H_5OH(g)}{0.03121\text{ mol }CH_3OH(g)}$$

$$\frac{\text{mol }C_2H_5OH(l)}{\text{mol }CH_3OH(l)} = \frac{(2.10549/0.03121)}{0.4801587} = 140.4994$$

$$\frac{(140.4994\text{ mol }C_2H_5OH)\left(\frac{46.07\text{ g }C_2H_5OH}{1\text{ mol }C_2H_5OH}\right)}{(1\text{ mol }CH_3OH)\left(\frac{32.04\text{ g }CH_3OH}{1\text{ mol }CH_3OH}\right)} = 202.0227 = \mathbf{2 \times 10^2}$$

13.96 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 x 10^{-6} = **1.98 x 10^{-6} M CH_3Cl**

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: **1.98 x 10^{-6} m CH_3Cl**

Still using 1.00 L of solution:

Moles CH_3Cl = (1.98098 x 10^{-6} mol/L) (1.00 L) = 1.98098 x 10^{-6} mol CH_3Cl (unrounded)

Moles H_2O = (1.00 kg) (10^3 g/1 kg) (1 mol H_2O / 18.02 g H_2O) = 55.49389567 mol H_2O (unrounded)

$X_{\text{chloroform}}$ = (1.98098 x 10^{-6} mol CH_3Cl) / [(1.98098 x 10^{-6} + 55.49389567) mol]

= 3.569726 x 10^{-8} = **3.57 x 10^{-8}**

13.97 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.

13.98 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.

$$\Pi V = nRT$$

$$\Pi = MRT$$

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

(M)(V) = moles

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

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

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

Mass solvent = mass of solution – mass of solute

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

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

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

$\Delta T_f = iK_fm$ = (1) (1.86°C/m) (1.83466 x 10^{-5} m) = 3.412 x 10^{-5} = **3.41 x 10^{-5}°C**

(So the solution would freeze at 0 – 3.41 x 10^{-5}°C = –3.41 x 10^{-5}°C.)

13.99 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.

13.100 Glyphosate will be abbreviated Gly.

a) Mass Gly = (16.0 fl oz) (1 gal/128 fl oz) (8.94 lb/1 gal) (1 kg/2.205 lb) (10^3 g/1 kg) (18.0%/100%)
= 91.224 = **91.2 g glyphosate**

b) Mass Gly = (3.00 fl oz) (1 gal/128 fl oz) (8.94 lb/1 gal) (1 kg/2.205 lb) (10^3 g/1 kg) (18.0%/100%)
= 17.10459 g glyphosate (unrounded)

Assume that the volume of solution is equal to the volume of solvent, because the volume of glyphosate is insignificant. Assume the density of H_2O is 1.00 g/mL.

Mass of water = (1.00 gal) (3.785 L/1 gal) (1 mL/10^{-3} L) (1.00 g/mL) = 3785 g H_2O (unrounded)

$$\text{Mass percent} = \frac{(17.10459\text{ g Gly})}{(17.10459 + 3785)\text{g}} \text{ x } 100\% = 0.44987 = \mathbf{0.450\%}$$

13.101 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)

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

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

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

$$\text{Mass in water} = \frac{(100.0\text{ mL})}{(100.0 + 8.35(30.0))\text{mL}}(2.853067\text{ mg})$$

= 0.813999 = **0.814 mg** remaining after second extraction

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

13.102 a) $$X_{N_2} = \frac{\text{moles of } N_2}{\text{total moles}}$$

Mixture A: $X_{N_2} = \dfrac{3\text{ moles }N_2}{3 + 2 + 3\text{ moles}} = 0.38$

Mixture B: $X_{N_2} = \dfrac{4\text{ moles }N_2}{4 + 4 + 2\text{ moles}} = 0.40$

Mixture C: $X_{N_2} = \dfrac{4\text{ moles }N_2}{4 + 3 + 5\text{ moles}} = 0.33$

Mixture C has the smallest mole fraction of N_2.

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

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

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

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

c) Mixture A: $X_{Cl_2} = \dfrac{3\text{ moles }Cl_2}{3 + 2 + 3\text{ moles}} = 0.38$

Mixture B: $X_{Cl_2} = \dfrac{2\text{ moles }Cl_2}{4 + 4 + 2\text{ moles}} = 0.20$

Mixture C: $X_{Cl_2} = \dfrac{5\text{ moles }Cl_2}{4 + 3 + 5\text{ moles}} = 0.42$

Mixture B < Mixture A < Mixture C

13.103 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:

Moles C = 32.3 g C (1 mol C/12.01 g C) = 2.6894 mol C (unrounded)
Moles H = 3.97 g H (1 mol H/1.008 g H) = 3.93849 mol H (unrounded)
Moles O = (100 – 32.3 – 3.97) g O (1 mol O/16.00 g O) = 3.9831 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: $\mathbf{C_2H_3O_3}$.

The molar mass comes from the freezing point depression:

$\Delta T_f = iK_fm$ (Assume the compound is a nonelectrolyte, $i = 1$.)

$m = \Delta T_f / iK_f = (1.26°\text{C})/[(1)\ (1.86°\text{C}/m)] = 0.677419\ m$ (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 $\mathbf{C_4H_6O_6}$.

13.104 $$\text{Moles} = (5.66\text{ g NH}_4\text{NO}_3)\left(\frac{1\text{ mol NH}_4\text{NO}_3}{80.05\text{ g NH}_4\text{NO}_3}\right)\left(\frac{1\text{ mol NH}_4^+}{1\text{ mol NH}_4\text{NO}_3}\right) = 7.07058\times10^{-2}\text{ mol NH}_4^+\text{ (unrounded)}$$

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

$= 8.89336 \times 10^{-2}$ mol NH_4^+ (unrounded)

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

$= 2.96445 \times 10^{-2}$ mol PO_4^{3-} (unrounded)

$M\ NH_4^+ = [(7.07058 \times 10^{-2}) + (8.89336 \times 10^{-2})]$ mol $NH_4^+/20.0$ L $= 7.98197 \times 10^{-3} = \mathbf{7.98 \times 10^{-3}\ \mathit{M}\ NH_4^+}$

$M\ PO_4^{3-} = (2.96445 \times 10^{-2}$ mol $PO_4^{3-})/20.0$ L $= 1.482225 \times 10^{-3} = \mathbf{1.48 \times 10^{-3}\ \mathit{M}\ PO_4^{3-}}$

13.105 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 H_3O^+ and Cl^- ions.

13.106 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 **CH_4N_2O**. 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.

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

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

Because the molecular weight equals the empirical weight, the molecular formula is also **CH_4N_2O**.

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

$= 13.512 = \mathbf{14\ g\ glucose}$

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

$= 1.97235 = \mathbf{2.0\ g\ NaCl}$

13.108 a) Use Henry's law:

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

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

$k_H = 1.64 \text{ x } 10^{-3}\ M/0.2095 \text{ atm}$

$k_H = 7.8282 \text{ x } 10^{-3} = \mathbf{7.83 \text{ x } 10^{-3}\ mol/L \bullet atm}$

b) $S_{gas} = k_H \times P_{gas}$

$S_{gas} = (7.8282 \text{ x } 10^{-5} \text{ mol/L} \bullet \text{atm})(0.005 \text{ atm})$

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

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

Oxygen is $4 \text{ x } 10^{-5}$ mol/L or $4 \text{ x } 10^{-5}$ mol in 1.0 L.

$$X_{O_2} = \frac{\text{moles of } O_2}{\text{moles of } O_2 + \text{moles of acrylic acid}} = \frac{4 \text{ x } 10^{-5} \text{ mol}}{((4 \text{ x } 10^{-5}) + 14.6) \text{ mol}} = 2.73972 \text{ x } 10^{-6} = \mathbf{3 \text{ x } 10^{-6}}$$

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

$$\text{Mass of oxygen} = \left(\frac{4 \text{ x } 10^{-5} \text{ mol } O_2}{\text{L}}\right)\left(\frac{32.0 \text{ g } O_2}{1 \text{ mol } O_2}\right) = 0.00128 \text{ g/L}$$

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

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

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

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

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

$m = \Delta T_f / iK_f = (5.0°\text{C}) / [(3)\ (1.86°\text{C}/m)] = 0.896057\ m$ $CaCl_2$ (unrounded)

$$\text{Mass} = \left(\frac{0.896057\ \text{mol CaCl}_2}{\text{kg}}\right)(5.5\ \text{kg})\left(\frac{110.98\ \text{g CaCl}_2}{\text{mol CaCl}_2}\right) = 546.944 = \mathbf{5.5 \times 10^2\ g\ CaCl_2}$$

13.110 a) $$\text{Moles CO}_2 = (355\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{3.3 \text{ x } 10^{-2}\ \text{mol}}{\text{L} \bullet \text{atm}}\right)(4\ \text{atm}) = 0.04686 = \mathbf{0.05\ mol\ CO_2}$$

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:

$$\text{Moles CO}_2 = (355\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{3.3 \text{ x } 10^{-2}\ \text{mol}}{\text{L} \bullet \text{atm}}\right)(3 \text{ x } 10^{-4}\ \text{atm}) = 3.5145 \text{ x } 10^{-6} = \mathbf{4 \times 10^{-6}\ mol\ CO_2}$$

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

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

$= 1.14638 = \mathbf{1\ L\ CO_2}$

13.111 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 + ½(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.