# 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 $CH_3(CH_2)_nCOOH$, 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 **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, even though HCl has a lower boiling point.

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 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 $CH_3CH_2OCH_2CH_3$ 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.

13.14 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.

13.15 **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 -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 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 disulfide linkages from covalent bonds between two sulfur atoms bond cysteine residues together. There may be salt links between ions $-COO^-$ and $-NH_3^+$ groups. There may be hydrogen bonding between the C=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 surrounding, 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.

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 surroundings.

13.21 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 $\Delta H_{\text{mix}}$ 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 ($\Delta H_{\text{mix}}$) would be combined to obtain $\Delta H_{\text{solution}}$.

13.26 a) Charge density is the ratio of the ions 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.

13.27 The solution cycle for ionic compounds in water consists of two enthalpy terms: the negative of the lattice energy, and the combined heats of hydration of the cation and anion.

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

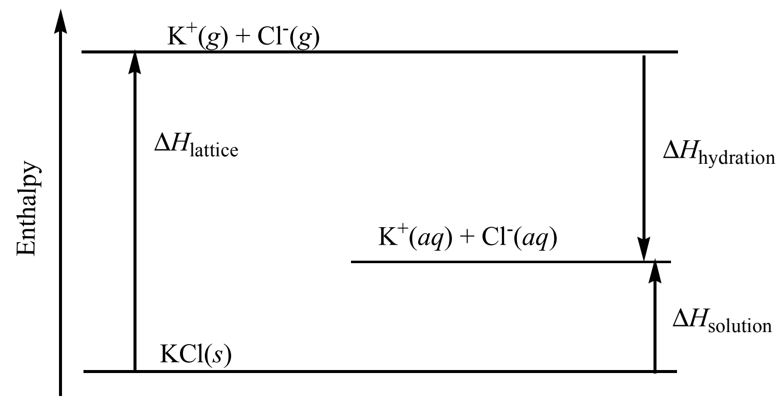
For a heat of solution to be zero (or very small)

$\Delta H_{\text{lattice}} \approx \Delta H_{\text{hydration of ions}}$, and they would have to have the same sign.

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.

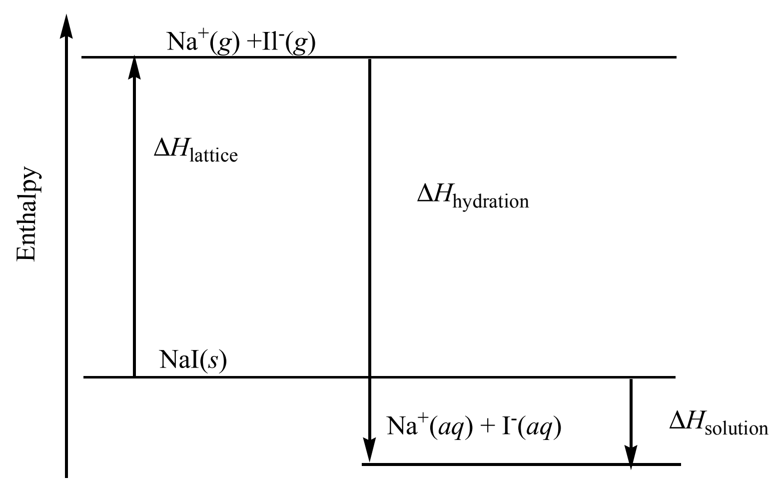
13.30




$\Delta H_{solution} = \Delta H_1 + \Delta H_2 + \Delta H_3$
$\Delta H_{solution} < 0$ (exothermic)
$\Delta H_{solution} > 0$ (endothermic)

You need to know the relative magnitudes of the intermolecular forces in the pure components (solute and solvent) and in the solution. In this problem, the endothermic result is specified.

13.31

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

You need to know the relative magnitudes of the intermolecular forces in the pure components (solute and solvent) and in the solution. In this problem, the exothermic result is specified.

13.32 Charge density is the ratio of an ion's charge (regardless of sign) to its volume.

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.

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

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

13.35 a) $\mathbf{I^-}$ b) $\mathbf{Ca^{2+}}$ c) $\mathbf{Br^-}$ d) $\mathbf{Cl^-}$ e) $\mathbf{Sc^{3+}}$

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

13.37 a) $\Delta H_{\text{hydration of ions}} = \Delta H_{solution} + \Delta H_{lattice}$
$\Delta H_{\text{hydration of ions}}$ = 17.3 kJ/mol + (-763 kJ/mol)
$\Delta H_{hyd}$ = -745.7 = **- 746 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.
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.

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

13.40 $\Delta H_{solution} = \Delta H_{\text{hydration of ions}} - \Delta H_{lattice}$
$\Delta H_{soln}$ = -799 kJ/mol - (-822 kJ/mol)
$\Delta H_{soln}$ = **23 kJ/mol**

13.41 Add a pinch of the solid solute to each solution. The supersaturated solution is unstable and addition of a "seed" crystal of solute causes the excess solute to crystallize immediately, leaving behind a saturated solution. The solution in which the added solid solute dissolves is the unsaturated solution of X. The solution in which the added solid solute remains undissolved is the saturated solution of X.

13.42 $KMnO_4(s) + H_2O(l) + \text{heat} \leftrightarrows KMnO_4(aq)$
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 ($\Delta H_{soln} < 0$).

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, 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: $S_{gas} = k_H \times P_{gas}$. $S_{gas}$ is expressed as mol/L, so convert moles of $O_2$ to mass of $O_2$ using the molar mass.

$$\text{Mass} = \left(1.28 \text{ x } 10^{-3} \frac{\text{mol}}{\text{L} \bullet \text{atm}}\right)(1.00 \text{ atm})\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol}}\right)(2.00 \text{ L}) = 0.08192 = \mathbf{0.0819 \text{ g } O_2}$$

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

$$\text{Mass} = \left(1.28 \text{ x } 10^{-3} \frac{\text{mol}}{\text{L} \bullet \text{atm}}\right)(0.209 \text{ atm})\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol}}\right)(2.00 \text{ L}) = 0.01712 = \mathbf{0.0171 \text{ g } O_2}$$

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

13.48 The solution is **saturated**.

13.49 Solubility for a gas is calculated from Henry's law: $S_{gas} = k_H \times P_{gas}$.
$S_{gas}$ = (3.7 x $10^{-2}$ mol/L•atm) (5.5 atm) = 0.2035 = **0.20 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 does not change with temperature.

13.52 Refer to the table in the text 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. Why is the identity of the solute not necessary for conversion?

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.

a) Molarity = $\left(\frac{42.3 \text{ g } C_{12}H_{22}O_{11}}{100. \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.30 \text{ g } C_{12}H_{22}O_{11}}\right)$ = 1.235758 = **1.24 *M* $C_{12}H_{22}O_{11}$**

b) Molarity = $\left(\frac{5.50 \text{ g LiNO}_3}{505 \text{ mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ mol LiNO}_3}{68.95 \text{ g LiNO}_3}\right)$ = 0.157956 = **0.158 *M* $LiNO_3$**

13.57 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* $NH_3$**

13.58 Dilution calculations can be done using $M_{conc}V_{conc} = M_{dil}V_{dil}$

a) $M_{conc}$ = 0.250 *M* NaOH   $V_{conc}$ = 75.0 mL   $M_{dil}$ = ?   $V_{dil}$ = 0.250 L

$$M_{dil} = M_{conc} V_{conc} / V_{dil} = \frac{(0.250 \text{ M})(75.0 \text{ mL})}{(0.250 \text{ L})}\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = \mathbf{0.0750\ \mathit{M}}$$

b) $M_{conc}$ = 1.3 *M* $HNO_3$   $V_{conc}$ = 35.5 mL   $M_{dil}$ = ?   $V_{dil}$ = 0.150 L

$$M_{dil} = M_{conc} V_{conc} / V_{dil} = \frac{(1.3 \text{ M})(35.5 \text{ mL})}{(0.150 \text{ L})}\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.3076667 = \mathbf{0.31\ \mathit{M}}$$

13.59 Dilution calculations can be done using $M_{conc}V_{conc} = M_{dil}V_{dil}$

a) $M_{conc}$ = 6.15 *M* HCl   $V_{conc}$ = 25.0 mL   $M_{dil}$ = ?   $V_{dil}$ = 0.500 L

$$M_{dil} = M_{conc} V_{conc} / V_{dil} = \frac{(6.15 \text{ M})(25.0 \text{ mL})}{(0.500 \text{ L})}\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) = 0.3075 = \mathbf{0.308\ \mathit{M}}$$

b) $M_{conc}$ = 2.00 x $10^{-2}$ *M* KI   $V_{conc}$ = 8.55 mL   $M_{dil}$ = ?   $V_{dil}$ = 10.0 mL

$$M_{dil} = M_{conc} V_{conc} / V_{dil} = \frac{(2.00 \text{ x } 10^{-2} \text{ M})(8.55 \text{ mL})}{(10.0 \text{ mL})} = \mathbf{0.0171\ \mathit{M}}$$

13.60 a) Find the number of moles $KH_2PO_4$ needed to make 355 mL of this solution. Convert moles to mass using the molar mass of $KH_2PO_4$ (Molar mass = 136.09 g/mol)

$$\text{Mass KH}_2\text{PO}_4 = (355 \text{ mL})\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)\left(\frac{8.74 \text{ x } 10^{-2} \text{ mol KH}_2\text{PO}_4}{\text{L}}\right)\left(\frac{136.09 \text{ g KH}_2\text{PO}_4}{1 \text{ mol KH}_2\text{PO}_4}\right)$$

= 4.22246 = 4.22 g $KH_2PO_4$

Add **4.22 g $KH_2PO_4$** to enough water to make 355 mL of aqueous solution.

b) Use the relationship $M_{conc}V_{conc} = M_{dil}V_{dil}$ to find the volume of 1.25 *M* NaOH needed.

$M_{conc}$ = 1.25 *M* NaOH   $V_{conc}$ = ?   $M_{dil}$ = 0.315 *M* NaOH   $V_{dil}$ = 425 mL

$V_{conc} = M_{dil} V_{dil} / M_{conc}$ = (0.315 *M*) (425 mL) / (1.25 *M*) = 107.1 = 107 mL

Add **107 mL** of 1.25 *M* NaOH to enough water to make 425 mL of solution.

13.61 a) Find the number of moles NaCl needed to make 3.5 L of this solution. Convert moles to mass using the molar mass of NaCl (Molar mass = 58.44 g/mol)

$$\text{Mass NaCl} = (3.5 \text{ L})\left(\frac{0.55 \text{ mol NaCl}}{\text{L}}\right)\left(\frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}}\right) = 112.497 = 1.1 \text{ x } 10^2 \text{ g NaCl}$$

Add **1.1 x $10^2$ g NaCl** to enough water to make 3.5 L of aqueous solution.

b) Use the relationship $M_{conc}V_{conc} = M_{dil}V_{dil}$ to find the volume of 2.2 $M$ urea needed.

$M_{conc}$ = 2.2 $M$ urea   $V_{conc}$ = ?   $M_{dil}$ = 0.3 $M$ urea   $V_{dil}$ = 17.5 L

$V_{conc} = M_{dil} V_{dil} / M_{conc}$ = (0.3 $M$) (17.5 L) / (2.2 $M$) = 2.38636 = 2 L

Add **2 L** of 2.2 $M$ urea to enough water to make 17.5 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 1.50 L of a 0.257 $M$ solution and convert to grams using molar mass of KBr.

$$\text{Mass KBr} = (1.50\ \text{L})\left(\frac{0.257\ \text{mol KBr}}{\text{L}}\right)\left(\frac{119.00\,\text{g KBr}}{1\ \text{mol KBr}}\right) = 45.8745 = 45.9\ \text{g KBr}$$

To make the solution, weigh **45.9 g KBr** and then dilute to 1.50 L with distilled water.

b) To find the volume of the concentrated solution that will be diluted to 355 mL, use $M_{conc}V_{conc} = M_{dil}V_{dil}$ and solve for $V_{conc}$.

$M_{conc}$ = 0.244 $M$ $LiNO_3$   $V_{conc}$ = ?   $M_{dil}$ = 0.0956 $M$ $LiNO_3$   $V_{dil}$ = 355 mL

$V_{conc} = M_{dil} V_{dil} / M_{conc}$ = (0.0956 $M$) (355 mL) / (0.244 $M$) = 139.090 = 139 mL

To make the 0.0956 $M$ solution, measure **139 mL** of the 0.244 $M$ solution and add distilled water to make a total of 355 mL.

13.63 a) To find the mass of $Cr(NO_3)_3$ needed, find the moles of $Cr(NO_3)_3$ in 67.5 mL of a 1.33 x $10^{-3}$ $M$ solution and convert to grams using molar mass of $Cr(NO_3)_3$.

$$\text{Mass Cr(NO}_3)_3 = (67.5\ \text{mL})\left(\frac{10^{-3}\ \text{L}}{1\ \text{mL}}\right)\left(\frac{1.33\ \text{x}\ 10^{-3}\text{mol Cr(NO}_3)_3}{\text{L}}\right)\left(\frac{238.03\ \text{g Cr(NO}_3)_3}{1\ \text{mol Cr(NO}_3)_3}\right)$$

$$= 0.0213691 = 0.0214\ \text{g Cr(NO}_3)_3$$

To make the solution, weigh **0.0214 g $Cr(NO_3)_3$** and then dilute to 67.5 mL with distilled water.

b) To find the volume of the concentrated solution that will be diluted to 3.5 x $10^3$ $m^3$ use $M_{conc}V_{conc} = M_{dil}V_{dil}$ and solve for $V_{conc}$.

$M_{conc}$ = 3.00 $M$ $NH_4NO_3$   $V_{conc}$ = ?   $M_{dil}$ = 1.55 $M$ $NH_4NO_3$   $V_{dil}$ = 6.8 x $10^3$ $m^3$

$V_{conc} = M_{dil} V_{dil} / M_{conc}$ = (1.55 $M$) (6.8 x $10^3$ $m^3$) / (3.00 $M$) = 3513.33 = 3.5 x $10^3$ $m^3$

To make the 1.55 $M$ solution, measure **3.5 x $10^3$ $m^3$** of the 3.00 $M$ solution and add distilled water to make 6.8 x $10^3$ $m^3$.

13.64 Molality = moles solute/kg solvent.

$$a)\ m\ \text{glycine} = \frac{88.4\ \text{g Glycine}\left(\frac{1\ \text{mol Glycine}}{75.07\ \text{g Glycine}}\right)}{(1.250\ \text{kg})} = 0.942054 = \mathbf{0.942}\ \boldsymbol{m}\ \textbf{glycine}$$

$$b)\ m\ \text{glycerol} = \frac{8.89\ \text{g Glycerol}\left(\frac{1\ \text{mol Glycerol}}{92.09\ \text{g Glycerol}}\right)}{(75.0\ \text{g})}\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right) = 1.2871466 = \mathbf{1.29}\ \boldsymbol{m}\ \textbf{glycerol}$$

13.65 Molality = moles solute/kg solvent.

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

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

$$= 2.41548 = \mathbf{2.42}\ \boldsymbol{m}\ \textbf{naphthalene}$$

13.66 Molality = moles solute/kg of solvent.

$$m \text{ benzene} = \frac{(34.0 \text{ mL } C_6H_6)\left(\frac{0.877 \text{ g}}{\text{mL}}\right)\left(\frac{1 \text{ mol } C_6H_6}{78.11 \text{ g } C_6H_6}\right)}{(187 \text{ mL } C_6H_{14})\left(\frac{0.660 \text{ g}}{\text{mL}}\right)}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 3.0930 = \mathbf{3.09}\ \boldsymbol{m}\ \mathbf{C_6H_6}$$

13.67 Molality = moles solute/kg of solvent.

$$m \text{ CCl}_4 = \frac{(2.77 \text{ mL } CCl_4)\left(\frac{1.59 \text{ g}}{\text{mL}}\right)\left(\frac{1 \text{ mol } CCl_4}{153.81 \text{ g } CCl_4}\right)}{(79.5 \text{ mL } CH_2Cl_2)\left(\frac{1.33 \text{ g}}{\text{mL}}\right)}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 0.2708155 = \mathbf{0.271}\ \boldsymbol{m}\ \mathbf{CCl_4}$$

13.68 a) The total weight of the solution is 3.00 x $10^2$ g, so

$\text{mass}_{solute} + \text{mass}_{solvent} = 3.00 \times 10^2$ g and

Grams $C_2H_6O_2$ / 1000 g $H_2O$ in 0.115 *m* = $\left(\frac{0.115 \text{ mol } C_2H_6O_2}{1 \text{ kg } H_2O}\right)\left(\frac{62.07 \text{ g } C_2H_6O_2}{1 \text{ mol } C_2H_6O_2}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$

= 7.13805 g $C_2H_6O_2$ / 1000 g $H_2O$ (unrounded)

Grams of this solution = 1000 g $H_2O$ + 7.13805 g $C_2H_6O_2$ = 1007.13805 (unrounded)

Mass $C_2H_6O_2$ = [7.13805 g $C_2H_6O_2$ / 1007.13805 g solution] (3.00 x $10^2$ g solution) = 2.1262378 = 2.13 g $C_2H_6O_2$

$\text{Mass}_{solvent}$ = 3.00 x $10^2$ g - $\text{mass}_{solute}$ = 3.00 x $10^2$ g - 2.1262378 g $C_2H_6O_2$ = 297.87376 = 2.98 x $10^2$ g $H_2O$

Therefore, **add 2.13 g $C_2H_6O_2$ to 298 g of $H_2O$** to make a 0.115 *m* solution.

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

(1.00 kg) (2.00% / 100%) = 0.0200 kg $HNO_3$ (solute)

Then determine the amount of the concentrated acid solution needed to get 0.0200 kg solute:

(Mass needed) (62.0% / 100%) = 0.0200 kg

Mass solute needed = 0.032258 = 0.0323 kg

Mass solvent = Mass solution - Mass solute = 1.00 kg - 0.032258 kg = 0.96774 = 0.968 kg

Add 0.0323 kg of the 62.0% (w/w) $HNO_3$ to 0.968 kg $H_2O$ to make 1.00 kg of 2.00% (w/w) HNO3.

13.69 a) The total weight of the solution is 1.00 kg, so

$\text{mass}_{solute} + \text{mass}_{solvent}$ = 1.00 kg and

g $C_2H_6O_2$ / 1000g $H_2O$ in 0.0555m = $\left(\frac{0.0555 \text{ mol } C_2H_5OH}{1 \text{ kg } H_2O}\right)\left(\frac{46.07 \text{ g } C_2H_5OH}{1 \text{ mol } C_2H_5OH}\right)\left(\frac{1 \text{ kg}}{10^3 \text{ g}}\right)$

= 2.556885 g $C_2H_5OH$ / 1000 g $H_2O$ (unrounded)

Grams of this solution = 1000 g $H_2O$ + 2.556885 g $C_2H_5OH$ = 1002.556885 g (unrounded)

Mass $C_2H_6O_2$ = [2.556885 g $C_2H_5OOH$ / 1002.556885 g solution] [1.00 kg solution ($10^3$ g / 1 kg)]

= 2.55036 = 2.55 g $C_2H_5OH$

$\text{Mass}_{solvent}$ = 1000 g - $\text{mass}_{solute}$ = 1000 g - 2.55036 g $C_2H_6O_2$ = 997.4496 = 998 g $H_2O$

Therefore, **add 2.55 g $C_2H_5OH$ to 998 g of $H_2O$** to make a 0.115 *m* solution.

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

(475 g) (15.0% / 100%) = 71.25 g HCl (solute) (unrounded)

Then determine the amount of the concentrated acid solution needed to get 0.0200 kg solute:

(Mass needed) (37.1% / 100%) = 74.25 g

Mass solute needed = 192.567586 = 193 g

Mass solvent = Mass solution - Mass solute = 475 g - 192.567586 g = 282.432432 = 282 g

**Add 192 g of the 37.1% (w/w) HCl to 282 g $H_2O$** (Note that the rounding has only given 474 grams of solution. In the laboratory, another gram of water would be added.)

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

$$X_{\text{isopropanol}} = \frac{0.30 \text{ mol Isopropanol}}{(0.30 + 0.80) \text{ mol}} = 0.2727272 = \mathbf{0.27} \text{ (Notice that mole fractions have no units.)}$$

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

Mass isopropanol = (0.30 mol isopropanol) (60.09 g/mol) = 18.027 g isopropanol (unrounded)

Mass water = (0.80 mol water) (18.02 g/mol) = 14.416 g water (unrounded)

$$\text{Percent isopropanol} = \frac{(18.027 \text{ g Isopropanol})}{(18.027 + 14.416) \text{ g}} \text{ x } 100\% = 55.5651 = \mathbf{56\%} \text{ isopropanol}$$

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

$$\text{Molality isopropanol} = \frac{0.30 \text{ mol Isopropanol}}{14.416 \text{ g Water}}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 20.8102 = \mathbf{21}\ \boldsymbol{m} \text{ isopropanol}$$

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

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

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)

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

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

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

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

Mass of water = (0.500 L) (1 mL/$10^{-3}$ L) (1.00 g/mL) (1 kg/$10^3$ g) = 0.500 kg

$$0.400\ m \text{ CsCl} = \frac{\text{moles CsCl}}{0.500 \text{ kg } H_2O}$$

Moles CsCl = 0.2000 mol (unrounded)

Mass of CsCl = (0.2000 mol CsCl) (168.4 g CsCl/1 mol CsCl) = 33.68 = **33.7 g CsCl**

Mass $H_2O$ = (0.500 kg solution) ($10^3$ g/kg) - 33.68 g CsCl = 466.32 g (unrounded)

Moles $H_2O$ = (466.32 g $H_2O$) (1 mol $H_2O$/18.02 g $H_2O$) = 25.87791 mol $H_2O$ (unrounded)

$$X_{\text{CsCl}} = \frac{0.2000 \text{ mol CsCl}}{(0.2000 + 25.87791) \text{mol}} = 7.6693 \text{ x } 10^{-3} = \mathbf{7.67 \text{ x } 10^{-3}}$$

$$\text{Percent CsCl} = \frac{(33.68 \text{ g CsCl})}{(33.68 + 466.32) \text{g}} \text{ x } 100\% = 6.736 = \mathbf{6.74\%} \text{ CsCl}$$

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

Mass of water = (0.400 L) (1 mL/$10^{-3}$ L) (1.00 g/mL) = 4.00 $10^2$ g $H_2O$

Moles $H_2O$ = (400. g $H_2O$) (1 mol $H_2O$/18.02 g $H_2O$) = 22.197558 mol $H_2O$ (unrounded)

Moles KBr = (0.30 g KBr) (1 mol KBr /119.00 g KBr) = 2.5210 x $10^{-3}$ mol KBr (unrounded)

$$X_{\text{KBr}} = \frac{2.5210 \text{ x } 10^{-3} \text{ mol KBr}}{(2.5210 \text{ x } 10^{-3} + 22.197558) \text{mol}} = 1.13558 \text{ x } 10^{-4} = \mathbf{1.1 \text{ x } 10^{-4}}$$

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

13.74 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:

Mass of $NH_3$ = (100.00 grams solution) (8.00% $NH_3$ / 100%) = 8.00 g $NH_3$

Mass $H_2O$ = mass of solution - mass $NH_3$ = (100.00 - 8.00) g = 92.00 g $H_2O$

Moles $NH_3$ = (8.00 g $NH_3$) (1 mol $NH_3$ / 17.03 g $NH_3$) = 0.469759 mol $NH_3$ (unrounded)

Moles $H_2O$ = (92.00 g $H_2O$) (1 mol $H_2O$ / 18.02 g $NH_3$) = 5.1054 mol $H_2O$ (unrounded)

Volume solution = (100.00 g solution) (1 mL / 0.9651 g) ($10^{-3}$ L / 1 mL) = 0.103616 L (unrounded)

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

Molality = Moles solute / 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 = **5.11 *m* $NH_3$**

Molarity = Moles solute / L solution = $\left(\frac{0.469759 \text{ mol NH}_3}{0.103616 \text{ L}}\right)$ = 4.53365 = **4.53 *M* $NH_3$**

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

13.75 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:

Molality = Moles solute / 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 = **2.49 *m* $FeCl_3$**

Molarity = Moles solute / L solution = $\frac{0.1775586 \text{ mol FeCl}_3}{0.078125 \text{ L}}$ = 2.272750 = **2.27 *M* $FeCl_3$**

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

13.76 The mass of 100.0 L of waste water solution is (100.0 L) (1.001 g/mL) (1 mL / $10^{-3}$ L) = 1.001 x $10^5$ g.

ppm $Ca^{2+}$ = [0.22 g $Ca^{2+}$ / 1.001 x $10^5$ g] ($10^6$) = 2.1978 = **2.2 ppm $Ca^{2+}$**

ppm $Mg^{2+}$ = [0.066 g $Mg^{2+}$ / 1.001 x $10^5$ g] ($10^6$) = 0.65934 = **0.66 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:

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 EG/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 = $\frac{17.94747865 \text{ mol EC}}{1.97570 \text{ L}}$ = 9.08411 = **9.08 *M* ethylene glycol**

d) Molality = Moles solute / kg solvent = $\frac{17.94747865 \text{ mol EG}}{1.00 \text{ x } 10^3 \text{g } H_2O}\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)$

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

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

= 0.244378 = **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 composition of the solute.

13.79 A nonvolatile nonelectrolyte is a covalently bonded molecule that does not dissociate 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 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 concentration deviates from the molar concentration. Thus, **0.050 *m* NaF** 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 $\Delta T_f$ would be greater.

13.88 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.

13.89 a) $NaMnO_4$ **strong electrolyte**
b) $CH_3COOH$ **weak electrolyte**
c) $CH_3OH$ **nonelectrolyte**
d) $Ca(C_2H_3O_2)_2$ **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) (0.2 mol KI/L) (2 mol particles/mol KI) (1 L) = **0.4 mol of particles**
KI consists of $K^+$ ions and $I^-$ ions, 2 particles for each KI.
b) (0.070 mol $HNO_3$/L) (2 mol particles/mol $HNO_3$) (1 L) = **0.14 mol of particles**
$HNO_3$ is a strong acid that forms hydronium ions and nitrate ions in aqueous solution.
c) ($10^{-4}$ mol $K_2SO_4$/L) (3 mol particles/mol $K_2SO_4$) (1 L) = **3 x $10^{-4}$ mol of particles**
Each $K_2SO_4$ forms 2 potassium ions and 1 sulfate ion in aqueous solution.
d) (0.07 mol $C_2H_5OH$/L) (1 mol particles/mol $C_2H_5OH$) (1 L) = **0.07 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, **0.07 mol** in 1 L.

13.91 a) (0.01 mol $CuSO_4$/L) (2 mol particles/mol $CuSO_4$) ($10^{-3}$ L / 1 mL) (1 mL) = **2 x $10^{-5}$ 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 x $10^{-5}$ = **2 x $10^{-5}$ 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 x $10^{-5}$ mol of particles**
d) (0.05 mol $(NH_4)_3CO_3$/L) (3 mol particles/mol $(NH_4)_2CO_3$) ($10^{-3}$ L / 1 mL) (1 mL)
= 1.5 x $10^{-4}$ = **2 x $10^{-4}$ mol of particles**

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

a) Molality of $CH_3OH$ = $\frac{(10.0 \text{ g } CH_3OH)}{(100. \text{ g } H_2O)}\left(\frac{1 \text{ mol } CH_3OH}{32.04 \text{ g } CH_3OH}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)$ = 3.1210986 = 3.12 *m* $CH_3OH$

Molality of $CH_3CH_2OH$ = $\frac{(20.0 \text{ g } CH_3CH_2OH)}{(200. \text{ g } H_2O)}\left(\frac{1 \text{ mol } CH_3CH_2OH}{46.07 \text{ g } CH_3CH_2OH}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)$
= 2.1706 = 2.17 *m* $CH_3CH_2OH$

The molality of methanol, $CH_3OH$, in water is 3.12 *m* whereas the molality of ethanol, $CH_3CH_2OH$, in water is 2.17 m. Thus, **$CH_3OH/H_2O$ solution** has the lower freezing point.

b) Molality of $H_2O$ = $\frac{(10.0 \text{ g } H_2O)}{(1.00 \text{ kg } CH_3OH)}\left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right)$ = 0.5549 = 0.55 *m* $H_2O$

Molality of $CH_3CH_2OH$ = $\frac{(10.0 \text{ g } CH_3CH_2OH)}{(1.00 \text{ kg } CH_3OH)}\left(\frac{1 \text{ mol } CH_3CH_2OH}{46.07 \text{ g } CH_3CH_2OH}\right)$ = 0.21706 = 0.217 *m* $CH_3CH_2OH$

The molality of $H_2O$ in $CH_3OH$ is 0.555 *m*, whereas $CH_3CH_2OH$ in $CH_3OH$ is 0.217 *m*. Therefore, **$H_2O/CH_3OH$ solution** has the lower freezing point.

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

a) Molality of $C_3H_8O_3$ = $\frac{(35.0 \text{ g } C_3H_8O_3)}{(250. \text{ g Ethanol})}\left(\frac{1 \text{ mol } C_3H_8O_3}{90.09 \text{ g } C_3H_8O_3}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)$ = 1.5540 = 1.55 *m* $C_3H_8O_3$

Molality of $C_2H_6O_2$ = $\frac{(35.0 \text{ g } C_2H_6O_2)}{(250. \text{ g Ethanol})}\left(\frac{1 \text{ mol } C_2H_6O_2}{62.07 \text{ g } C_2H_6O_2}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)$ = 2.2555 = 2.26 *m* $C_2H_6O_2$

The molality of $C_2H_6O_2$, in ethanol is 2.26 *m* whereas the molality of $C_3H_8O_3$, in ethanol is 1.55 m. Thus, **$C_2H_6O_2$/ethanol solution** has the higher boiling point.

b) Molality of $C_2H_6O_2$ = $\frac{(20.\ g\ C_2H_6O_2)}{(0.50\ kg\ H_2O)}\left(\frac{1\ mol\ C_2H_6O_2}{62.07\ g\ C_2H_6O_2}\right)$ = 0.64443 = 0.64 *m* $C_2H_6O_2$

Molality of NaCl = $\frac{(20.\ g\ NaCl)}{(0.50\ kg\ H_2O)}\left(\frac{1\ mol\ NaCl}{58.44\ g\ NaCl}\right)$ = 0.68446 = 0.68 *m* NaCl

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

(2 particles/NaCl) (0.68446 mol NaCl/kg) = 1.36892 = 1.4 *m* particles

The molality of $C_2H_6O_2$ in $H_2O$ is 0.64 m, whereas NaCl in $H_2O$ is 1.4 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.

(I) 0.100 *m* $NaNO_3 \times$ 2 mol particles/mol $NaNO_3$ = 0.200 *m* ions

(II) 0.200 *m* glucose $\times$ 1 mol particles/mol glucose = 0.200 *m* molecules

(III) 0.100 *m* $CaCl_2 \times$ 3 mol particles/mol $CaCl_2$ = 0.300 *m* ions

a) Osmotic pressure: $\pi_I = \pi_{II} < \pi_{III}$

b) Boiling point: $bp_I = bp_{II} < bp_{III}$

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 0.02 *m* $AgNO_3$ x 2 mol particles / 1 mol $AgNO_3$ = 0.04 *m* ions

III 0.02 *m* $CuSO_4$ x 2 mol particles / 1 mol $CuSO_4$ = 0.04 *m* ions

All have 0.04 *m* particle concentrations and have the same colligative properties.

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

Moles $C_3H_8O_3$ = (44.0 g $C_3H_8O_3$) (1 mol $C_3H_8O_3$ / 92.09 g $C_3H_8O_3$)
= 0.47779 mol $C_3H_8O_3$ (unrounded)

Moles $H_2O$ = (500.0 g $H_2O$) (1 mol $H_2O$ / 18.02 g $H_2O$) = 27.7469 mol $H_2O$ (unrounded)

$X_{solvent}$ = (27.7469 mol $H_2O$) / [(0.47779) + (27.7469)] mol = 0.98307 (unrounded)

$P_{solvent} = X_{solvent}P°_{solvent}$ = (0.98307) (23.76 torr) = 23.35774 = **23.36 torr**

13.97 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) (41torr) = 38.2382 = **38 torr**

13.98 The change in freezing point is calculated from $\Delta T_f = iK_f m$, 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, i = 1. Once $\Delta T_f$ 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*) (0.111 *m*) = 0.20646°C (unrounded)

The freezing point is 0.00°C - 0.20646°C = -0.20646 = **-0.206°C**

13.99 $\Delta T_b = iK_b\, m$ = (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**

13.100 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*).

$$\text{Molality of Vanillin} = \frac{(3.4\text{ g Vanillin})\left(\frac{1\text{ mol Vanillin}}{152.14\text{ g Vanillin}}\right)}{(50.0\text{ g Ethanol})}\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)$$

= 0.44695675 *m* Vanillin (unrounded)

$\Delta T_b = iK_b\, m$ = (1) (1.22°C/*m*) (0.44695675 *m*) = 0.545287°C (unrounded)

The boiling point is 78.5°C + 0.545287°C = 79.045287 = **79.0°C**

13.101 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**

13.102 The molality of the solution can be determined from the relationship $\Delta T_f = iK_fm$ with the value 1.86°C/*m* inserted for $K_f$, i = 1 for the nonelectrolyte ethylene glycol, and the given $\Delta T_f$ of -10.0°F converted to °C. Multiply the molality by the given mass of solvent to find the mass of ethylene glycol that must be in solution. Note that ethylene glycol is a covalent compound that will form one particle per molecule when dissolved.

°C = (5/9) (°F - 32.0) = (5/9) ((-10.0)°F - 32.0) = -23.3333°C (unrounded)

$\Delta T_f$ = (0.00 - (-23.3333))°C = 23.3333°C

*m* = $\Delta T_f / K_f$ = (23.3333°C) / (1.86°C/*m*) = 12.54478 *m* (unrounded)

Ethylene glycol will be abbreviated as EG

$$\text{Mass ethylene glycol} = \left(\frac{12.54478\text{ mol EG}}{1\text{ kg }H_2O}\right)(14.5\text{ kg }H_2O)\left(\frac{62.07\text{ g EG}}{1\text{ mol EG}}\right)$$

= 1.129049 x $10^4$ = **1.13 x $10^4$ g ethylene glycol**

To prevent the solution from freezing, dissolve a minimum of 1.13 x $10^4$ 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°C/m inserted for $K_f$, i = 1 for the nonelectrolyte glycerol, and the given $\Delta T_f$ of -25°C.

m = $\Delta T_f / K_f$ = (25°C) / (1.86°C/m) = 13.44086 *m* (unrounded)

Glycerol will be abbreviated as GLY

$$\text{Mass glycerol} = \left(\frac{13.44086\text{ mol GLY}}{1\text{ kg }H_2O}\right)(11.0\text{ mg }H_2O)\left(\frac{10^{-3}\text{ g}}{1\text{ mg}}\right)\left(\frac{1\text{ kg}}{10^3\text{ g}}\right)\left(\frac{92.09\text{ g GLY}}{1\text{ mol GLY}}\right)$$

= 0.013615 = **0.014 g glycerol**

To prevent the solution from freezing, dissolve a minimum of 0.014 g glycerol in 11.0 mg water.

13.104 Convert the mass percent to molality and use $\Delta T = iK_bm$ to find the expected freezing point depression.

a) Assume exactly 100 grams of solution. Thus, the solution contains 1.00 grams of NaCl in 99.00 grams of water.

$$\text{Molality NaCl} = \left(\frac{1.00\text{ g NaCl}}{99.00\text{ g }H_2O}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol NaCl}}{58.44\text{ g NaCl}}\right) = 0.172844 = \mathbf{0.173\ \textit{m}\ NaCl}$$

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

$\Delta T_f = iK_f\, m$

i = $\Delta T_f / K_f\, m$ = (0.593°C) / [(1.86°C/*m*) (0.172844 *m*)] = 1.844537 = **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. Thus, the solution contains 0.500 grams of $CH_3COOH$ in 99.500 grams of water.

$$\text{Molality } CH_3COOH = \left(\frac{0.500\text{ g } CH_3COOH}{99.500\text{ g } H_2O}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol } CH_3COOH}{60.05\text{ g } CH_3COOH}\right)$$

= 0.083682 = **0.0837 *m* $CH_3COOH$**

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

$\Delta T_f = iK_f\, m$

$i = \Delta T_f / K_f\, m = (0.159°C) / [(1.86°C/m)\,(0.083682\; m)] = 1.02153 = \mathbf{1.02}$

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.

13.105 Convert the mass % to molality and use $\Delta T = iK_b m$ to find the expected freezing point depression.

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 $\Delta T = (0.000 - (-0.234))°C = 0.234°C$

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

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

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 $\Delta T = (0.000 - (-0.423))°C = 0.423°C$

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

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

Sulfuric acid is a strong acid and dissociates to give a hydrogen ions and a hydrogen sulfate ions. The hydrogen sulfate ions may further dissociate to more hydrogen ions and sulfate ions. 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.)K = 293 K

Molarity = (3.42 g sucrose/L) (1 mol sucrose / 342.30 g sucrose)

= $9.9912 \times 10^{-3}$ *M* sucrose (unrounded)

$\Pi = i\,MRT = (1)\,(9.9912 \times 10^{-3}\text{ mol/L})\,(0.0821\text{ L•atm/mol•K})\,(293\text{ K}) = 0.24034 = \mathbf{0.240\ atm}$

A pressure greater than 0.240 atm must be applied to obtain pure water from a 3.42 g/L solution.

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

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

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) = $5.4445 \times 10^3$ = **$5.44 \times 10^3$ g/mol**

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

$X_{CH_2Cl_2}$ = (1.50 mol) / [(1.50 + 1.00)mol] = 0.600

$X_{CCl_4}$ = (1.00 mol) / [(1.50 + 1.00)mol] = 0.400

$P_A = X_A P°_A$

= (0.600) (352 torr) = 211.2 = **211 torr $CH_2Cl_2$**

= (0.400) (118 torr) = **47.2 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 Foam comes from gas bubbles trapped in the cream. The best gas for foam would be one that is not very soluble in the cream. Since nitrous oxide is a good gas to make foam and carbon dioxide a bad one, nitrous oxide must be less soluble in the cream than carbon dioxide. (In fact, at room temperature, $CO_2$ is almost twice as soluble in water as $N_2O$.) Henry's law constant is larger for a gas that is more soluble, so the constant for carbon dioxide must be larger than the constant for nitrous oxide.

13.115 Assume exactly 100 grams of solution.

Thus, the solution contains

(100 g) (10% glucose / 100%) = 10. g glucose

(10. g glucose) (1 mol glucose / 180.16 g glucose) = 0.055506216 mol glucose (unrounded)

and 90. g of water

(90. g $H_2O$) (1 kg / $10^3$ g) = 0.090 kg

The volume of the solution is

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

Molarity glucose = (0.055506216 mol glucose) / (0.096246 L) = 0.57671 *M* glucose (unrounded)

Molality glucose = (0.055506216 mol glucose) / (0.090 kg) = 0.6167357 *m* glucose (unrounded)

Glucose is a nonelectrolyte so i = 1.

T = (273 + 20) = 293 K

$\Delta T_f = iK_f m$ = (1) (1.86°C/m) (0.6167357 m) = 1.1471°C

Freezing point = (0.00 - 1.1471) = - 1.1471 = **- 1.1°C**

$\Delta T_b = iK_b m$ = (1) (0.512°C/m) (0.6167357 m) = 0.3157687°C

Boiling point = (100.00 + 0.3157687) = 100.3157687 = **100.32°C**

$\Pi = iMRT$ = (1) (0.57671 mol/L) (0.0821 L•atm/mol•K) (293 K) = 13.8729 = **14 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 (65.39 / 63.55) must be applied to the zinc contribution.

a) Density of alloy = (90.0% Cu / 100%) (8.95 $g/cm^3$) + (10.0% Zn / 100%) (65.39 / 63.55) (8.95 $g/cm^3$)
= 8.9759 = **8.98 $g/cm^3$**

b) Density of alloy = (62.0% Cu / 100%) (8.95 $g/cm^3$) + (38.0% Zn / 100%) (65.39 / 63.55) (8.95 $g/cm^3$)
= 9.04847 = **9.05 $g/cm^3$**

13.117 **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.

13.118 $P_A = X_A P°_A$
= (0.14) (11 torr) = 1.54 = **1.5 torr octane**

13.119 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.

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

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

$$\text{Mass of seawater} = \left[\frac{31.1 \text{ g}}{1.1 \times 10^{-2}} \times 10^9\right]\left(\frac{100\%}{79.5\%}\right) = 3.5563 \times 10^{12} \text{ g seawater (unrounded)}$$

Volume seawater = ($3.5563 \times 10^{12}$ g) (1 mL / 1.025 g) ($10^{-3}$ L / 1 mL) = $3.46956 \times 10^9$ = **$3.5 \times 10^9$ L**

13.120 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.

13.121 0.0°C

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

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

20.0°C

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

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

40.0°C

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

= $1.996883 \times 10^{-4}$ = **$2.00 \times 10^{-4}$ 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 Price NaCl (\$/ion) = (\$0.22 / kg NaCl) (1 kg / $10^3$ g) (58.44 g NaCl / 1 mol NaCl) (1 mol NaCl / 2 ions)
= \$ $6.4284 \times 10^{-3}$ / ion (unrounded)

Price $CaCl_2$ (\$/kg) = (\$ $6.4284 \times 10^{-3}$ / 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 $CaCl_2$**

13.124 Mass CO = (4.0 x $10^{-6}$ mol/L) (12 L/min) (60 min/1 h) (8.0 h) (28.01 g CO/mol) = 0.645350 = **0.65 g CO**

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

13.126 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

$m = \Delta T_f / iK_f$ = (0.201°C) / [(1) (1.86°C/m)] = 0.1080645 $m$ (unrounded)

$$\text{Molar mass} = \left(\frac{0.243\text{ g}}{25.0\text{ mL}}\right)\left(\frac{\text{mL}}{1.00\text{ g}}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ kg}}{0.108065\text{ mol}}\right) = 89.94586591 = \mathbf{89.9\ g/mol}$$

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.

Moles C = (53.31 g C) (1 mol C / 12.01 g C) = 4.43880 mol C (unrounded)

Moles H = (11.18 g C) (1 mol H / 1.008 g H) = 11.09127 mol H (unrounded)

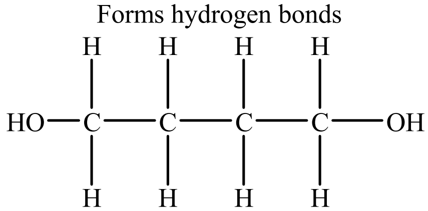
Moles O = (35.51 g O) (1 mol O / 16.00 g O) = 2.219375 mol O (unrounded)

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.

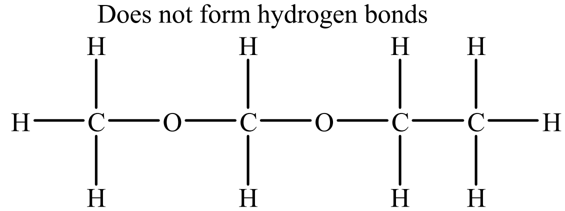
Since the molar mass of the compound is twice the molar mass of the empirical formula, the **molecular formula is $C_4H_{10}O_2$**.

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

Forms hydrogen bonds



Does not form hydrogen bonds



13.127 a) $\left(\frac{5.00 \times 10^{-5}\text{ mol F}^-}{\text{L}}\right)(5000.\text{ L})\left(\frac{1\text{ mol NaF}}{1\text{ mol F}^-}\right)\left(\frac{41.99\text{ g NaF}}{1\text{ mol NaF}}\right)$ = 10.4975 = **10.5 g NaF**

b) $\left(\frac{5.00 \times 10^{-5}\text{ mol F}^-}{\text{L}}\right)(2.0\text{ L})\left(\frac{19.00\text{ g F}^-}{1\text{ mol F}}\right)$ = **0.0019 g $F^-$**

13.128 a) To shorten the settling time, lime (CaO) and cake alum ($Al_2(SO_4)_3$) are added to form a fluffy, gel-like precipitate of $Al(OH)_3$.

b) Water that contains large amounts of divalent cations (such as $Ca^{2+}$, $Mg^{2+}$, and $Fe^{2+}$) 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 $\Delta H_{hydr}$ increases with increasing charge density.

a) **$Mg^{2+}$** has the higher charge density because it has a smaller ion volume.

b) **$Mg^{2+}$** has the higher charge density because it has both a smaller ion volume and greater charge.

c) **$CO_3^{2-}$** has the higher charge density for the same reasons in b).

d) **$SO_4^{2-}$** has the higher charge density for the same reasons in b).

e) **$Fe^{3+}$** has the higher charge density for the same reasons in b).

f) **$Ca^{2+}$** has the higher charge density for the same reasons in b).

13.130 Calculate the individual partial pressures from P = X P°. 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 (pinene) = (0.53100) (100.3 torr) = 53.2593 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 (terpineol) = (0.4689985) (9.8 torr) = 4.5961855 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.

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

$m = \Delta T / iK_b$ = (100.45 - 100.00)°C / [(1) (0.512°C)] = 0.878906 $m$ = 0.878906 mol/kg (unrounded)

$$\left(\frac{1.50\text{ g}}{25.0\text{ mL}}\right)\left(\frac{\text{mL}}{0.997\text{ g}}\right)\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{\text{kg}}{0.878906\text{ mol}}\right) = 68.4721 = \mathbf{68\ g/mol}$$

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 68.4721 = 205.416 = **2.1 x $10^2$ g/mol**.

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

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.

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

$\Delta T = iK_b m$

$$i = \Delta T / K_b\, m = \frac{((100.45 - 100.00)°\text{C})}{(0.512°/\text{m})(0.3667309\text{ m})} = 2.396597 = \mathbf{2.4}$$

13.132 $\frac{\text{mol } C_2H_5OH(g)}{\text{mol } CH_3OH(g)} = \frac{\text{mol } C_2H_5OH(l)}{\text{mol } CH_3OH(l)}\left(\frac{60.5 \text{ torr}}{126.0 \text{ torr}}\right) = \frac{\text{mol } C_2H_5OH(l)}{\text{mol } CH_3OH(l)}$ (0.4801587) (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.)

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

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

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

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

$$\left(\frac{100. \text{ g}}{10^9 \text{ g Solution}}\right)\left(\frac{1 \text{ mol } CH_3Cl}{50.48 \text{ g } CH_3Cl}\right)\left(\frac{1.00 \text{ g Solution}}{\text{mL}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)$$

$= 1.98098 \times 10^{-6} = \mathbf{1.98 \times 10^{-6}}$ ***M*** $\mathbf{CH_3Cl}$

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: $\mathbf{1.98 \times 10^{-6}}$ ***m*** $\mathbf{CH_3Cl}$

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)

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

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

13.134 a) **Yes**, equilibrium is a dynamic process.

b) Radioactivity would be found in all the solid.

$Na_2{}^{14}CO_3(s) + H_2O(l) \rightarrow Na_2{}^{14}CO_3(aq)$

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.

$\Pi V = nRT$

$\Pi = (n/V)RT = MRT$

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

$$\text{Molar Mass} = \left(\frac{10.0 \text{ mg}}{30.0 \text{ mL}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mL}}{10^{-3} \text{ L}}\right)\left(\frac{1 \text{ L}}{1.828546 \times 10^{-5} \text{ mol}}\right)$$

$= 1.82294 \times 10^4 = \mathbf{1.82 \times 10^4 \text{ g/mol}}$

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 = [(30.0 mL) (0.997 g/mL) - (10.0 mg) ($10^{-3}$ g / 1 mg)] (1 kg / $10^3$ g) = 0.0299 kg

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 x $10^{-7}$ mol (unrounded)

Molality = (5.485638 x $10^{-7}$ mol) / (0.0299 kg) = 1.83466 x $10^{-5}$ *m* (unrounded)
$\Delta T_f = iK_f m$ = (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 -3.41 x $10^{-5}$°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.

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

$$P_{gas} = S_{gas} / k_H = \left(\frac{0.75 \text{ mg/L}}{0.033 \text{ mol/L} \bullet \text{atm}}\right)\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right)\left(\frac{1 \text{ mol } C_2H_2Cl_2}{96.94 \text{ g } C_2H_2Cl_2}\right)$$

= 2.3446799 x $10^{-4}$ atm (unrounded)

PV = nRT

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

(9.7130 x $10^{-6}$ mol/L) (96.94 g $C_2H_2Cl_2$ / mol $C_2H_2Cl_2$) (1 ng / $10^{-9}$ g) = 9.41578 x $10^5$ = **9.4 x $10^5$ ng/L**

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

$$\frac{1.17021 \text{ x } 10^3 \text{cm}^2}{2.50 \text{ mg}}\left(\frac{1 \text{ mg}}{10^{-3} \text{ g}}\right)\left(\frac{283 \text{ g}}{\text{mol}}\right)\left(\frac{1 \text{ mol}}{6.022 \text{ x } 10^{23} \text{ molecules}}\right)$$

= 2.19973 x $10^{-16}$ = **2.20 x $10^{-16}$ $cm^2$/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 x $10^3$ ppm solutions and 14 ± 1.7 units for the 10.00 x $10^3$ 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 x $10^3$ ppm) = 8.75 x $10^3$ ppm

Assume the mass of $CaCl_2$ present is negligible relative to the mass of the solution.

$$\text{Molality } CaCl_2 = \left(\frac{8.75 \text{ x } 10^3 \text{ g } CaCl_2}{10^6 \text{ g Solution}}\right)\left(\frac{1 \text{ mol } CaCl_2}{110.98 \text{ g } CaCl_2}\right)\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) = 0.0788430 = \mathbf{0.0788\ \textit{m}\ CaCl_2}$$

Moles $CaCl_2$ = (8.75 x $10^3$ g $CaCl_2$) (1 mol $CaCl_2$ / 110.98 g $CaCl_2$) = 78.84303 mol $CaCl_2$ (unrounded)
Moles $H_2O$ = (1.00 x $10^6$ g $H_2O$) (1 mol $H_2O$ / 18.02 g $H_2O$) = 5.5493896 x $10^4$ mol $H_2O$ (unrounded)

$$\text{Mole fraction } CaCl_2 = X = \frac{\left(78.84303 \text{ mol } CaCl_2\right)}{\left(78.84303 + 5.5493896 \text{ x } 10^4\right)\text{mol}} = 1.4187 \text{ x } 10^{-3} = \mathbf{1.42 \text{ x } 10^{-3}}$$

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 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)

$$\text{Mass percent} = \frac{(17.10459\ \text{g Gly})}{(17.10459 + 3785)\text{g}} \times 100\% = 0.44987 = \mathbf{0.0450\%}$$

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)

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

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.

13.142 Molality is defined as moles of solute per kg of solvent, so 0.150 *m* means 0.150 mol $NaHCO_3$ per kg of water. The total mass of the solution would be 1 kg + 0.150 mol x molar mass of $NaHCO_3$.

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

= 12.6015 g $NaHCO_3$ / 1000 g solvent

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

Grams $H_2O$ = 250. g - 3.111 g = 246.889 g $H_2O$

To make 250. g of a 0.150 *m* solution of $NaHCO_3$, **weigh 3.11 g $NaHCO_3$ 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:

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:

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

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

13.144 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 g methanol / 2.000 g solution) = (1.000 g ethanol / 2.000 g solution)

Moles methanol = (1.000 g ethanol) (1 mol ethanol / 32.04 g ethanol) = 0.031210986 mol methanol (unrounded)

Moles ethanol = (1.000 g ethanol) (1 mol ethanol / 46.07 g ethanol) = 0.021706 mol ethanol (unrounded)

Mole fraction methanol = (0.031210986 mol methanol) / [(0.031210986) + (0.021706)] mol
= 0.589810 (unrounded)

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

13.145 a) The molar mass comes from the boiling point elevation:

The boiling point and elevation constant values come from Table 13.6.

$\Delta T_b = (77.5 - 76.5) = 1.0°C$

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

$m = \Delta T_b / iK_b = (1.0°C) / [(1)(5.03°C/m)] = 0.198807\ m$ (unrounded)

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

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:
(mol solute/L solution) = (mol solute/kg solvent) (kg solvent/L solution) = $M = m$ (kg solvent/L solution)
For a very dilute solution, the assumption that mass of solvent ≅ mass of solution is valid. This equation then becomes

$M = m$ (kg solvent/L solution) = $m$ x $d_{solution}$

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

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

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

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

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

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

$M$ $PO_4^{3-}$ = (2.96445 x $10^{-2}$ mol $PO_4^{3-}$) / 20.0 L = 1.482225 x $10^{-3}$ = **1.48 x $10^{-3}$ $M$ $PO_4^{3-}$**

13.148 a) $M$ $SO_2$ = P $k_H$ = (2.0 x $10^{-3}$ atm) (1.23 mol/L•atm) = 2.46 x $10^{-3}$ = **1.5 x $10^{-3}$ $M$ $SO_2$**
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.
Moles C = 20.1 g C (1 mol C / 12.01 g C) = 1.6736 mol C (unrounded)
Moles H = 6.7 g H (1 mol H / 1.008 g H) = 6.6468 mol H (unrounded)
Moles N = 46.5 g N (1 mol N / 14.01 g N) = 3.31906 mol N (unrounded)
Moles O = (100 - 20.1 - 6.7 - 46.5) g O (1 mol O / 16.00 g O) = 1.66875 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 $\Pi = M\text{RT}$ 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.

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

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

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

13.150 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 = **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) 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 = **2.0 g NaCl**

13.151 Total pressure = $P_{nitrogen} + P_{oxygen} + P_{helium}$ = 4.00 atm

$M = P\,k_H$

At 1.00 atm (this is the partial pressure of each gas, and not the partial pressure of each gas in air at 1.00 atm. The problem does not specify air, but $O_2$ and $N_2$ at 1.00 atm.)

$M_{oxygen}$ = (1.00 atm) (1.1 x $10^{-3}$ mol/L•atm) = 1.1 x $10^{-3}$ *M* oxygen

$M_{nitrogen}$ = (1.00 atm) (6.4 x $10^{-4}$ mol/L•atm) = 6.4 x $10^{-4}$ *M* nitrogen

At 4.00 atm total pressure

$P_{oxygen} = M / k_H$ = the same pressure as the original gas.

$P_{helium}$ must be equal to the total pressure minus the sum of the partial pressures of the other gases or **2.00 atm**.

13.152 a) The solubility of a gas is proportional to its partial pressure. If the solubility at a pressure of 0.10 MPa is 0.147 $cm^3$, then the solubility at a pressure that is 0.6 of 0.10 MPa will be 0.6 of 0.147 $cm^3$.

[(0.147 $cm^3$/g $H_2O$) / 0.10 MPa] (0.060 MPa) = 0.0882 = **0.088 $cm^3$/g $H_2O$**

b) Henry's law states that $S_{gas} = k_H \times P_{gas}$.

$k_H$ = (0.147 $cm^3$/g $H_2O$) / 0.10 MPa = 1.47 = **1.5 $cm^3$/g $H_2O$•MPa**

c) At 5.0 MPa, Henry's law would give a solubility of

(1.47 $cm^3$/g $H_2O$•MPa) (0.50 MPa) = 0.735 = **0.74 $cm^3$/g $H_2O$**

Percent error = [(0.825 - 0.735) / 0.825] (100) = 10.909 = **11%**

An error of approximately 11% occurs when Henry's law is used to calculate solubility at high pressures around 1 MPa.

13.153 Mass percents:

Iodine in chloroform

[2.7 g $I_2$ / (2.7 + 100.0)g] x 100% = 2.6290 = **2.6% $I_2$**

Iodine in carbon tetrachloride

[2.5 g $I_2$ / (2.5 + 100.0)g] x 100% = 2.4390 = **2.4% $I_2$**

Iodine in carbon disulfide

[16 g $I_2$ / (16 + 100.0)g] x 100% = 13.793 = **14% $I_2$**

Mole fraction:

Iodine in chloroform

Moles $I_2$ = (2.7 g $I_2$) (1 mol $I_2$ / 253.8 g $I_2$) = 0.010638 mol $I_2$ (unrounded)

Moles solvent = (100.0 g $CHCl_3$) (1 mol $CHCl_3$ / 119.37 g $CHCl_3$)

= 0.8377314 mol $CHCl_3$(unrounded)

Mole fraction = (0.010638 mol $I_2$) / [(0.010638) + (0.8377314)] mol = 0.01245966 = **0.012**

Iodine in carbon tetrachloride

Moles $I_2$ = (2.5 g $I_2$) (1 mol $I_2$ / 253.8 g $I_2$) = 0.009850 mol $I_2$ (unrounded)

Moles solvent = (100.0 g $CCl_4$) (1 mol $CCl_4$ / 153.81 g $CCl_4$)

= 0.650152785 mol $CCl_4$ (unrounded)

Mole fraction = (0.009850 mol $I_2$) / [(0.009850) + (0.650152785)] mol = 0.014924 = **0.015**

Iodine in carbon disulfide

Moles $I_2$ = (16 g $I_2$) (1 mol $I_2$ / 253.8 g $I_2$) = 0.0630418 mol $I_2$ (unrounded)

Moles solvent = (100.0 g $CS_2$) (1 mol $CS_2$ / 76.15 g $CS_2$) = 1.3131976 mol $CS_2$ (unrounded)

Mole fraction = (0.0630418 mol $I_2$) / [(0.0630418) + (1.3131976)] mol = 0.045807 = **0.046**

Molality:

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

Iodine in chloroform

Molality = (0.010638 mol $I_2$) / (0.1000 kg) = 0.10638 = **0.11 *m* $I_2$**

Iodine in carbon tetrachloride

Molality = (0.009850 mol $I_2$) / (0.1000 kg) = 0.09850 = **0.098 *m* $I_2$**

Iodine in carbon disulfide

Molality = (0.0630418 mol $I_2$) / (0.1000 kg) = 0.630418 = **0.63 *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:

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

$P_1$ = 1.00 atm  $T_1$ = (273 + 100) K = 373 K

$P_2$ = ?  $T_2$ = (273 + 200.)K = 473 K

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

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

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

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

$P_2$ = 16.03365 = **16.0 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°C) / [(2) (1.86°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°C) / [(3) (1.86°C/m)] = 0.896057 *m* $CaCl_2$ (unrounded)

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

13.157 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.

13.158 a) Molarity of $N_2$ = (1.00 atm) (78% $N_2$ / 100%) (7.0 x $10^{-4}$ mol/L•atm) = 5.46 x $10^{-4}$ = **5.5 x $10^{-4}$ *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.

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

= 1.47535 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 atm) (78% $N_2$ / 100%) (7.0 x $10^{-4}$ mol/L•atm)

= 1.35154 x $10^{-3}$ = **1.4 x $10^{-3}$ *M* $N_2$**

c) Moles of $N_2$ per liter at the surface = 5.56 x $10^{-4}$ mol $N_2$. Moles of $N_2$ per liter at 50. ft = 1.35154 x $10^{-3}$ mol $N_2$.

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

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

= 19.4635 = **19 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) Moles $N_2$ dissolved = (1.00 x $10^4$ L) (1.20 atm) (7.0 x $10^{-4}$ mol/L•atm) = 8.4 mol $N_2$

$$PV = nRT \text{ so } V = nRT / P = \frac{(8.4\ mol)\left(0.0821\frac{L \bullet atm}{mol \bullet K}\right)((273 + 25)K)}{(1.20\ atm)} = 171.26 = \mathbf{1.7\ x\ 10^2\ L\ N_2}$$

b) Moles $CO_2$ dissolved = (1.00 x $10^4$ L) (1.20 atm) (2.3 x $10^{-2}$ mol/L•atm) = 276 mol $CO_2$ (unrounded)

$$V = nRT / P = \frac{(276\ mol)\left(0.0821\frac{L \bullet atm}{mol \bullet K}\right)((273 + 25)K)}{(1.20\ atm)} = 5627.1 = \mathbf{5.6\ x\ 10^3\ L\ CO_2}$$

c) Carbon dioxide reacts with water to form carbonic acid. The reaction allows more carbon dioxide to dissolve than the unreactive nitrogen.

13.161 a) $\text{Concentration} = \left(\frac{58\ mL}{7.0\ L}\right)\left(\frac{40\%}{100\%}\right)\left(\frac{0.789\ g}{mL}\right)\left(\frac{22\%}{100\%}\right)\left(\frac{10^{-3}\ L}{1\ mL}\right)$ = 5.7529 x $10^{-4}$ = **5.8 x $10^{-4}$ g/mL**

b) $\left(\frac{58\ mL}{5.7529\ x\ 10^{-4}\ g/mL}\right)(0.0030\ g/mL)$ = 302.456 = **3.0 x $10^2$ mL**

13.162 a) 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 = **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:

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 x $10^{-6}$ = **4 x $10^{-6}$ 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 = **1 L $CO_2$**