An aqueous solution is 8.50% ammonium chloride by mass. The density of the solution is 1.024 g/mL
Find: molality and molarity.

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

$$\frac{8.50 \text{ g } NH_4Cl}{100 \text{ g solution}} \longrightarrow \frac{\text{mol } NH_4Cl \text{ (1)}}{\text{Kg } H_2O \text{ (2)}}$$

mass percent (defintion) molality (defintion)

Assume a basis of 100 grams solution, meaning we have 8.5 grams ammonium chloride.
So we need to find (1) moles of ammonium chloride and (2) mass of water.

For (1), just convert 8.50 grams ammonium chloride to moles, Use FORMULA WEIGHT.

$$8.50 \text{ g } NH_4Cl \times \frac{\text{mol } NH_4Cl}{53.491 \text{ g } NH_4Cl} = 0.1589052364 \text{ mol } NH_4Cl$$

For (2), subtract the mass of ammonium chloride from the total mass of solution. That will give the mass of water since the solution contains only ammonium chloride and water.

$$100 \text{ g solution} - 8.50 \text{ g } NH_4Cl = 91.50 \text{ g } H_2O = 0.09150 \text{ kg } H_2O$$

After converting mass of water to kg, just divide to get the molality...

$$\frac{\text{mol } NH_4Cl}{\text{Kg } H_2O} = \frac{0.1589052364 \text{ mol } NH_4Cl}{0.09150 \text{ kg } H_2O} = \boxed{1.74 \text{ m } NH_4Cl}$$

An aqueous solution is 8.50% ammonium chloride by mass. The density of the solution is 1.024 g/mL
Find: molality and molarity.

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

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

mass percent (defintion)   molarity (defintion)

As before, assume a basis of 100 grams solution. This lets us know the mass of ammonium chloride (8.50 g) and gives us a place to start.

For (1), we convert 8.50 g ammonium chloride to moles. (We did that already on the last page)

0.1589052364 mol NH_4Cl

For (2), we need to convert 100 g solution to volume. Use DENSITY.

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

After changing the volume to liters, just divide to get the molarity.

$$\frac{\text{mol } NH_4Cl}{\text{L solution}} = \frac{0.1589052364\text{ mol } NH_4Cl}{0.09765625\text{ L}} = \boxed{1.63\text{ M } NH_4Cl}$$

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

$$\underbrace{\Delta T_f}_{\text{Freezing point depression: The amount the freezing temperature is LOWERED by the solute.}} = \underbrace{K_f}_{\text{Freezing point depression constant (for SOLVENT)}} \times \underbrace{C_m}_{\text{concentration of solute (molality)}}$$

- 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?

$K_{f,\text{benzene}} = 5.065\ ^{o}C/m$, $T_{f,\text{benzene}} = 5.455^{o}C$ (see p500 9th p509, 10th)

$$\underbrace{\Delta T_f}_{5.455^oC - 4.880^oC = 0.575^oC} = \underbrace{K_f}_{5.065^oC/m} \times C_m$$

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

kg benzene ← 0.1000 kg (100g)

We need to find Cm so we can find the moles of unknown (needed for molecular weight)

Start by finding molal cocnentration of unknown, Cm:

$$(0.575^oC) = (5.065\ ^oC/m) \times C_m$$

$$0.1135241856\ m \text{ unknown} = C_m$$

Plug into definintion of Cm to find the MOLES UNKNOWN:

$$C_m = \frac{\text{mol unknown}}{\text{kg benzene}} \rightarrow 0.1135241856\ m \text{ unknown} = \frac{\text{mol unknown}}{0.1000 \text{ kg benzene}}$$

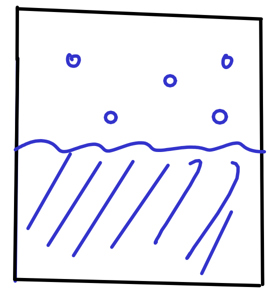
$$\text{mol unknown} = 0.01135241 86 \text{ mol}$$

Now, find the molecular weight.

$$MW = \frac{\text{mass unknown}}{\text{mol unknown}} = \frac{2.500\text{ g}}{0.0113524186 \text{ mol}} = \boxed{22\bar{0}\text{ g/mol}}$$

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.

$$\Delta T_b = K_b \times C_m$$

C_m: concentration of solute (molality)

K_b: Boiling point elevation constant (for SOLVENT) (p506, 9th; p509, 10th)

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