MOLARITY and the other concentration units

- To convert between molarity and the other two concentration units we've studied, you have to know more about the solution. For example:

$$\text{molarity} \quad \frac{\text{moles A}}{\text{L solution}} \longrightarrow \text{molality} \quad \frac{\text{moles A}}{\text{kg solvent}}$$

* To perform this conversion, you can assume a liter of solution, which will give you the number of moles present. But you've then got to have a way to convert the volume of SOLUTION to the mass of the SOLVENT. How?

* You need DENSITY (which depends on temperature). The density of the solution will allow you to find the total mass of the solution.

* If you subtract out the mass of the SOLUTE, then what you have left is the mass of the SOLVENT. Express that in kilograms, and you have all the information you need to find molality!

* You'll run into the same situation when you use any of the other mass or mole based units. DENSITY is required to go back and forth between MOLARITY and these units.

Example: If a solution is 0.688 m citric acid, what is the molar concentration (M) of the solution? The density of the solution is 1.049 g/mL

$H_3C_6H_5O_7$: 192.125 g/mol "CA"

$\frac{0.688 \text{ mol CA}}{\text{Kg solvent}}$ → $\frac{? \text{ mol CA}}{\text{L solution}}$

definition of molality

definition of molarity

To start the problem, ASSUME A BASIS for the calculation. We will assume that we have an amount of solution that contains 1 kg of the solvent. This means that the solution also must contain 0.688 moles of CA. Now all we have to do is to find out the volume of the solution in liters. We need to use DENSITY to relate the mass of the solution to the volume we need. But first, we need to calculate the mass of solution. It contans both solvent and CA!

$$0.688 \text{ mol CA} \times \frac{192.125 \text{ g CA}}{\text{mol CA}} = 132.182 \text{ g CA}$$

$$\text{mass solution} = 1000 \text{ g solvent} + 132.182 \text{ g CA} = 1132.182 \text{ g}$$

$$1132.182 \text{ g solution} \times \frac{\text{mL}}{1.049 \text{ g}} = 1079.296473 \text{ mL} = 1.079296473 \text{ L}$$

$$M = \frac{\text{mol CA}}{\text{L solution}} = \frac{0.688 \text{ mol CA}}{1.079296473 \text{ L}} = \boxed{0.637 \text{ M CA}}$$

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

$$\frac{\text{mol } NH_4Cl \ (1)}{\text{Kg } H_2O \ (2)}$$

definition of molality

1) Convert 8.50 grams ammonium chloride to moles. Use formula weight.

2) Find mass of water in solution by subtraction.

(1) $8.50 \text{ g } NH_4Cl \times \frac{\text{mol } NH_4Cl}{53.491 \text{ g } NH_4Cl} = 0.158905236 \text{ mol } NH_4Cl$

(2) $100. \text{ g solution} - 8.50 \text{ g } NH_4Cl = 91.5 \text{ g } H_2O = 0.0915 \text{ Kg } H_2O$

$$m = \frac{\text{mol } NH_4Cl}{\text{K } H_2O} = \frac{0.158905236 \text{ mol } NH_4Cl}{0.0915 \text{ Kg } H_2O} = \boxed{1.74 \, m \; NH_4Cl}$$

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

$$\frac{\text{mol } NH_4Cl \quad (1)}{\text{L solution} \quad (2)}$$

definition of molarity

1) Convert 8.50 grams ammonium chloride to moles. Use formula weight. (Already done for molality calculation)

2) Convert 100 grams solution to volume using density.

(1) 0.1589052364 mol NH_4Cl

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

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

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

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

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?

$K_{f,\text{benzene}} = 5.12\ {}^{o}C/m$, $T_{f,\text{benzene}} = 5.5\ {}^{o}C$ (Values from p619 OpenStax Table 11.2)

$$\underbrace{\Delta T_f}_{5.5^oC - 4.880^oC = 0.62^oC} = \underbrace{K_f}_{5.5^oC/m} \times C_m$$

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

Molecular weight is the grams of the sample divided by the moles of sample. We already know the weight of the sample, so moles is what we'll need to find. To get moles, first find Cm (the molal concentration) ... and then we can use it to find moles of unknown!

$0.62^oC = (5.5\ {}^oC/m) \times C_m$; $C_m = 0.112727272727\ m$ unknown

but $C_m = \frac{\text{mol unknown}}{\text{Kg benzene}}$; $0.1127272727\ m = \frac{\text{mol unknown}}{0.1000\text{ Kg benzene}}$

$0.01127272273 = \text{mol unknown}$

$$MW = \frac{\text{g unknown}}{\text{mol unknown}} = \frac{2.500\text{g}}{0.01127272273\text{ mol}} = \boxed{220\text{ g/mol}}$$