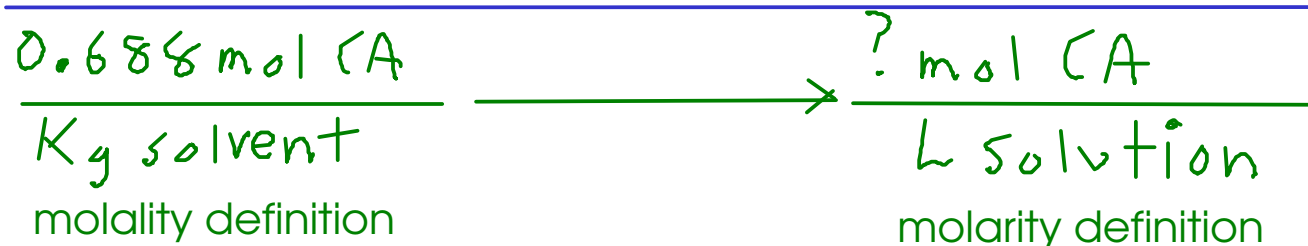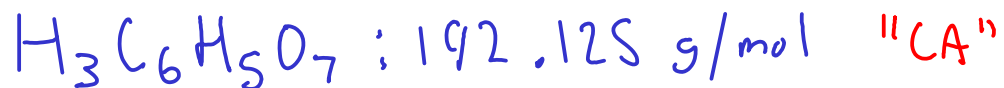Example: If a solution is 0.688 m citric acid, what is the molar concentration (M) of the solution? The density of the solution is 1.049 g/mL

$$H_3C_6H_5O_7 : 192.125 \text{ g/mol} \quad "CA"$$

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

molality definition   molarity definition

To solve the problem, we'll ASSUME A BASIS of 1 kg of solvent. This means we now know the moles of CA (0.688 mol) dissolved and we can start working the problem from there. We now just need to calculate the volume of the solution! To figure that out, remember that we assumed 1 kg (1000g) solvent. If we add the mass of CA to that, we'll have the MASS of the solution. Then we can use density to find volume.

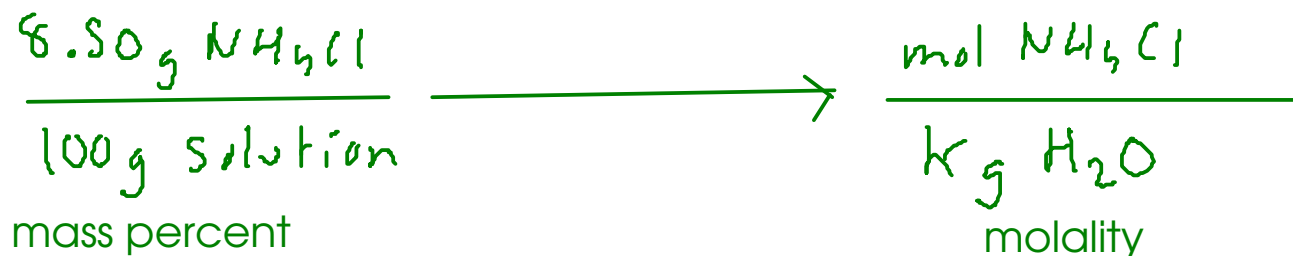$$0.688 \text{ mol } CA \times \frac{192.125 \text{ g } CA}{\text{mol } CA} = 132.182 \text{ g } CA$$

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

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

Molarity equals the moles CA / the volume solution

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

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

$NH_4Cl : 53.491 \ g/mol \qquad H_2O: 18.016 \ g/mol$

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

mass percent $\qquad\qquad\qquad\qquad$ molality

Assume a basis of 100g solution. This means that the solution contains 8.50 grams of ammonium chloride. We can convert the mass of ammonium chloride to moles:

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

Now that we have moles ammonium chloride, find the mass of water:

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

So the molality is:

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

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

$NH_4Cl : 53.491 \text{ g/mol} \qquad H_2O : 18.016 \text{ g/mol}$

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

mass percentage $\qquad\qquad\qquad$ molarity

Again, assume a basis of 100 g solution. That means we laready know the mass (8.50 g) and
moles of ammonium chloride, and all we have to do is change the mass of solution
to volume using density.

$0.1589052364 \text{ mol } NH_4Cl$ (calculated on previous page)

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

Molarity equals moles ammonium chloride / volume of solution in L

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

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

$$\Delta T_f = K_f \times \frac{C_m}{}$$

⎿ concentration of solute (molality)

⎿ Freezing point depression constant (for SOLVENT)

⎿ 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?

$$K_{f, benzene} = 5.065 \,^{\circ}C/m \quad , \quad T_{f, benzene} = 5.455 \,^{\circ}C \quad \left(\begin{array}{l} \text{see} \\ \text{p 500 9th} \\ \text{p 509, 10 th} \end{array}\right)$$

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

$$\qquad 5.065 \,^{\circ}C/m$$

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

To find molecular weight, we need to know the MOLES of the unknown. Then, we just divide 2.500 g / ?? mol

We can calculate the molal concentration, Cm:

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

$$C_m = 0.1135241856 \text{ mol unk}/Kg \text{ benzene}$$ ... now we need to find moles unknown!

Calculate moles unknown from the amount of benzene actually used! (100.0 g, or 0.1000 kg)

$$0.1000 Kg \text{ benzene} \times \frac{0.1135241856 \text{ mol unk}}{Kg \text{ benzene}} = 0.0113524186 \text{ mol unk}$$
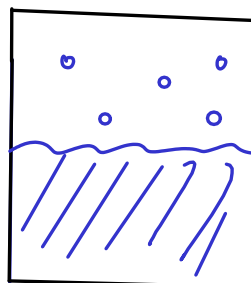
Now find the molecular weight:

$$MW = \frac{g \text{ unknown}}{mol \text{ unk}} = \frac{2.500g \text{ unk}}{0.0113524186 \text{ mol unk}} = \boxed{220 \text{ g/mol}}$$

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.

mole fraction of component A

vapor pressure of pure component A (depends on temperature)

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...

$$\Delta P = P_A^* \times X_B$$

mole fraction of component B (the SOLUTE in a two-component mixture)

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.

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

concentration of solute (molality)

Boiling point elevation constant (for SOLVENT) $\left(\begin{array}{c} p \, S \, 0 \, 6 \, , 9 \, th \\ p \, S \, 0 \, 9 \, , 10 \, th \end{array} \right)$

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?

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

$$\Delta T_b = K_b \times [m] \frac{mol\ S_8}{Kg\ HC_2H_3O_2} \qquad S_8: \frac{8 \times 32.07}{256.56\,g\,S_8 = mol\ S_8}$$
$$\underset{3.08\,°C/m}{}$$

First, find moles sulfur and mass (in kg) of acetic acid used... That will give us Cm.

$$2.817\,g\ S_8 \times \frac{mol\ S_8}{256.56\,g\ S_8} = 0.01097798877\ mol\ S_8$$

$$100.0\,g\ HC_2H_3O_2 = 0.1000\ Kg\ HC_2H_3O_2$$

$$C_m = \frac{0.01097798877\ mol\ S_8}{0.1000\ Kg\ HC_2H_3O_2} = 0.1097988775\ m\ S_8$$

Now, we can find delta Tb:

$$\Delta T_b = \left(3.08\,°C/m\right)\left(0.1097988775\ m\ S_8\right) = 0.338\,°C$$

Calculate the new boiling temperature:

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

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.

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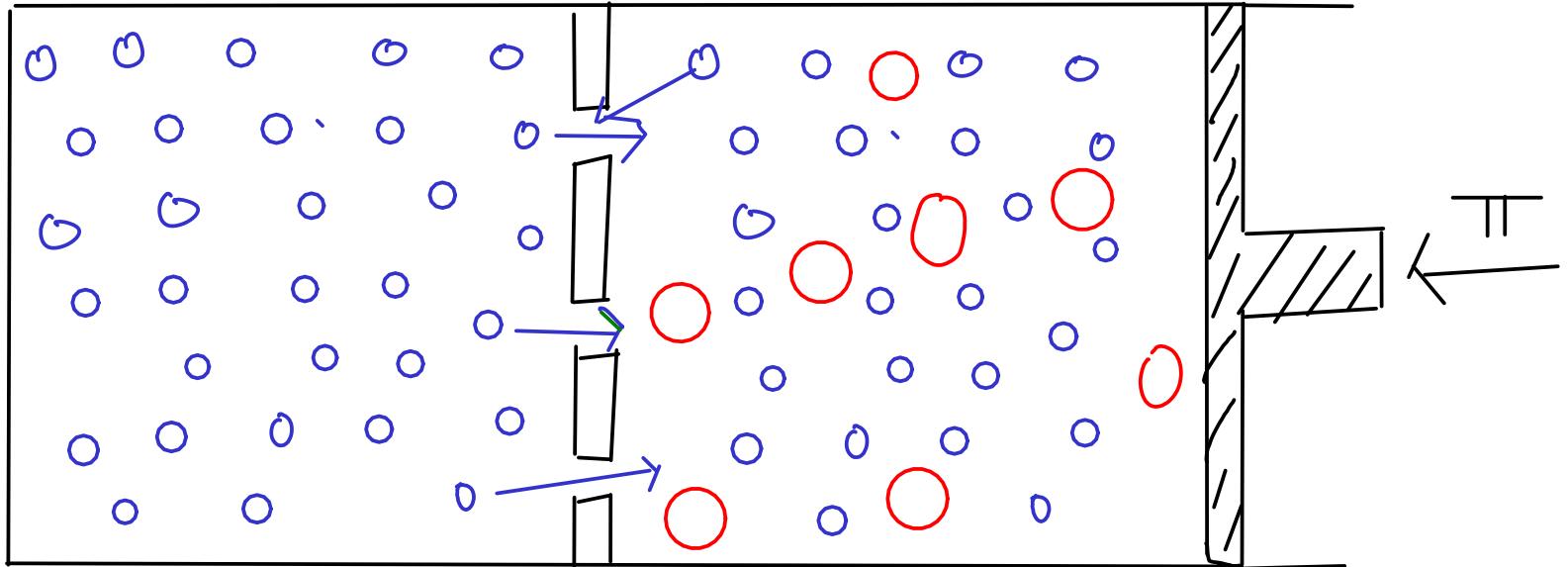
The rate of solvent migration towards the RIGHT is greater than that towards the LEFT.

If you apply enough pressure to the piston, osmosis will not occur. This pressure is called the OSMOTIC PRESSURE
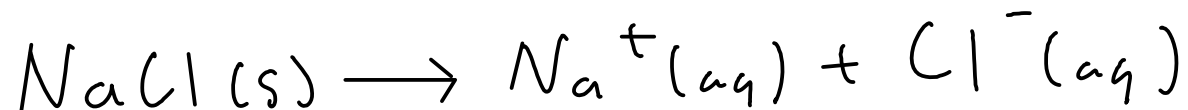
$$\pi = M \times R \times T$$

temperature

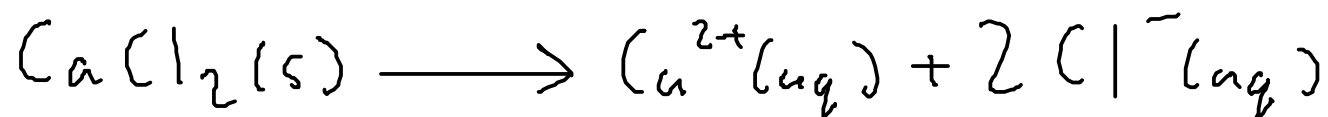ideal gas constant

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!

If you are at an altitude high enough for the boiling point of water to be 95.00 C, what amount of sodium chloride would you need to add to 1.000 kg of water to raise the boiling point to 100.00 C?

$$K_b = 0.512 \ ^\circ C/m \qquad NaCl: 58.443 \ g/mol$$

$$\Delta T_b = K_b \times C_m$$
$$ = 0.512 \ ^\circ C/m$$
$$100.00 \ ^\circ C - 95.00 \ ^\circ C = 5.00 \ ^\circ C$$

Let's find Cm first ...

(concentration of ions, not the same as concentration of NaCl!)

$$5.00 \ ^\circ C = (0.512 \ ^\circ C/m) \times C_m$$
$$C_m = 9.765625 \ m \ \underline{\underline{ions}}$$

Now, find how many moles of ions need to be dissolved...

$$1.000 \ Kg \ H_2O \times \frac{9.765625 \ mol \ ions}{Kg \ H_2O} = 9.765625 \ mol \ ions$$

The ions are formed from NaCl dissolving! $\quad NaCl \rightarrow Na^+ + Cl^- \ (2 \ ions!)$

$$9.765625 \ mol \ ions \times \frac{mol \ NaCl}{2 \ mol \ ions} = 4.8828125 \ mol \ NaCl$$

Find the mass of NaCl:

$$4.8828125 \ mol \ NaCl \times \frac{58.443g \ NaCl}{mol \ NaCl} = \boxed{285g \ NaCl}$$