More on MOLARITY

To prepare a solution of a given molarity, you generally have two options:

1. Weigh out the appropriate amount of solute, then dilute to the desired volume with solvent (usually water)"

"stock solution"

2. Take a previously prepared solution of known concentration and DILUTE it with solvent to form a new solution

- Use DILUTION EQUATION

The dilution equation is easy to derive with simple algebra.

$$M \times V$$

$$\frac{mol}{L} \times L = \text{moles solute}$$

... but when you dilute a solution, the number of moles of solute REMAINS CONSTANT. (After all, you're adding only SOLVENT)

$$M_1V_1 = M_2V_2$$

before diution

after dilution

Since the number of moles of solute stays the same, this equality must be true!

$$M_1V_1 = M_2V_2$$ ... the "DILUTION EQUATION"

M_1 = molarity of concentrated solution

V_1 = volume of concentrated solution

M_2 = molarity of dilute solution

V_2 = volume of dilute solution ←(TOTAL VOLUME, NOT the volume water added!)

The volumes don't HAVE to be in liters, as long as you use the same volume UNIT for both V_1 and V_2

Example: Take the 0.500 M sodium sulfate we discussed in the previous example and dilute it to make 150. mL of 0.333 M solution. How many mL of the original solution will we need to dilute?

$$M_1V_1 = M_2V_2$$

$M_1 = 0.500\ M$

$V_1 = ?$

$M_2 = 0.333\ M$

$V_2 = 150.\ mL$

$$(0.500\ M)\,V_1 = (0.333\ M)(150.\ mL)$$

$$V_1 = \boxed{99.9\ mL \text{ of } 0.500\ M \text{ stock}}$$

Measure out 99.9 mL of 0.500 M sodium sulfate, then dilute to 150. mL with DI water. (You can do this in a single 250 mL graduated cylinder if you're in a hurry!)

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:

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

* 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"

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

molality (definition) molarity (defintion)

To start the problem, ASSUME A BASIS of some amount of CA solution. This will work because concentration is a RATIO and doesn't actually depend on the amount of solution we have. ASSUME A BASIS of 1 kg of solvent. This means that we have 0.688 grams of CA and only have to figure out what volume of solution we have.

We can calculate the volume of the solution using DENSITY, but we first need to know what the solution weighs. The solution is made of solvent and CA, so let's find the mass of CA.

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

So the solution weighs ...

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

Use DENSITY ...

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

Find MOLARITY...

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

An aqueous solution is 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

Find molality. Start with defintions.

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

mass percent   molality

ASSUME A BASIS of 100 g solution. That means we have 8.50 grams of ammonium chloride. We can change 8.50 grams of ammonium chloride to moles to find the top half of molality, then we just have to figure out the mass of water to get the bottom half.

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

Now find mass water. Subtract.

$$100\text{ g solution} - 8.50\text{ g } NH_4Cl = 91.50\text{ g } H_2O = 0.0915\text{ kg } H_2O$$

FInd MOLALITY.

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

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

Find molarity. Start with defintions.

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

mass percent molarity

As before, assume 100 g solution as our basis. This means we have 8.50 grams of ammonium chloride. We'll need to calculate moles of ammonium chloride (which we already did to find molality), then find the volume of solution.

0.158905236 mol NH_4Cl (from previous page)

FInd volume of solution.

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

FInd MOLARITY

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

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?

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

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

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

We want to find Cm and then use it to find moles unknown.

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

$$0.1135241856\ m = C_m$$

Now find moles unknown. Do that by multiplying the amount of benzene used (in kg) by the molality.

$$0.1000\ \text{kg benzene} \times \frac{0.1135241856\ \text{mol unknown}}{\text{kg benzene}} = 0.01135241856\ \text{mol unknown}$$

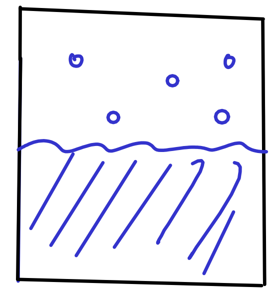
Since we know both the moles and mass of the unknown, we can calculate molecular weight.

$$MW = \frac{\text{g unknown}}{\text{mol unknown}} = \frac{2.500\ \text{g unknown}}{0.01135241856\ \text{mol unknown}} = \boxed{220.\ \text{g/mol}}$$

VAPOR PRESSURE LOWERING

- Described by RAOULT'S LAW

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



P_A = partial pressure of the VAPOR of solvent molecules.

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?

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

$$\Delta T_b = \underbrace{K_b}_{3.08°C/m} \times C_m \qquad C_m = \frac{\text{mol } S_8}{\text{kg } HC_2H_3O_2}$$

We need to calculate Cm. To get that, we need to find out how many moles of sulfur was dissolved and the mass of acetic acid.

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

S_8: S—8 x 32.07

$256.56 \text{ g } S_8 = \text{mol } S_8$

$$C_m = \frac{0.0113524186 \text{ mol unknown}}{0.1000 \text{ kg } HC_2H_3O_2} = 0.1097988775 \text{ m } S_8$$

Now find boiling point ELEVATION.

$$\Delta T_b = (3.08°C/m)(0.1097988775 \text{ m } S_8) = 0.338°C$$

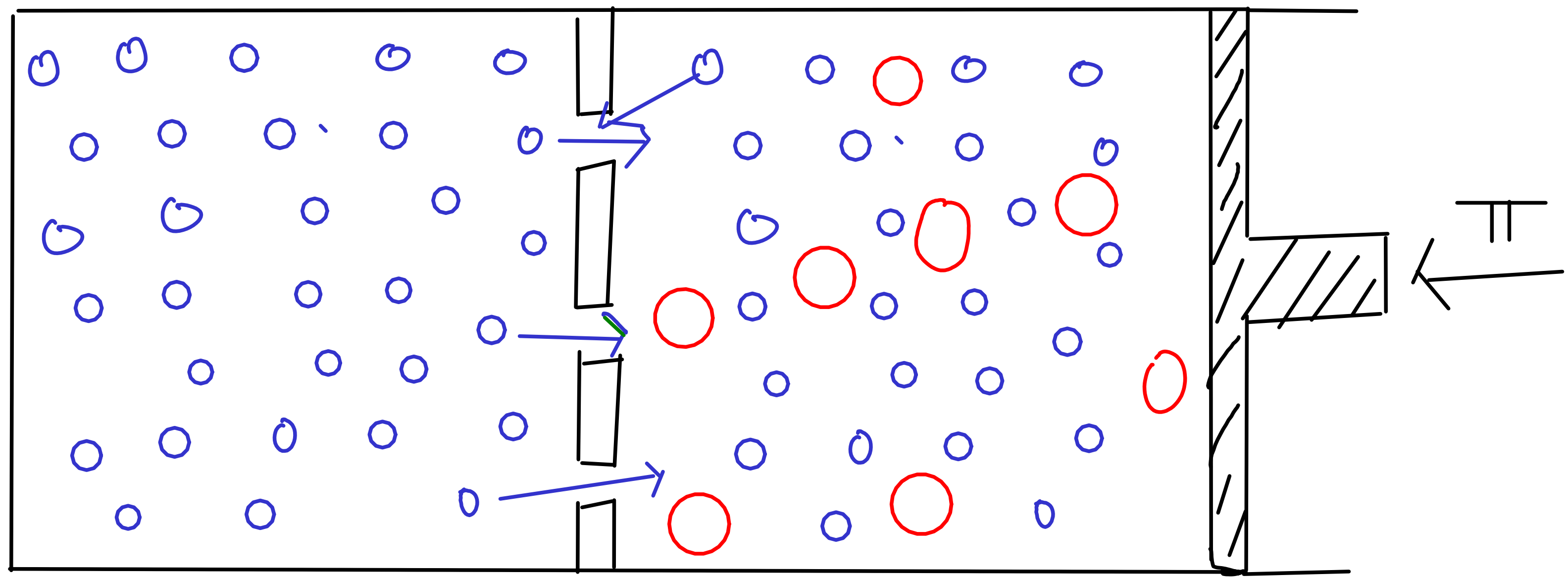
Find the new boiling point.

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

OSMOTIC PRESSURE

permits flow of solvent, but not solute particles

- OSMOSIS: the flow of solvent molecules through a SEMIPERMEABLE membrane to equalize concentration of solute on each side of the membrane.



The rate of solvent migration towards the RIGHT is greater than that towards the LEFT.

If you apply enough pressure to the piston, osmosis will not occur. This pressure is called the OSMOTIC PRESSURE

$$\pi = M \times R \times T$$

temperature

ideal gas constant

molar concentration of solute

IONIC COMPOUNDS and colligative properties

- Ionic compounds DISSOCIATE in water into their component ions. Each ion formed can act as a solute and influence the colligative properties!

$$NaCl(s) \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$$

2 ions!

... so the concentration of IONS here is TWICE the nominal NaCl concentration.

$$CaCl_2(s) \longrightarrow Ca^{2+}(aq) + 2\,Cl^{-}(aq)$$

3 ions!

... so the concentration of IONS here is THREE TIMES the nominal calcium chloride concentration.

- Ions interact with each other in solution, so unless an ionic solution is DILUTE, the effective concentrations of ions in solution will be less than expected. A more advanced theory (Debye-Huckel) covers this, but we'll assume that our solutions are dilute enough so that we can use the concentration of the ions in solution to determine the colligative properties!