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

$$\frac{8.50\ g\ NH_4Cl}{100\ g\ solution} \longrightarrow \frac{?\ mol\ NH_4Cl}{?\ kg\ H_2O}$$

mass percent — molality

ASSUME A BASIS of 100 g solution. This means that there's 8.50 grams of ammonium chloride.
To start, convert 8.50 grams ammonium chloride to moles.

$$8.50\ g\ NH_4Cl \times \frac{mol\ NH_4Cl}{53.491\ g\ NH_4Cl} = 0.158905236\ 4\ mol\ NH_4Cl$$

Find the mass of water by subtraction (the solution is ammonium chloride in water).

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

So the molality is ...

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

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mass percent   molarity

We'll keep the basis we used for the last calculation (100 g solution). Doing that lets us take advantage of the fact that we already calculated moles of ammonium chloride.

0.158905236 mol NH_4Cl

Find volume of solution using the given density of solution. (We already know solution mass ... 100 g)

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

Find molarity ...

$$M = \frac{0.158905236\ mol\ NH_4Cl}{0.09765625\ L} = \boxed{1.63\ M\ NH_4Cl}$$

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?

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

$\Delta T_f = K_f \times C_m$

K_f → 5.065 °C/m

ΔT_f → 5.455°C − 4.880°C = 0.575°C

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

Since molecular weight is the mass of unknown per mole of unknown, we need to find how many moles unknown we have in the solution. We already know the mass of unknown is 2.500 grams.

First, calculate the MOLAL CONCENTRATION, Cm:

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

$$C_m = 0.113524186\ m$$

The molal concentration, Cm, is the number of moles of unknown PER KILOGRAM benzene. To find the actual number of moles of unknown, we need to multiply by the mass of benzene used!

100.0 g benzene → 0.1000 kg benzene

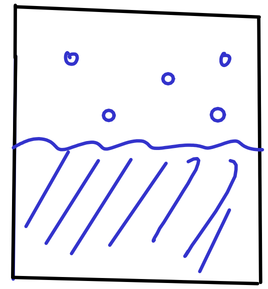
$$0.1000\text{ kg benzene} \times \frac{0.113524186\text{ mol unknown}}{\text{kg benzene}} = 0.0113524186\text{ mol unk.}$$

Now find 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) (p500, 9th / p509, 10th)

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?

$T_b = 118.5°C$ $K_b = 3.08 °C/m$ (see p500 for data p509, 10th)

$\Delta T_b = K_b \times C_m$ (K_b = 3.08 °C/m)

$C_m = \frac{\text{mol } S_8}{\text{kg AA}}$ solute / solvent

S_8: 8×32.07 = 256.56 g S_8 = mol S_8

Find Cm, using the information given in the problem.

$$2.817 \text{ g } S_8 \times \frac{\text{mol } S_8}{256.56 \text{ g } S_8} = 0.0109798677 \text{ mol } S_8$$

$$\frac{0.0109798677 \text{ mol } S_8}{0.1000 \text{ kg AA}} = 0.109798775 \text{ m} = C_m$$

Find DELTA Tb...

$$\Delta T_b = (3.08 °C/m)(0.109798775 \text{ m}) = 0.338°C$$

This is the boiling point ELEVATION

To get the new bolliling point, add the original boiling point (Tb) to the elevation ...

$$118.5°C + 0.338°C = \boxed{118.8°C}$$