FREEZING POINT DEPRESSION

$$\Delta T_f = K_f \times c_m$$

concentration of solute (molality)

Freezing point depression constant (for SOLVENT)

Freezing point depression: The amount the freezing temperature is LOWERED by the solute.

- Applications: In chemistry, this effect is often used to determine the molecular weight of an unknown molecule.

A solution of 2.500g of unknown dissolved in 100.0 g of benzene has a freezing point of 4.880 C. What is the molecular weight of the unknown?

$K_{f,\text{benzene}} = 5.065\ {}^{\circ}\text{C}/m$, $T_{f,\text{benzene}} = 5.455^{\circ}\text{C}$ (see p500)

$$\underbrace{\Delta T_f}_{5.455^{\circ}\text{C} - 4.880^{\circ}\text{C} = 0.575^{\circ}\text{C}} = \underbrace{K_f}_{5.065^{\circ}\text{C}/m} \times \boxed{C_m} = \frac{\text{mol unknown}}{\text{kg benzene}}$$

To find molecular weight, we will find the moles of unknown. To find moles of unknown, we need to find the molal concentration Cm!

We calculate MOLAL CONCENTRATION (Cm) from the freezing point depression

$$0.575^{\circ}\text{C} = 5.065\ {}^{\circ}\text{C}/m \times C_m$$

$$C_m = 0.113524185\ m = 0.113524185\ \frac{\text{mol unknown}}{\text{kg benzene}}$$

To find moles unknown, multiply the mass of benzene (IN KG) by the molality:

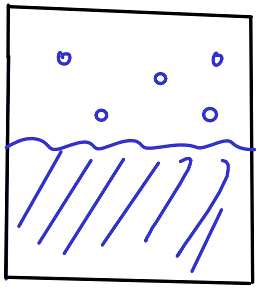
$$0.1000\ \text{kg benzene} \times 0.113524185\ \frac{\text{mol unknown}}{\text{kg benzene}} = 0.0113524186\ \text{mol unk}$$

MOLECULAR WEIGHT is mass per mole, so

$$MW = \frac{\text{mass unk}}{\text{mol unk}} = \frac{2.500\text{g}}{0.0113524186\ \text{mol unk}} = \boxed{220.\ \text{g/mol}}$$

VAPOR PRESSURE LOWERING

- Described by RAOULT'S LAW



P_A = partial pressure of the VAPOR of solvent molecules.

$$\underline{P_A} = \underline{P_A^*} \times \underline{X_A}$$

mole fraction of component A

vapor pressure of pure component A (depends on temperature)

partial pressure of component A in a solution

... but component "A" above is actually the SOLVENT. If we want to describe this as a colligative property, we want to express Raoult's law in terms of the SOLUTE! Assuming a two-component mixture, we get...

$$\underline{\Delta P} = P_A^* \times \underline{X_B}$$

mole fraction of component B (the SOLUTE in a two-component mixture)

Vapor pressure lowering. This is the DECREASE in the vapor pressure of the solvent due to the presence of solute.

BOILING POINT ELEVATION

- Since the vapor pressure is lowered by the presence of a solute, AND since boiling occurs when the vapor pressure of a liquid equals the external pressure - solutes also cause BOILING POINT ELEVATION.

- The equation for boiling point elevation looks almost exactly like the equation for the freezing point depression, and is used in almost the same way.

$$\underline{\Delta T_b} = \underline{K_b} \times \underline{C_m}$$

concentration of solute (molality)

Boiling point elevation constant (for SOLVENT) (p500)

Boiling point elevation: The amount the boiling temperature is RAISED by the solute.

What is the boiling point of a solution that contains 2.817 g of molecular sulfur S_8) dissolved in 100.0 grams of acetic acid?

$T_b = 118.5°C$ $\quad K_b = 3.08\,°C/m$ (see p500 for data)

$$\Delta T_b = \underbrace{K_b}_{3.08°C/m} \times C_m$$

$$C_m = \frac{\text{mol } S_8}{\text{Kg } HC_2H_3O_2} \xrightarrow{} 0.1000 \text{ kg } HC_2H_3O_2$$

We need to convert mass sulfur to moles. Then, we can find Cm and calculate the boiling point elevation.

S_8 : 8×32.07] $256.56 \text{ g } S_8 = \text{mol } S_8$

$$2.817 \text{ g } S_8 \times \frac{\text{mol } S_8}{256.56 \text{ g } S_8} = 0.0109798877 \text{ mol } S_8$$

FInd Cm:

$$C_m = \frac{0.0109798877 \text{ mol } S_8}{0.1000 \text{ kg } HC_2H_3O_2} = 0.1097988775 \ \underline{m}\ S_8$$

Find boiling point ELEVATION:

$$\Delta T_b = 3.08°C/m \times 0.1097988775 \ \underline{m}\ S_8 = 0.3382°C$$

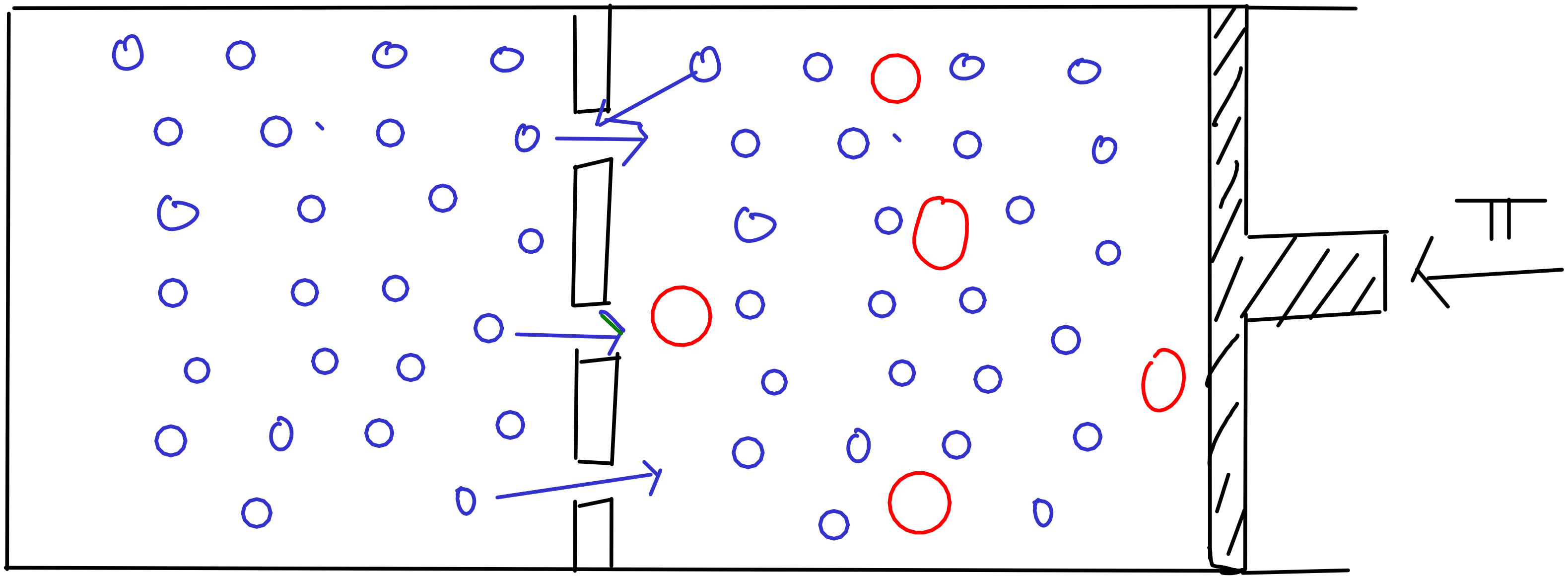
Find the boiling point by ADDING the change to the original solvent boiling point:

$$T_{b,\text{solution}} = 118.5°C + 0.3382°C = \boxed{118.8°C}$$

OSMOTIC PRESSURE

permits flow of solvent, but not solute particles

- OSMOSIS: the flow of solvent molecules through a SEMIPERMEABLE membrane to equalize concentration of solute on each side of the membrane.



The rate of solvent migration towards the RIGHT is greater than that towards the LEFT.

If you apply enough pressure to the piston, osmosis will not occur. This pressure is called the OSMOTIC PRESSURE

$$\Pi = M \times R \times T$$

temperature

ideal gas constant

molar concentration of solute

IONIC COMPOUNDS and colligative properties

- Ionic compounds DISSOCIATE in water into their component ions. Each ion formed can act as a solute and influence the colligative properties!

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

2 ions!

... so the concentration of IONS here is TWICE the nominal NaCl concentration.

$$CaCl_2(s) \longrightarrow Ca^{2+}(aq) + 2\,Cl^-(aq)$$

3 ions!

... so the concentration of IONS here is THREE TIMES the nominal calcium chloride concentration.

- Ions interact with each other in solution, so unless an ionic solution is DILUTE, the effective concentrations of ions in solution will be less than expected. A more advanced theory (Debye-Huckel) covers this, but we'll assume that our solutions are dilute enough so that we can use the concentration of the ions in solution to determine the colligative properties!

If you are at an altitude high enough for the boiling point of water to be 95.00 C, what amount of sodium chloride would you need to add to 1.000 kg of water to raise the boiling point to 100.00 C?

$K_b = 0.512\ ^\circ C/m$ NaCl: 58.443 g/mol

$$\underbrace{\Delta T_b}_{100.00^\circ C - 95.00^\circ C = 5.00^\circ C} = \underbrace{K_b}_{0.512^\circ C/m} \times C_m$$

$$C_m = \frac{\text{mol ions}}{\underbrace{\text{kg } H_2O}_{1.000\text{ kg}}}$$

Find Cm ... molal concentration of IONS

$5.00^\circ C = 0.512^\circ C/m \times C_m$; $C_m = 9.765625\ \underline{m}$ ions

Find moles IONS:

$$1.000\text{ kg } H_2O \times \frac{9.765625\text{ mol ions}}{\text{kg } H_2O} = 9.765625\text{ mol ions}$$

$$NaCl(s) \longrightarrow Na^+(aq) + Cl^-(aq)$$

mol NaCl = 2 mol ions

Each mole of sodium chloride dissociates into TWO MOLES OF IONS!

Find moles NaCl, then convert to mass:

$$9.765625\text{ mol ions} \times \frac{\text{mol NaCl}}{2\text{ mol ions}} \times \frac{58.443\text{ g NaCl}}{\text{mol NaCl}} = \boxed{285\text{ g NaCl}}$$

EXTERNAL FACTORS AFFECTING SOLUBILITY

- There are a few external factors that affect the solubility. (By external, we mean other than the chemical identity of the solute and solvent).

① TEMPERATURE

-For gases dissolved in liquids, the solubility DECREASES as the temperature INCREASES

- This is why THERMAL POLLUTION is bad! Hot water holds less oxygen than cooler water.

- For solids dissolved in liquids, solubility USUALLY increases with temperature. This is not true for ALL solid/liquid solutions.

② PRESSURE

- For gases dissolved in liquids, solubility INCREASES when the partial pressure of the solute gas over the solution INCREASES.

- Consider soft drinks. They go flat after opening because the pressure of carbon dioxide over the liquid goes down.

- No significant pressure effects for solid/liquid solutions.

Some sample colligative propoerties problems from the book...

12.97

p521

What is the freezing point of a 41% solution of urea in water?

$(NH_2)_2CO$: urea

$$\Delta T_f = \underbrace{K_f}_{\text{p500, } 1.858\,^{\circ}\text{C}/m} \times C_m$$

$$T_{f,H_2O} = 0.000\,^{\circ}\text{C}$$

$$C_m = \frac{\text{mol urea}}{\text{kg water}}$$

We need to find mol urea and kg water. We know mass percentage, so we will have to start there...

$$41\%\ \text{urea} : \frac{41\,\text{g urea}}{100\,\text{g solution}}$$

... but we need mass WATER for molality!

; $100\,\text{g solution} - 41\,\text{g urea} = 59\,\text{g}\ H_2O$

Find moles urea from mass and formula weight

$(NH_2)_2CO$:
N: 2 x 14.01
H: 4 x 1.008
C: 1 x 12.01
O: 1 x 16.00
60.062 g/mol

$$41\,\text{g urea} \times \frac{\text{mol urea}}{60.062\,\text{g urea}} = 0.6826279511\ \text{mol urea}$$

So, Cm is

$$C_m = \frac{\text{mol urea}}{\text{kg water}} = \frac{0.6826279511\ \text{mol urea}}{0.059\ \text{kg}\ H_2O} = 11.56996527\ \underline{m}\ \text{urea}$$

$$C_m = \frac{\text{mol urea}}{\text{kg water}} = \frac{0,6826279511 \text{ mol urea}}{0.059 \text{ kg } H_2O} = 11.56996527 \ \underline{m} \text{ urea}$$

Now, find freezing point depression:

$$\Delta T_f = K_f \times C_m = 1,858\,^{\circ}C/m \times 11.56996527 \ \underline{m} \text{ urea}$$

$$= 21\,^{\circ}C$$

So, $T_f = 0.000\,^{\circ}C - 21\,^{\circ}C = \boxed{-21\,^{\circ}C}$