11
Solutions

I. Concentration Units
A. Mass percent
1. Formula for determining mass percent:

2. *Exercises*
   a. Calculate the mass percent of water in a solution made up of 123 g of NaOH and 289 g of water.
   Solution: Since

$$\text{mass percent of water} = \frac{\text{mass of water}}{\text{mass of solution}} \times 100$$

we get

$$\text{mass percent of water} = \frac{289\,g}{123\,g + 289\,g} \times 100 = 70.1\%$$

   b. An aqueous solution of potassium iodide is often made up by mixing KI and I_2 and dissolving the mixture in water. A 5.06% KI solution has 18 g of KI and I_2 and 35 g of water. How much KI is in the mixture? **(E1)**

B. Parts per million (ppm)
   1. Formula for determining parts per million:

   2. Mathematical relation between parts per million and mass percent:

   3. Exercises
      a. Methyl mercury(II) chloride, CH_3HgCl, is a toxic pollutant usually found in waters around industrial sites. If a sample of 250.0 mL of water (d = 1.00 g/mL) is found to contain 0.200 mg of CH_3HgCl, what is the concentration of methyl mercury(II) chloride in parts per million?
      Solution: The formula we use is

$$ppm = \frac{\text{mass of solute}}{\text{total mass of solution}} \times 10^6$$

      The mass of the solute, CH_3HgCl, is 0.200 mg, which is 2.00×10^{-4} g. The mass of the solution should also be in grams. Since we are given a volume and density, we obtain

$$250.0 \text{ mL of solution} \times \frac{1.00 \text{ g solution}}{1.00 \text{ mL solution}} = 250.0 \text{ g solution}$$

      Substituting now into the formula for ppm, we get

$$ppm = \frac{2.00 \times 10^{-4} \text{ g}}{250.0 \text{ g}} \times 10^6 = 0.800$$

      b. A water solution contains 0.56 ppm of lead as a pollutant. If you drank a cup (1/2 pint) of this water, how many mg of lead would you ingest? (density of water = 1.00 g/mL) **(E2)**

C. Mole fraction

Recall from Chapter 5 that the mole fraction of a component in a mixture is given by the relation

$$X_A = \frac{\text{number of moles of A}}{\text{total number of moles of all components}}$$

D. Molality

1. Symbol for molality: _____

Do not confuse the symbol for molality (m) with the symbol for mass (m) that we used earlier. The symbol for molality is always in italics.

2. Formula for determining molality:

3. *Exercises*

a. A solution contains sucrose, $C_{12}H_{22}O_{11}$, in water. The solution has a total mass of 738 g and the mass percent of sucrose in the solution is 38.4%. Calculate the molality of the solution.

Solution: To calculate the molality of the solution, you need to know two things: the mass of the solute and the mass of the solvent. You can calculate the mass of the solute, sucrose, by using the information on the mass percent of the solution.

$$38.4 = \frac{\text{mass of sucrose}}{738 \text{ g solution}} \times 100$$

Solving for the mass of sucrose, we get

$$m_{\text{sucrose}} = \frac{38.4 \times 738}{100} = 283 \text{ g}$$

Since the solution consists of water and sucrose, the mass of water is

$$738 \text{ g} - 283 \text{ g} = 455 \text{ g}$$

Since molality is moles of solute/kilogram of solvent, we need to convert grams of solute to moles of solute, and grams of solvent to kilograms of solvent. Starting with the solute, we need the molar mass of sucrose if we are to calculate moles from mass. The molar mass of sucrose is

$$12(12.01) + 22(1.01) + 11(16.0) = 342 \text{ g}$$

Using this molar mass, we now convert to moles.

$$283 \text{ g sucrose} \times \frac{1 \text{ mol sucrose}}{342 \text{ g sucrose}} = 0.827 \text{ mol}$$

The unit for the solvent has to be changed from grams to kilograms.

$$455 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.455 \text{ kg}$$

The molality is therefore

$$m = \frac{0.827 \text{ mol solute}}{0.455 \text{ kg}} = 1.82$$

b. A 0.633 *m* solution of vitamin B_6 (pyridoxine) in water is prepared by adding 53.5 g of vitamin B_6 to 5.00×10^2 g of water. Calculate the molar mass of vitamin B_6. **(E3)**

E. Molarity

    1. Symbol for molarity: _____

    2. Formula for molarity:

    3. What a given molarity can tell you
        a. The number of moles of solute in a liter of solution
            Example: A 0.250 M solution of Na_2SO_4 has 0.250 mol of Na_2SO_4 in one liter of solution. Remember solution is not the same as solvent. A solution that is 0.250 M does *not* mean 0.250 mol of Na_2SO_4 in one liter of water.

b. The moles of solute in a given volume of solution
You can use the molarity of a solution as a conversion factor, with which you can convert a given volume of solution to the number of moles of solute in that solution.
Example: How many moles of Na_2SO_4 are there in 0.300 L of solution that is 0.250 M?

$$0.250\,M \quad \text{means} \quad \frac{0.250\,\text{mol}\ Na_2SO_4}{1\,\text{L solution}}$$

Thus,

$$0.300\,L \times \frac{0.250\,\text{mol}\ Na_2SO_4}{1\,L} = 0.0750\,\text{mol}$$

c. The volume of solution that is necessary, given the number of moles of solute and the molarity of the solution
Example: What is the volume of a 0.250 M solution that has 0.700 mol of Na_2SO_4 as solute?
Again, we use the conversion factor

$$\frac{0.250\,\text{mol}\ Na_2SO_4}{1\,\text{L solution}}$$

to convert 0.700 mol of solute to a volume of solution. Thus,

$$0.700\,\text{mol} \times \frac{1\,\text{L solution}}{0.250\,\text{mol}\ Na_2SO_4} = 2.80\,L$$

d. The number of grams of solute in a liter of solution
Example: How many grams of Na_2SO_4 are in a liter of a solution that is 0.250 M?
Here you have to use the conversion factor

$$\frac{142\,g\ Na_2SO_4}{1\,\text{mol}\ Na_2SO_4}$$

Thus, since 0.250 M means 0.250 mol Na_2SO_4/L of solution, we get

$$\frac{0.250\,\text{mol}\ Na_2SO_4}{\text{L solution}} \times \frac{142\,g\ Na_2SO_4}{1\,\text{mol}\ Na_2SO_4} = \frac{35.5\,g\ Na_2SO_4}{\text{L solution}}$$

e. The number of grams of solute in a given volume of solution
Example: How many grams of Na_2SO_4 are in 0.800 L of a 0.250 M solution?
Again the conversion factors are:

$$\frac{142\,g\ Na_2SO_4}{1\,\text{mol}\ Na_2SO_4} \quad \text{and} \quad \frac{0.250\,\text{mol}\ Na_2SO_4}{\text{L solution}}$$

Here 0.800 L has to be converted to grams of Na_2SO_4. So

$$0.800\,L \times \frac{0.250\,\text{mol}\ Na_2SO_4}{\text{L solution}} \times \frac{142\,g\ Na_2SO_4}{1\,\text{mol}\ Na_2SO_4} = 28.4\,g\ Na_2SO_4$$

f. The volume of a given molarity required for a given mass of solute
   Example: What is the volume of a solution that has 157 g of Na_2SO_4 if the solution is 0.250 M?
   The conversion factors again are:

$$\frac{142\,g\,Na_2SO_4}{1\,mol\,Na_2SO_4} \quad \text{and} \quad \frac{0.250\,mol\,Na_2SO_4}{L\,solution}$$

This time we have to change 157 g Na_2SO_4 to liters of solution. Thus,

$$157\,g\,Na_2SO_4 \times \frac{1\,mol\,Na_2SO_4}{142\,g\,Na_2SO_4} \times \frac{L\,solution}{0.250\,mol\,Na_2SO_4} = 4.42\,L\,solution$$

4. *Exercises*
   a. Calculate the molarity of a solution prepared by adding enough water to 73.2 g of ammonium nitrate to make 0.835 L of solution. **(E4)**

   b. How many moles of calcium chloride do you need in 356 mL of solution to make a 0.125 M solution? **(E5)**

   c. How many grams of potassium permanganate are required in 565 mL of solution to obtain a 0.380 M solution? **(E6)**

   d. What volume of solution is required for a 1.25 M solution of $Pb(NO_3)_2$ that contains 58.5 g of solute? **(E7)**

   e. How many moles of phosphate and ammonium ions are there in 0.250 L of a 0.150 M solution of ammonium phosphate? **(E8)**

Caution: It can be very confusing to do molarity problems using ratios and proportions. The factor conversion method is simpler and clearer.

6.

5. Preparing solutions by dilution
   a. Formula used for dilution problems:

   b. *Exercises*
      (1) How would you prepare 100.0 mL of a 0.200 M solution of NaOH from
      a bottle that is labeled 0.4871 M?
      Solution: The concentrated solution is the one that is on hand. In
      this case it is the one labeled 0.4871 M. The dilute solution is the one
      that has to be prepared. In this case it is 0.200 M. What you want to
      find is the volume of concentrated solution to take so that you can add
      enough water to make 100.0 mL of a 0.200 M solution. Thus, V_c is
      unknown, M_c is 0.4871 M, V_d is 100.0 mL, and M_d is 0.200 M. Using
      the formula, we get

$$V_c = \frac{M_d \times V_d}{M_c} = \frac{0.200\,M \times 100.0\,mL}{0.4871\,M} = 41.1\,mL$$

      You would prepare the final solution by taking 41.1 mL of the initial
      solution and adding enough water to make the final volume 100.0 mL.
      Note that in dilution problems, the volumes of the solutions do not have
      to be in liters. You just have to be consistent. In calculating molarity for
      other problems, however, you must always have the volume in liters.
      (2) Calculate the molarity of a solution prepared by taking 25.00 mL of a
      0.400 M solution of HNO_3 and adding enough water to make 75.00
      mL of solution. **(E9)**

F. Relating concentration units to each other

    1. If you examine all concentration units closely, you see that the numerator of each unit is either in grams of solute or in moles of solute. The denominator in each unit is the one that varies. When you want to convert from one concentration unit to another, it is wise to separate the numerator from the denominator and work with them as separate entities. After you have gotten the required units for both numerator and denominator, put them together.

    2. Given a concentration unit, it is also a good idea to arbitrarily choose one unit of the denominator. For example, if you were asked to calculate the molarity of a 2.50 molal solution, then choose one kilogram as the unit for the solvent since that is the denominator unit of the given molality. The numerator is of course 2.50 moles of solute, which is also the numerator for molarity.

    3. It is also important to know the density in most of these interconversions.

    4. You will find that organizing your data by filling in a prepared table is the easiest way to do these interconversions. You may not need all the entries, but you will find that the data is much more accessible arranged this way. A blank table should look like this:

	mass (g)	moles	volume (L)
solute			X
solvent			X
solution			

    We put X's for the volumes of solute and solvent because we will never need those entries. The use of the table will be shown in the following exercise.

    5. *Exercises*

      a. What is the molarity of an aqueous solution of HCl that is 35.0% by mass and whose density is 1.19 g/mL?

      Solution: Since we are given mass percent (mass of solute/100 g of solution), we arbitrarily choose to work with 100.0 grams of solution. Doing that, we can deduce that there are 35.0 g of solute and 65.0 g of solvent. Your table should now look like this:

	mass (g)	moles	volume (L)
solute	35.0		X
solvent	65.0		X
solution	100.0		

    Knowing that the solute is HCl ($\mathcal{M}$ = 36.45 g/mol) and that the solvent is water, ($\mathcal{M}$ = 18.02 g/mol), we can fill in the table with the results of the following calculations.

$$35.0 \text{ g HCl} \times \frac{1 \text{ mol}}{36.45 \text{ g}} = 0.960 \text{ mol HCl}$$

$$65.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol}}{18.02 \text{ g}} = 3.61 \text{ mol H}_2\text{O}$$

$$\text{mol solution} = 0.960 + 3.61 = 4.57$$

	mass (g)	moles	volume (L)
solute	35.0	0.960	X
solvent	65.0	3.61	X
solution	100.0	4.57	

	mass (g)	moles	volume (L)
solute	35.0	0.960	X
solvent	65.0	3.61	X
solution	100.0	4.57	0.0840

Note that we did all this without worrying about what was asked for. Now, no matter what concentration unit is asked for, you can easily calculate it, provided that you know the definition for that concentration unit. In this case, the unit asked for is molarity. Thus using the definition for molarity and picking out the appropriate data from the table, we get

$$M = \frac{mol\ solute}{L\ solution} = \frac{0.960\ mol\ HCl}{0.0840\ L\ solution} = 11.4$$

We could, however, just as easily have solved for molality or mole fraction of HCl.

b. How would you prepare 5.00 L of 2.00 M H_2SO_4 from a bottle labeled: H_2SO_4 — 98.0% by mass; density — 1.84 g/mL? **(E10)**

c. Calculate the molality of a 2.500 M solution of sulfuric acid. Its density is 1.200 g/mL.

   Solution: We choose 1.000 L of solution, which therefore means that we have 2.500 moles of solute (from the definition of molarity). We list these as our initial entries in the table.

	mass (g)	moles	volume (L)
solute		2.500	X
solvent			X
solution			1.000

There are many ways in which you can now proceed. We choose to determine the mass of the solution by using the density of the solution that is given.

$$1000 \text{ mL solution} \times \frac{1.200 \text{ g solution}}{1 \text{ mL solution}} = 1.200 \times 10^3 \text{ g solution}$$

Our table now looks like this:

	mass (g)	moles	volume (L)
solute		2.500	X
solvent			X
solution	1.200×10^3		1.000

We now determine the mass of the solute using the molar mass of H_2SO_4.

$$2.500 \text{ mol } H_2SO_4 \times \frac{98.08 \text{ g}}{1 \text{ mol}} = 245.2 \text{ g } H_2SO_4$$

We enter that value in our table.

	mass (g)	moles	volume (L)
solute	245.2	2.500	X
solvent			X
solution	1.200×10^3		1.000

Now, we can determine the mass of solvent ($1.200 \times 10^3 - 245.2$) and the number of moles of solvent, water. Finally, we determine the moles of solution by adding moles of solute and moles of solvent. The completed table now looks like this.

	mass (g)	moles	volume (L)
solute	245.2	2.500	X
solvent	955	53.0	X
solution	1.200×10^3	55.5	1.000

Since the problem asks for the molality of the solution, we pick the moles of solute from the table (2.500 mol) and the mass of solvent in *kg* (0.955). Thus

$$m = \frac{2.500 \text{ mol solute}}{0.955 \text{ kg solvent}} = 2.62$$

d. Calculate the mol fraction of H_2SO_4, its mass percent, and parts per million in the solution in *c.* **(E11)**

II. Principles of Solubility
A. Solute-solvent interactions
1. Solubility is dependent on both the nature of the solute and the solvent.
2. Electrical forces that determine the extent to which ionic compounds dissolve in water are

    a. _____

    _____

    b. _____

    _____

3. "Like dissolves like" principle
    a. Ionic substances tend to dissolve in _____ solvents.
    b. Nonpolar or slightly polar solutes are most soluble in solvents that are

    (polarity) _____

B. Temperature effects on solubility

    1. Solid solute in liquid solvent

    Ordinarily, dissolving a solid is an endothermic process. Thus, an increase in temperature most often favors the solubility of a solid solute in a liquid solvent.

    2. Saturation

        a. Unsaturated solution

        One that has less solute than it is capable of dissolving at a particular temperature.

        b. Saturated solution

        One that has the exact amount of solute that it is capable of dissolving at a particular temperature.

        c. Supersaturated solution

        Definition: _____

        _____

    3. Gas solute in liquid solvent

    Ordinarily, dissolving a gas in a liquid is an exothermic process. Therefore, a decrease in temperature favors the solubility of the gas in a liquid.

C. Pressure effects on solubility

    1. Solid solute in a liquid solvent

    No pressure effects noticeable.

    2. Gas solute in liquid solvent

    An increase in pressure increases the solubility of a gas in a liquid.

    3. Henry's Law

    Mathematical statement:

III. Colligative Properties of Solutions

    A. Definition of colligative properties:

    _____

    _____

    B. Colligative properties of nonelectrolytes

        1. Vapor pressure lowering

        a. Raoult's Law

        Mathematical expression:

b. *Exercises*

(1) The vapor pressure of pure ethanol, C_2H_5OH, at 20°C is 39.98 mm Hg. What is the vapor pressure of ethanol over a solution containing 5.0% by mass of iodine?

Solution: To be able to use Raoult's Law, which says that the vapor pressure of the solvent over the solution is equal to the mole fraction of the solvent multiplied by the vapor pressure of the pure solvent, we must express the concentration of the solvent (ethanol) in terms of mole fraction.

Assume 100 g of solution (since mass percent is given). The solution has 5.0% iodine and 95.0% ethanol. In grams, it has 5.0 g of iodine and 95.0 g of ethanol. The mole fraction of ethanol (X_1) is then

$$X_1 = \frac{\text{mol ethanol}}{\text{mol ethanol} + \text{mol I}_2} = \frac{\dfrac{95.0}{46.1}}{\dfrac{95.0}{46.1} + \dfrac{5.0}{253.8}} = 0.990$$

We now plug into Raoult's Law:

$$P_1 = 0.990 \times 39.98 = 39.6 \, \text{mm Hg}$$

(2) At 64°C the vapor pressure of pure benzene is 400.0 mm Hg. Calculate the vapor pressure of benzene above a solution prepared by mixing 25.0 g of benzene (C_6H_6) and 15.0 g of naphthalene ($C_{10}H_8$). What is the vapor pressure lowering? **(E12)**

2. Osmotic pressure
    a. Osmosis — definition: _____

    _____

    _____

    b. Symbol used for osmotic pressure: _____
    c. Mathematical relationship between osmotic pressure and molarity:

3. Freezing point depression and boiling point elevation
    a. Mathematical expression for the determination of the freezing point depression of a solution:

    b. Mathematical expression for the determination of the boiling point elevation of a solution:

    c. *Exercises*
      (1) The freezing point of benzene is 5.50°C. Its freezing point constant is 5.12°C/m. How much p-dichlorobenzene ($C_6H_4Cl_2$) must be added to 50.0 g of benzene (C_6H_6) to lower the freezing point of the solution to that of pure water?
      Solution: The formula you need is

$$\Delta T_f = k_f \times m$$

ΔT_f means freezing point depression, which is the difference between the freezing point of the pure solvent (5.50°C) and that of the solution (0.00°C).

$$\Delta T_f = 5.50°C - 0.00°C = 5.50°C$$

k_f is given: 5.12°C/m. Hence, we must calculate the molality of the solution before we can determine the mass of p-dichlorobenzene (PCB).

$$m = \frac{\Delta T_f}{k_f} = \frac{5.50}{5.12} = 1.07$$

1.07 m means 1.07 mol PCB/kg benzene. We now obtain the mass of PCB in 50.0 g (0.0500 kg) benzene.

$$0.050 \text{ kg } C_6H_6 \times \frac{1.07 \text{ mol PCB}}{1 \text{ kg } C_6H_6} \times \frac{147 \text{ g PCB}}{1 \text{ mol PCB}} = 7.86 \text{ g PCB}$$

(2) Methylene chloride (CH_2Cl_2) has a boiling point of 40.7°C and a boiling point constant of 2.49 °C/m. What is the boiling point of a 20.0 % (by mass) solution of ferrocene, ($Fe(C_5H_5)_2$), in methylene chloride? **(E13)**

4. Molar masses from colligative properties

Colligative properties are most often used to determine the molar masses of unknown solutes.

Exercises

a. Calculate the molar mass of an unknown substance, if when 6.32 g of the unknown is added to 50.0 g of camphor, the solution formed has a boiling point of 212.00°C.

Solution: You need to use the formula

$$\Delta T_b = k_b \times m$$

According to Table 10.2 of your text, k_b for camphor is 5.61°C/m. Since the boiling point of pure camphor is 207.42°C, the boiling point elevation is 4.58°C. We can therefore calculate the molality of the solution.

$$m = \frac{\Delta T_b}{k_b} = \frac{4.58}{5.61} = 0.816$$

Using the definition for molality we get

$$0.816\,m = \frac{\text{mols of unknown solute}}{0.0500\,\text{kg camphor}}$$

and

$$\text{mols of unknown solute} = 0.816 \times 0.0500 = 0.0408\,\text{mol}$$

We now find the molar mass by substituting into the definition for molar mass

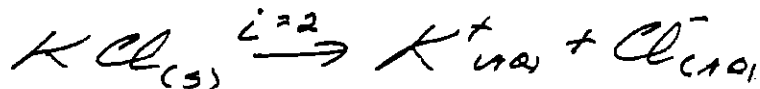$$\text{molar mass} = \frac{6.32\,\text{g}}{0.0408\,\text{mol}} = 155\,\frac{\text{g}}{\text{mol}}$$

b. A 0.0344-g sample of starch in 10.0 mL of water at 25°C has an osmotic pressure of 4.0 mm Hg. What is the approximate molar mass of the starch? **(E14)**

C. Colligative properties of electrolytes
1. Colligative properties of dilute solutions are directly proportional to the concentration of solute particles. Because electrolytes in solution split into ions, there are more particles in solution in an electrolyte than there are in solutions of nonelectrolytes. For this reason, electrolytes have a greater effect on colligative properties than do nonelectrolytes.
2. The equations for determining the colligative properties of electrolytes are the same as those for nonelectrolytes. The only difference is that in determining the freezing point depression, the boiling elevation, and the osmotic pressure for electrolytes, a multiplier i is added to the right hand side of the equation. The multiplier i stands for the number of particles in solution. For a nonelectrolyte, the multiplier is always one. For an electrolyte like NaCl, i = 2 because NaCl splits up into two ions. In determining the vapor pressure lowering for an electrolyte in solution, the difference is a bit more subtle. You will have to consider the number of moles of each ion in determining the mole fraction of the solute.

$$\Delta t_f = (k_f \, m) \, i \left. \begin{array}{c} \\ \\ \end{array} \right\} \quad i = \text{VAN'T HOFF FACTOR}$$
$$\Delta t_b = (k_b \, m) \, i$$

$Ex:$

$$KCl_{(s)} \xrightarrow{i=2} K^+_{(aq)} + Cl^-_{(aq)}$$

$$CaCl_{2(s)} \xrightarrow{i=3} Ca^{+2}_{(aq)} + 2\,Cl^-_{(aq)}$$

$$AlCl_{3(s)} \xrightarrow{i=4} Al^{+3}_{(aq)} + 3\,Cl^-_{(aq)}$$

Henry's Law Problem

A bottle of soda is packaged so that a bottle at $25°C$ contains CO_2 gas at a pressure of 5.0 atm over the liquid. Assuming the partial pressure of CO_2 in the atmosphere is 4.0×10^{-4} atm, calculate the equilibrium concentrations of CO_2 in the soda <u>both</u> before and after opening the bottle. The Henry's Law constant for CO_2 in aqueous solution is $32 \frac{L \cdot atm}{mol}$ or $32 \frac{atm}{M}$ at $25°C$.

Osmotic Pressure Problem

How would you prepare 1.0 L of an aqueous solution of sucrose $(C_{12}H_{22}O_{11})$ having an osmotic pressure of 15 atm at a temperature of $22°C$?

Colligative Properties Worksheet

1) If I add 45 grams of sodium chloride to 500 grams of water, what will the melting and boiling points be of the resulting solution? $K_b(H_2O) = 0.52$ $^0C/m$ and $K_f(H_2O) = 1.86$ $^0C/m$.

2) What is the vapor pressure of the solution in problem #1 at 25^0 C? The vapor pressure of pure water at 25^0C is 3.17 kPa.

3) Which solution will have a higher boiling point: A solution containing 105 grams of sucrose ($C_{12}H_{22}O_{11}$) in 500 grams of water or a solution containing 35 grams of sodium chloride in 500 grams of water?

18

$$\text{Mass \%} = \frac{\text{mass part}}{\text{mass whole sol'n}} \times 100$$

$$\text{ppm} = (\text{mass \%})(1 \times 10^4)$$

$$m \text{ (molality)} = \frac{\text{moles solute}}{kg \text{ solvent}}$$

$$M \text{ (molarity)} = \frac{\text{moles solute}}{L \text{ of solution}}$$

Mole Fraction (X)

$$X_A = \frac{\text{moles A}}{\text{moles of sol'n}}$$

Dilutions: $M_c V_c = M_0 V_0$

Henry's Law: $C = KP \leftarrow$ Pressure   or $P = kC$ (depending on units for k)

$\uparrow$ concentration $\nwarrow$ constant

$\underset{\text{Pressure}}{\uparrow}$ $\underset{\text{constant}}{\uparrow}$ $\underset{\text{concentration}}{\nwarrow}$

Raoult's Law: $P_{\substack{\text{solvent in} \\ \text{sol'n}}} = (X_{\text{solv}})(P^0_{\substack{\text{pure} \\ \text{solvent}}})$

Osmotic Pressure $(\pi) = MRT$ M = Molarity, T = Temp, $R = .0821 \frac{L \cdot atm}{mol \cdot K}$

Freezing Pt Depression $(\Delta t_f) = k_f m$ k_f = freezing pt constant, m = molality

Boiling Pt Elevation $(\Delta t_b) = k_b m$ k_b = boiling pt elevation, m = molality

Van't Hoff
Eq.

$\left. \begin{array}{l} \Delta t_f = (k_f m) i \\ \Delta t_B = (k_B m) i \end{array} \right\}$ i = Van't Hoff Factor
(the # of electrolytes)