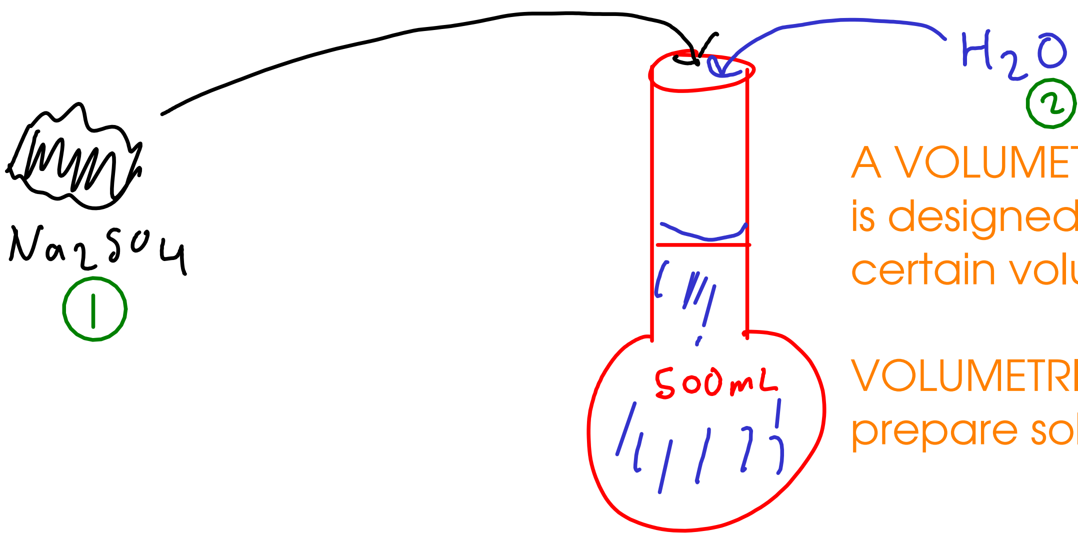
... we use MOLARITY so much because it's easy to work with. It is easier to measure the VOLUME of a liquid solution than it is to measure mass.

Example: How would we prepare 500. mL of 0.500 M sodium sulfate in water?

Na_2SO_4 : 142.04 g/mol

Dissolve the appropriate amount of sodium sulfate into enough water to make 500. mL of solution.



A VOLUMETRIC FLASK is a flask that is designed to precisely contain a certain volume of liquid.

VOLUMETRIC FLASKS are used to prepare solutions.

volumetric flask

Calculate the moles of sodium sulfate present in 500 mL of 0.500 M solution, then convert to grams!

M = mol solute/L 0.500 mol Na_2SO_4 = L

$$500.\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} \times \frac{0.500\text{ mol } Na_2SO_4}{\text{L}} = 0.250\text{ mol } Na_2SO_4$$

$$0.250\text{ mol } Na_2SO_4 \times \frac{142.04\text{ g } Na_2SO_4}{\text{mol } Na_2SO_4} = \boxed{35.5\text{ g } Na_2SO_4}$$

Dissolve 35.5 g of sodium sulfate in enough water to make 500. mL of solution.

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)

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)

$$M_1 V_1 = M_2 V_2$$

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

before diution

after dilution

$$M_1V_1 = M_2V_2$$ ... 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

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

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?

$M_1V_1 = M_2V_2$ $\quad M_1 = 0.500\,M$ $\quad V_1 = ???$

$M_2 = 0.333\,M$ $\quad V_2 = 150.\,mL$

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

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

Take 99.9 mL of 0.500 M solution, and dilute with water to 150. mL total volume.

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:

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

* 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 (m) → molarity (M)

1- Assume we have 1 kg of solvent. This makes the number of moles of CA = 0.688 mol.
2 - Find the volume of the solution. We can use density, but there is a small catch: We know the mass of the solvent, and need the mass of the SOLUTION. So, we have to find out the mass of the solute first.

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

mass of solution = 132.182 g CA + 1000 g solvent = 1132.182 g solution

Now, we can use the density to find the volume!

$$1132.182 \text{ g solution} \times \frac{\text{mL}}{1.049 \text{ g}} \times \frac{10^{-3} \text{ L}}{\text{mL}} = 1.079296 \text{ L}$$

$$\underline{M} = \frac{0.688 \text{ mol CA}}{1.079296 \text{ L}} = \boxed{0.637\ \underline{M}\ \text{CA}}$$

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

NH_4Cl : 53.491 g/mol H_2O: 18.016 g/mol

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

mass percent molality

Find moles ammonium chloride: Convert 8.50 g to moles!

$$8.50\text{ g } NH_4Cl \times \frac{\text{mol}}{53.491\text{ g } NH_4Cl} = 0.15891\text{ mol } NH_4Cl$$

Find mass water.

$$100\text{ g solution} - 8.50\text{ g } NH_4Cl = 91.5\text{ g water} = 0.0915\text{ kg}$$

Find molality.

$$m = \frac{0.15891\text{ mol } NH_4Cl}{0.0915\text{ kg}} = \boxed{1.74\text{ m } NH_4Cl}$$

$$\frac{8.50\text{ g } NH_4Cl}{100\text{ g solution}} \longrightarrow \frac{\text{mol } NH_4Cl}{\text{mol } NH_4Cl + \text{mol } H_2O} \leftarrow 0.15891\text{ mol } NH_4Cl$$

mass percent mole fraction

Find moles water: Use mass we calculated before. $91.5\text{ g } H_2O \times \frac{\text{mol}}{18.016\text{ g}} = 5.07881\text{ mol } H_2O$

$$X_{NH_4Cl} = \frac{0.15891\text{ mol } NH_4Cl}{0.15891\text{ mol } NH_4Cl + 5.07881\text{ mol } H_2O} = \boxed{0.0303 = X_{NH_4Cl}}$$

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

NH_4Cl : 53.491 g/mol H_2O: 18.016 g/mol

$$\frac{8.50\text{ g } NH_4Cl}{100\text{ g solution}} \longrightarrow \frac{?\text{ mol } NH_4Cl}{?\text{ L solution}} \leftarrow 0.15891\text{ mol } NH_4Cl$$

mass percent   molarity

Find the volume of 100g of solution using density

$$100\text{ g solution} \times \frac{\text{mL}}{1.024\text{ g}} \times \frac{10^{-3}\text{ L}}{\text{mL}} = 0.09765625\text{ L solution}$$

Find molarity:

$$\underline{M} = \frac{0.15891\text{ mol } NH_4Cl}{0.09765625\text{ L solution}} = \boxed{1.63\ \underline{M}\ NH_4Cl}$$

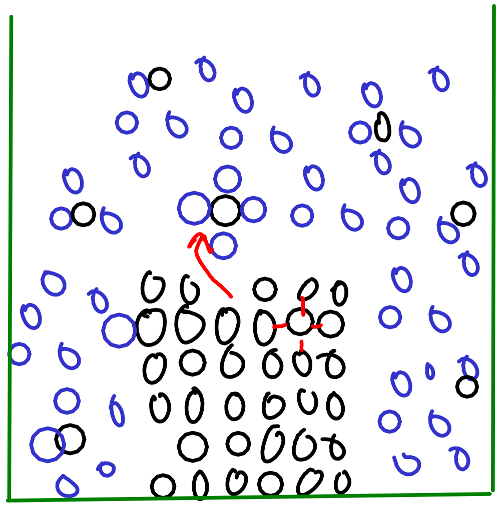
HOW THINGS DISSOLVE

- Let's look at how things dissolve into water, since aqueous solutions are quite common.

sucrose (table sugar)

$$C_{12}H_{22}O_{11}(s) \underset{}{\overset{H_2O}{\rightleftarrows}} C_{12}H_{22}O_{11}(aq)$$

... what happens?



- Water molecules pull the sugar molecules out of the sugar crystal and into solution.

- Attractions between sugar molecules and water allow this to happen.

- The solubility of the sugar depends on how well water and sugar interact (HYDRATION) versus how well the sugar molecules are held in the crystal (LATTICE ENERGY)

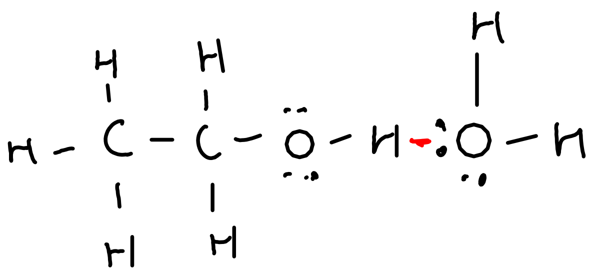
- "like dissolves like": Substances held together by similar (or at least compatible) kinds of attractive forces can dissolve in each other. Substances that are held together by very different kinds of attractive forces will not dissolve in one another!

Consider WATER:

HYDROGEN BONDS

↓

Water mixes well with other substances that can hydrogen bond, like ETHANOL!



POLAR

↓

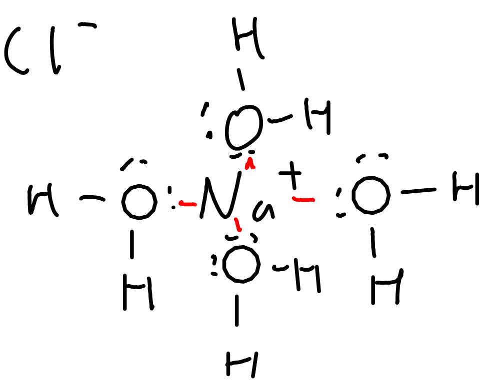
Water can dissolve polar substances!

(SUCROSE is polar!)

↓

Since IONIC BONDS are also interactions between opposite charges (You can think of an ionic bond here as an extreme case of dipole-dipole interaction), many IONIC SUBSTANCES will also dissolve in water!

$Na^+ - Cl^-$



"ion-dipole" interactions

SMALL (little London force)

↓

large and/or nonpolar solutes do not dissolve well in water!

(example: OILS and WAXES)

MOLECULAR AND IONIC SOLUTIONS

- MOLECULAR solutions:

Contain MOLECULES dissolved in one another.

1 - Any mixture of GASES

- all gases mix with one another, since gas molecules (effectively) do not interact with one another.

2 - Liquids

- Liquids dissolve well in one another only if they are held together by similar kinds of forces

3 - Solids and liquids

- MOLECULAR SOLIDs will dissolve well in liquids if they are held together by similar forces.

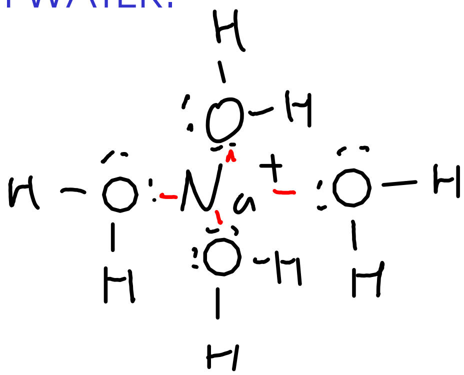
- IONIC SOLIDS will sometimes dissolve in POLAR liquids, but not in nonpolar liquids

- COVALENT NETWORK solids don't generally dissolve well in other substances

IONIC solutions

- form when ions from IONIC SUBSTANCES interact with POLAR solvents - often WATER.

The charged ends of the water molecule HYDRATE the ions.



VS.

- The solubility of an ionic compound depends on whether HYDRATION (attraction of water molecules for an ion) is greater than LATTICE ENERGY - the attrraction of ions in a crystal lattice for one another..

- SMALLER IONS are usually easier to enclose in water than larger ones, and ions with larger charges are attracted to water molecules.

- But solubility is also determined by LATTICE ENERGY - which holds the solid ionic compound together. Ions with high charges tend to be strongly attracted to other ions in a crystal, meaning lattice energy is high. Smaller ions also tend to have higher lattice energies. Lattice energy and hydroation are competing trends!

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

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

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?

$$K_{f,\,benzene} = 5.065\ {}^{\circ}C/m \quad , \quad T_{f,\,benzene} = 5.455\ {}^{\circ}C$$

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

$$C_m = \frac{\text{moles unknown}}{\text{kg solvent}} \leftarrow 0.1000 \text{ kg solvent}$$

$$\Delta T_f = 5.455\ {}^{\circ}C - 4.880\ {}^{\circ}C = 0.575\ {}^{\circ}C$$

* If we know the freezing point depression, we can calculate the CONCENTRATION of the solution we prepared in molality units.
* If we know the concentration, we can find out how many MOLES of unknown are in the solution, and then get the molecular weight.

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

$$0.113524\ m = C_m$$

Find the moles of unknown, starting from the amount of solvent used

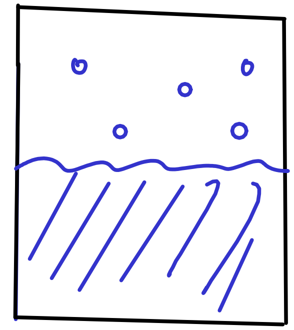
$$0.1000 \text{ kg benzene} \times \frac{0.113524 \text{ mol unk}}{\text{kg benzene}} = 0.0113524 \text{ mol unk}$$

Find molecular weight by dividing mass and moles.

$$MW = \frac{\text{mass unk}}{\text{mol unk}} = \frac{2.500 \text{ g}}{0.0113524 \text{ mol}} = \boxed{220 \text{ g/mol}}$$

VAPOR PRESSURE LOWERING

- Described by RAOULT'S LAW



P_A = partial pressure of the VAPOR of solvent molecules.

$$\underline{P_A} = \underline{P_A^*} \times \underline{X_A}$$

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...

$$\underline{\Delta P} = P_A^* \times \underline{X_B}$$

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)

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?

$T_b = 118.5°C$ $K_b = 3.08 °C/m$

$\Delta T_b = K_b \times C_m$ | $C_m = \dfrac{\text{moles } S_8}{\text{kg } HC_2H_3O_2}$ ← 0.1000 kg

First, find the moles of sulfur.

S_8: $8 \times 32.07 = 256.56$ g/mol

$$2.817 \text{ g } S_8 \times \frac{\text{mol } S_8}{256.56 \text{ g } S_8} = 0.0109799 \text{ mol } S_8$$

Find Cm.

$$C_m = \frac{0.0109799 \text{ mol } S_8}{0.1000 \text{ kg}} = 0.109799 \, m$$

Once you know Cm, you can calculate the boiling point ELEVATION

$$\Delta T_b = (3.08 °C/m)(0.109799 \, m) = 0.338 °C$$

Calculate the new boiling point by adding the boiling point elevation to the original boiling point

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

← It didn't change MUCH in this example!