MOLARITY and the other concentration units

- To convert between molarity and the other three 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"

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

definition of molality definition of molarity

To solve the problem, let's ASSUME A BASIS of 1 kg of solvent (the "bottom" of the unit that we start with). If we do that, we know that there are 0.688 moles of CA, and all we have to do is to find the VOLUME of solution. To find volume, we'll need to use the DENSITY ... which means we'll need to calculate the mass of the solution (the solven PLUS the CA).

192.125 g CA = mol CA

$$0.688 \text{ mol CA} \times \frac{192.125 \text{ g CA}}{\text{mol CA}} = 132.182 \text{ g CA}$$ Now find total mass!

(1 kg)

1000 g solvent + 132.182 g CA = 1132.182 g solution

Find volume!

$$1132.182 \text{ g solution} \times \frac{\text{mL}}{1.049 \text{ g}} = 1079.3 \text{ mL} = 1.0793 \text{ L}$$

$$M = \frac{\text{mol CA}}{\text{L soln}} = \frac{0.688 \text{ mol CA}}{1.0793 \text{ L}} = \boxed{0.637 \text{ M CA}}$$

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

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

mass percent (definition)   molality (definition)

ASSUME A BASIS of 100 g solution. This means we know we have 8,50 grams of ammonium chloride. Find moles ammonium chloride by converting 8.50 grams (use formula weight). Find mass water by subtraction.

$$8.50\text{ g } NH_4Cl \times \frac{\text{mol } NH_4Cl}{53.491\text{ g } NH_4Cl} = 0.1589052364\text{ mol } NH_4Cl \quad ①$$

The mass of water is ...

$$100\text{ g solution} - 8.50\text{ g } NH_4Cl = 91.50\text{ g } H_2O = 0.0915\text{ Kg } H_2O \quad ②$$

So molality is ...

$$m = \frac{\text{mol } NH_4Cl}{\text{Kg } H_2O} = \frac{0.1589052364\text{ mol } NH_4Cl}{0.0915\text{ Kg } H_2O} = \boxed{1.74\text{ m } NH_4Cl}$$

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

$$\frac{8.50 \text{ g } NH_4Cl}{100 \text{ g solution}} \longrightarrow \frac{\text{mol } NH_4Cl}{\text{L solution}}$$

mass percent (definition)   molarity (definition)

Like before, assume 100 grams of solution as our basis. That means we have 8.50 grams ammonium chloride (and that we know the moles of ammonium chloride, too ... since we calculated it!). Now all we need to do is to find the VOLUME of solution. Use the density along with the 100 g solution we assumed initially.

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

From our previous work, 8.50 grams of ammonium chloride is equal to ...

0.1589052364 mol NH_4Cl

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

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?

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

$\Delta T_f = K_f \times C_m$

K_f: $5.065\ {}^{o}C/m$

$0.575\,{}^{o}C = 5.455\,{}^{o}C - 4.880\,{}^{o}C$

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