An aqueous solution is contains 8.50 grams of ammonium chloride in each 100. grams of solution. The density of the solution is 1.024 g/mL. Find the molality and molarity of the solution.

NH_4Cl : 53.491 g/mol H_2O: 18.016 g/mol

$$\frac{8.50\text{ g } NH_4Cl}{100\text{ g. solution}} \longrightarrow \frac{\text{mol } NH_4Cl \quad (1)}{\text{L solution} \quad (2)}$$

Given information

Definition of molarity

As before, keep the basis of 100 grams of solution. We need to find (1) and (2)

(1) Convert 8.50 grams ammonium chloride to moles. (Done to get molality!)
(2) Convert grams solution to volume using DENSITY. Then convert to L.

(1) 0.158905364 mol NH_4Cl

(2) $100.\text{g solution} \times \frac{\text{mL}}{1.024\text{ g}} = 97.65625\text{ mL} = 0.09765625\text{ L solution}$

$$M = \frac{\text{mol } NH_4Cl}{\text{L solution}} = \frac{0.158905364\text{ mol } NH_4Cl}{0.09765625\text{ L solution}} = \boxed{1.63\ M\ NH_4Cl}$$

COLLIGATIVE PROPERTIES

- properties unique to solutions.

- depend only on the CONCENTRATION of a solution and not the IDENTITY of the solute**

  **ionic solutes: Remember that they dissociate into MULTIPLE IONS!

(1) Freezing point depression

- The freezing temperature of a SOLUTION gets lower as the CONCENTRATION of a solution increases.

(2) Vapor pressure lowering

- The vapor pressure of a solution (pressure of sovent vapor over a liquid surface) goes DOWN as solution concentration goes UP

(3) Boiling point elevation

- The boiling temperature of a solution increases as the concentration of the solution increases.

(4) Osmotic pressure

- The pressure required to PREVENT the process of osmosis

FREEZING POINT DEPRESSION

$$\Delta T_f = K_f \times C_m$$

- C_m — concentration of solute (molality)
- K_f — Freezing point depression constant (for SOLVENT)
- ΔT_f — 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?

$K_{f,benzene} = 5.12\ ^{o}C/m$, $T_{f,benzene} = 5.5\ ^{o}C$

(Values from p619 OpenStax Table 11.2)

$$\Delta T_f = K_{f,benzene} \times C_m$$

$C_m = \frac{\text{mol unknown}}{\text{kg } C_6H_6}$

$K_{f,benzene} = 5.12\ ^{o}C/m$

$\Delta T_f = 5.5\ ^{o}C - 4.880\ ^{o}C = 0.62\ ^{o}C$

Our initial goal is to find Cm. Why? Because it will allow us to find the moles of unknown. (Multiply the kg of benzene used in the experiment by Cm and you get moles unknown!) Then, the formula weight is just mass unknown used / mol unknown.

$0.62\ ^{o}C = (5.12\ ^{o}C/m) \times C_m$; $C_m = 0.12109375\ m = 0.12109375\ \frac{\text{mol unknown}}{\text{kg } C_6H_6}$

We have 100.0 grams of benzene: $100.0\text{ g } C_6H_6 = 0.1000\text{ kg } C_6H_6$

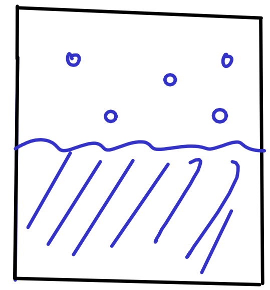
$$0.1000\text{ kg } C_6H_6 \times 0.12109375\ \frac{\text{mol unknown}}{\text{kg } C_6H_6} = 0.012109375\text{ mol unknown}$$

$$MW = \frac{\text{g unknown}}{\text{mol unknown}} = \frac{2.500\text{ g unknown}}{0.012109375\text{ mol unknown}} = 206.4516128\text{ g/mol}$$

$$= \boxed{210\text{ g/mol}}$$

VAPOR PRESSURE LOWERING

- Described by RAOULT'S LAW



P_A = partial pressure of the VAPOR of solvent molecules.

$$P_A = P_A^* \times X_A$$

- X_A: mole fraction of component A
- P_A^*: vapor pressure of pure component A (depends on temperature)
- P_A: 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 Raolt's law in terms of the SOLUTE! Assuming a two-component mixture, we get...

$$\Delta P = P_A^* \times X_B$$

- X_B: mole fraction of component B (the SOLUTE in a two-component mixture)
- ΔP: 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.

$$\Delta T_b = K_b \times c_m$$

(Find constant values on p619 OpenStax Table 11.2)

- c_m: concentration of solute (molality)
- K_b: Boiling point elevation constant (for SOLVENT)
- ΔT_b: 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?

$T_b = 118.1°C$ $K_b = 3.07\ °C/m$ (Values from p619 OpenStax Table 11.2)

$\Delta T_b = K_b \times C_m$

K_b: $3.07°C/m$

$C_m = \frac{mol\ S_8}{Kg\ HC_2H_3O_2}$ ← $100.0\ g = 0.1000\ Kg$

$S_8: 8 \times 32.07 = 256.56\ g/mol$

Calculate moles sulfur, use it to find Cm ...

$2.817\ g\ S_8 \times \frac{mol\ S_8}{256.56\ g\ S_8} = 0.01097988777\ mol\ S_8$

$C_m = \frac{0.01097988777\ mol\ S_8}{0.1000\ Kg\ HC_2H_3O_2} = 0.1097988775\ m\ S_8$

Now find DELTA Tb:

$\Delta T_b = (3.07°C/m)(0.1097988775\ m) = 0.3370825538\ °C$

FInd boiling point by addition:

$118.1°C + 0.3370825538\ °C = 118.4°C$