<u>Key Worksheet 5: Colligative Properties</u>

Objectives: To understand and be able to apply the following colligative properties.

<u>Freezing Point Depression</u>: When a solute is added to a solvent, the freezing point (same as the melting point) of the solvent is lowered. We can calculate this using the equation:

$$\Delta T_F = imK_F$$

Here ΔT_F is the change in freezing point, $T_{final} - T_{initial}$, i is the van't Hoff factor for the solute, m is the molality of the solute in the solution, and K_F is the freezing point depression constant for that solvent.

<u>Boiling Point Elevation</u>: When a solute is added to a solvent, the boiling point of the solvent is elevated. We can calculate this using the equation:

$$\Delta T_B = imK_B$$

Here ΔT_B is the change in boiling point, $T_{final} - T_{initial}$, i is the van't Hoff factor for the solute, m is the molality of the solute in the solution, and K_B is the boiling point depression constant for that solvent.

<u>Osmotic Pressure</u>: When a solution is in contact through a semipermeable membrane with some of the pure solvent, the solvent will travel through that semipermeable membrane in an attempt to dilute the solution. It will exert a certain pressure, called the osmotic pressure, when doing this. We can calculate the osmotic pressure of a solution using the following equation:

$$\Pi = iMRT$$

Here Π is the osmotic pressure, we will use units of atmospheres, i is the van't Hoff factor for the solute, M is the molarity of the solution, R is the gas constant (we will use 0.08206 L·atm/K·mol), and T is the temperature of the solution in Kelvins.

1.) 39.2 g of urea, $(NH_2)_2CO$, is dissolved in 142.2 g of water. Urea is a nonelectrolyte.
Given that for water K_F = −1.86 $°C/m$ and K_B = 0.512 $°C/m$, calculate the boiling point and
the freezing point of the resulting solution.

$$\frac{39.2 \text{ g } (NH_2)_2CO}{60.056 \text{ g/mol}} = 0.652_7 \text{ mol } (NH_2)_2CO \qquad \frac{0.652_7 \text{ mol } (NH_2)_2CO}{0.1422 \text{ kg } H_2O} = 4.59_0 \text{ } m$$

$$\Delta T_F = -(4.59_0 \text{ } m) \cdot 1.86 \frac{°C}{m} = -8.53_7 °C \qquad T_F = -8.53_7 °C + 0.00 °C = -8.53_7 °C$$

$$\Delta T_B = (4.59_0 \text{ } m) \cdot 0.512 \frac{°C}{m} = 2.35_0 °C \qquad T_B = 2.35_0 °C + 100.00 °C = 102.35_0 °C$$

B.P. = <u>102.35 ºC</u>

F.P. = <u>−8.54 ºC</u>

2.) Ethanol (CH_3CH_2OH) is hygroscopic (it absorbs water from the air). The normal
boiling point of ethanol is 78.37 $°C$. If a 25.00 gram sample of ethanol that was left
standing in the air boils at 79.89 $°C$, how many grams of water did it absorb?
K_B=1.22 $°C/m$ for ethanol.

$$\Delta T_B = 79.89 °C - 78.37 °C = +1.52 °C$$

$$\Delta T_B = im K_B \Rightarrow m = \frac{\Delta T_B}{i K_B} = \frac{1.52 °C}{(1)(1.22 \frac{°C}{m})} = 1.24_5 \text{ } m$$

$$\frac{1.24_5 \text{ mol } H_2O}{1 \text{ kg } CH_3CH_2OH} \times 0.02500 \text{ kg } CH_3CH_2OH = 0.0311_4 \text{ mol } H_2O$$

$$0.00311_4 \text{ mol } H_2O \times \frac{18.015 \text{ g } H_2O}{1 \text{ mol } H_2O} = 0.561_1 \text{ g } H_2O$$

Mass of H_2O = <u>0.561 g</u>

3.) A solution contains 4.39 g of a nonvolatile hydrocarbon as a solute in 101.2 g of
acetone. The resulting solution has a boiling point of 56.69 $°C$. Given that the normal
boiling point of pure acetone is 55.95 $°C$ and K_B = 1.71 $°C$ for acetone, calculate the molar
mass of the solute.

$$\Delta T_B = 56.69 °C - 55.95 °C = +0.74 °C$$

$$\Delta T_B = im K_B \Rightarrow m = \frac{\Delta T_B}{i K_B} = \frac{0.74 °C}{(1)(1.71 \frac{°C}{m})} = 0.43_2 \text{ } m$$

$$\frac{0.43_2 \text{ mol hydrocarbon}}{1 \text{ kg acetone}} \times 0.1012 \text{ kg acetone} = 0.043_7 \text{ mol unknown}$$

$$\frac{4.39 \text{ g unknown}}{0.043_7 \text{ mol unknown}} = 1.0_0 \times 10^2 \frac{\text{g}}{\text{mol}}$$

Molar Mass = <u>1.0x10² g/mol</u>

4.) Calculate the osmotic pressure of a solution prepared by dissolving 52.34 g of $MgCl_2$ in 545.8 mL of water at 23.12 °C. The van't Hoff factor for magnesium chloride is $i = 2.7$.

$$\frac{52.34 \text{ g } MgCl_2}{95.211 \text{ g/mol}} = 0.5497_2 \text{ mol } MgCl_2$$

$$\frac{0.5479_2 \text{ mol } MgCl_2}{0.5458 \text{ L solution}} = 1.007_1 \text{ M}$$

$$23.12 \text{ °C} + 273.15 = 296.27 \text{ K}$$

$$\Pi = i\text{MRT} = (2.7) \cdot \left(1.007_1 \frac{\text{mol}}{\text{L}}\right) \cdot \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}\right) \cdot (296.27 \text{ K}) = 66._1 \text{ atm}$$

$$\Pi = \underline{66 \text{ atm}}$$

5.) A 0.03234 g sample of a nonelectrolyte is dissolved in enough water to make 125.0 mL of solution. The osmotic pressure of the resulting solution is 80.79 mmHg at 27.0 °C. Calculate the molar mass of the nonelectrolyte solute.

$$\Pi = 80.79 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.1063_0 \text{ atm} \quad T = 27.0 \text{ °C} + 273.15 = 300.0_5 \text{ K}$$

$$\Pi = i\text{MRT} \Rightarrow \text{M} = \frac{\Pi}{i\text{RT}} = \frac{0.1063_0 \text{ atm}}{(1) \cdot \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}\right)(300.0_5 \text{ K})} = 0.004317_3 \frac{\text{mol}}{\text{L}}$$

$$\left(0.004317_3 \frac{\text{mol unknown}}{\text{L solution}}\right) \cdot (0.1250 \text{ L solution}) = 5.396_7 \times 10^{-4} \text{ mol unknown}$$

$$\frac{0.03234 \text{ g unknown}}{5.396_7 \times 10^{-4} \text{ mol unknown}} = 59.92_5 \frac{\text{g}}{\text{mol}}$$

$$\text{M.M.} = \underline{59.93 \text{ g/mol}}$$