SOLUTIONS

- a SOLUTION is a HOMOGENEOUS MIXTURE.

  Uniform properties throughout!

- parts of a solution:

  1. SOLUTE(S)
     - component(s) of a solution present in small amounts.
  2. SOLVENT
     - the component of a solution present in the GREATEST amount
     - in solutions involving a solid or gas mixed with a LIQUID, the liquid is typically considered the solvent.
     - solutions are usually the same phase as the pure solvent. For example, at room temperature salt water is a liquid similar to pure water.

SOLVENTS

- We traditionally think of solutions as involving gases or solids dissolved in liquid solvents. But ANY of the three phases may act as a solvent!

(1) GAS SOLVENTS

- Gases are MISCIBLE, meaning that they will mix together in any proportion.
- This makes sense, since under moderate conditions the molecules of a gas don't interact wth each other.
- Gas solvents will only dissolve other gases.

(2) LIQUID SOLVENTS

- Can dissolve solutes that are in any phase: gas, liquid, or solid.
- Whether a potential solute will dissolve in a liquid depends on how compatible the forces are between the liquid solvent and the solute.

(3) SOLID SOLVENTS

- Solids can dissolve other solids, and occasionally - liquids.
- Solid-solid solutions are called ALLOYS. Brass (15% zinc dissolved in copper) is a good example.
- AMALGAM is a solution resulting from dissolving mercury into another metal.

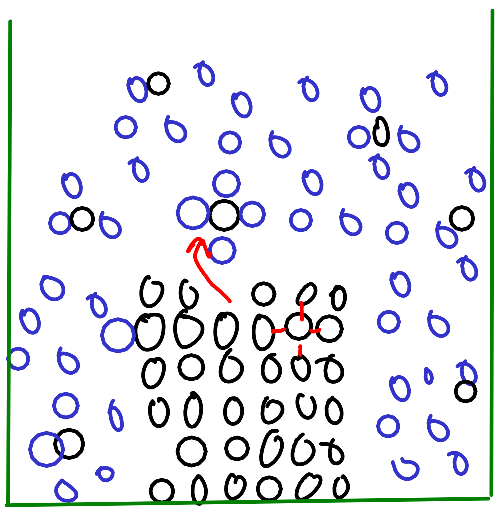
HOW THINGS DISSOLVE

- 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)

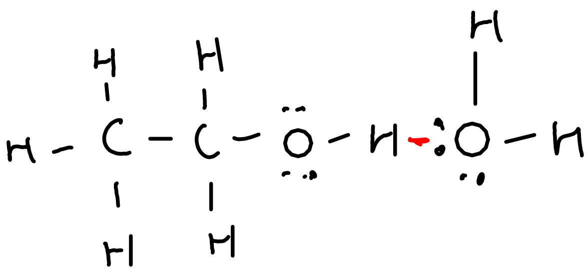
- "like dissolves like": Substances held together by similar (or at least compatible) kinds of attractive forces can dissolve in each other. Substances that are held together by very different kinds of attractive forces will not dissolve in one another!

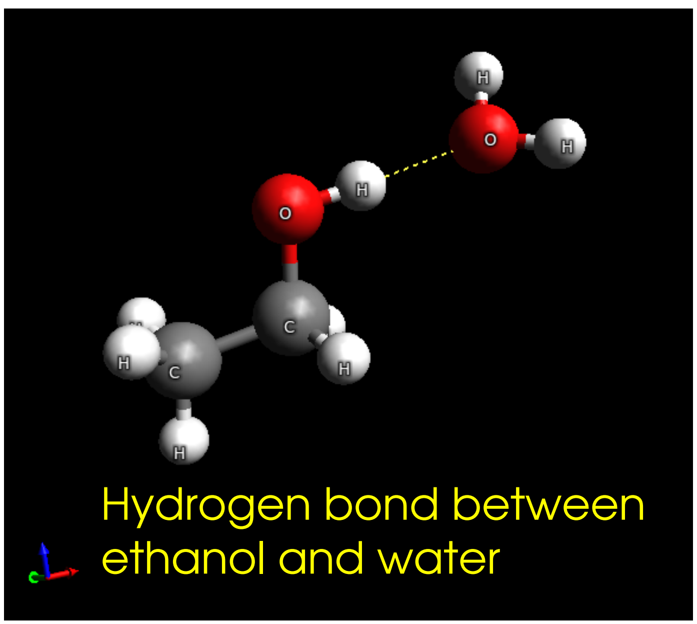
Consider WATER:

HYDROGEN BONDS

↓

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





Hydrogen bond between ethanol and water

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!

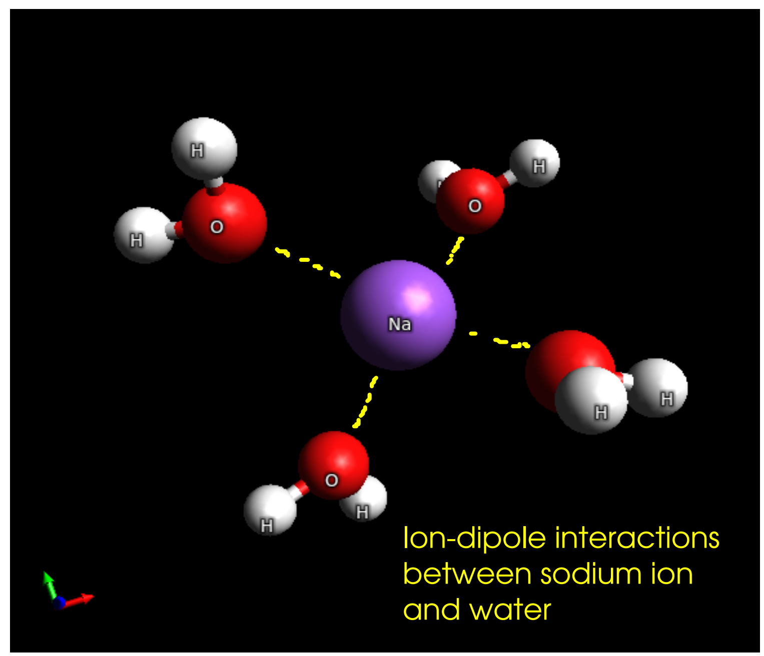
$Na^+ - Cl^-$

SMALL (little London force)

↓

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

(example: OILS and WAXES)



Ion-dipole interactions between sodium ion and water

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

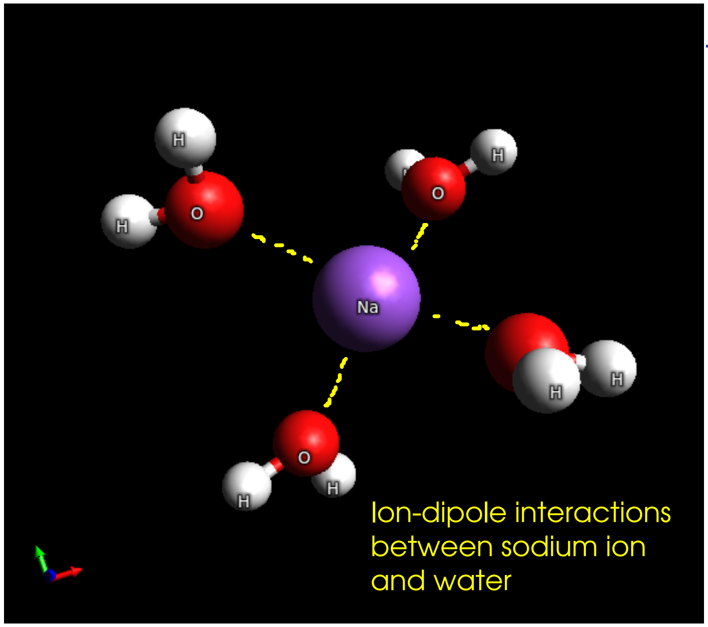
- MOLECULAR SOLIDS will dissolve well in liquids if they are held together by similar forces.

- 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. $Cl^- - Na^+ - Cl^-$

$Na^+ - Cl^- - Na^+$

- 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 hydration are competing trends!

EXTERNAL FACTORS AFFECTING SOLUBILITY

- There are a few external factors that affect the solubility. (By external, we mean other than the chemical identity of the solute and solvent).

(1) TEMPERATURE

-For gases dissolved in liquids, the solubility DECREASES as the temperature INCREASES

- This is why THERMAL POLLUTION is bad! Hot water holds less oxygen than cooler water.

- For solids dissolved in liquids, solubility USUALLY increases with temperature. This is not true for ALL solid/liquid solutions.

(2) PRESSURE

- For gases dissolved in liquids, solubility INCREASES when the partial pressure of the solute gas over the solution INCREASES.

- Consider soft drinks. They go flat after opening because the pressure of carbon dioxide over the liquid goes down.

- No significant pressure effects for solid/liquid solutions.

CONCENTRATION

- When you discuss a solution, you need to be aware of:
    - what materials are in the solution
    - how much of each material is in the solution

- CONCENTRATION is the amount of one substance compared to the others in a solution. This sounds vague, but that's because there are many different ways to specify concentration!

- We will discuss four different concentration units in CHM 111:

(1) MASS PERCENTAGE

$$= \frac{\text{mass solute}}{\text{mass solution}} \times 100\% \qquad \%\,,\ \%\ w/w$$

(2) MOLARITY

$$= \frac{\text{moles solute}}{\text{L solution}} \qquad M \text{ or } \underline{M}$$

(3) MOLALITY

$$= \frac{\text{moles solute}}{\text{kg solvent}} \qquad m$$

(4) MOLE FRACTION

$$= \frac{\text{moles component A}}{\text{moles solution}} \qquad X_A$$

How would you prepare 455 grams of an aqueous solution that is 6.50% sodium sulfate by mass?

$$\text{mass \%} = \frac{\text{mass } Na_2SO_4}{\text{mass solution}} \times 100$$

Start by writing the definition of the unit(s) you're using. Try to fill in the terms you know!

We know everything in the definition except the mass of sodium sulfate ... So let's solve for that!

$$6.50 = \frac{\text{mass } Na_2SO_4}{455\,g} \times 100$$

$$29.6\,g = \text{mass } Na_2SO_4$$

Calculate the mass of water by subtraction...

$$455\,g \text{ solution} - 29.6\,g\ Na_2SO_4 = 425.4\,g \text{ water}$$

To prepare the solution, mix 29.6 grams sodium sulfate with 425.4 grams water!

What's the MOLALITY and MOLE FRACTION OF SOLUTE of the previous solution?

29.6 g Na_2SO_4, 425.4 g water ← previous solution

To find molality, write out the definition of the unit first...

$$\frac{\text{mol } Na_2SO_4 \; (1)}{\text{Kg } H_2O \; (2)}$$

molality

1 - Convert 29.6 grams of sodium sulfate to moles. Use FORMULA WEIGHT.

2 - Change 425.4 grams water to kilograms.

Find formula weight of sodium sulfate: Na_2SO_4 - Na: 2 x 22.99

S: 1 x 32.07

O: 4 x 16.00

142.05 g Na_2SO_4 = mol Na_2SO_4

$$(1)\ 29.6\,g\,Na_2SO_4 \times \frac{\text{mol } Na_2SO_4}{142.05\,g\,Na_2SO_4} = 0.2083773314 \text{ mol } Na_2SO_4$$

Convert mass water to kg. $Kg = 10^3\,g$

$$(2)\ 425.4\,g\,H_2O \times \frac{Kg}{10^3\,g} = 0.4254\,Kg\,H_2O$$

So the molality is:

$$\frac{\text{mol } Na_2SO_4}{Kg\,H_2O} = \frac{0.2083773314 \text{ mol } Na_2SO_4}{0.4254\,Kg\,H_2O} = \boxed{0.490\,m\,Na_2SO_4}$$

29.6 g Na_2SO_4, 425.4 g water ← previous solution

Write the definition of the mole fraction ...

$$X_{Na_2SO_4} = \frac{\text{mol } Na_2SO_4}{\text{mol solution}} = \frac{\text{mol } Na_2SO_4 \quad ①}{\text{mol } Na_2SO_4 + \text{mol } H_2O \quad ②}$$

1 - We've already calculated moles sodium sulfate (from the 29.6 g), so we'll just re-use that number.

2 - We can find moles water by converting 425.5 grams to moles.

① 0.2083773314 mol Na_2SO_4

② H_2O: H − 2 × 1.008
O − 1 × 16.00
18.016 g H_2O = mol H_2O

$$425.4 \text{ g } H_2O \times \frac{\text{mol } H_2O}{18.016 \text{ g } H_2O} = 23.61234458 \text{ mol } H_2O$$

Find mole fraction...

$$\frac{\text{mol } Na_2SO_4}{\text{mol } Na_2SO_4 + \text{mol } H_2O} = \frac{0.2083773314 \text{ mol } Na_2SO_4}{0.2083773314 \text{ mol } Na_2SO_4 + 23.61234458 \text{ mol } H_2O}$$

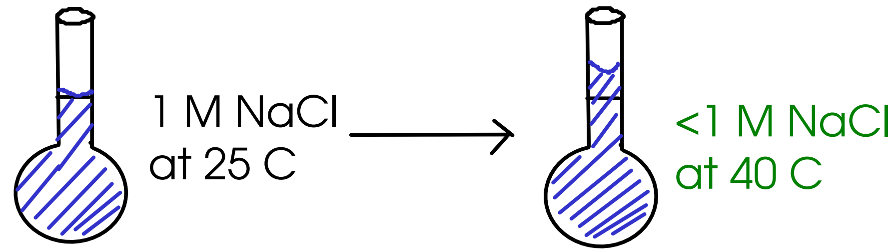
$$= \boxed{0.00875 = X_{Na_2SO_4}}$$

MOLARITY

- In the previous example, we converted between three of the four units that we discussed: mass percent, molality, and mole fraction.

- We didn't do MOLARITY, because the information given in the previous problem was not sufficient to determine molarity!

$$M = \frac{\text{moles solute}}{\text{L solution}}$$



Molarity is based on VOLUME, while the other three units are based on MASS. (moles and mass can be directly converted)

Volume depends on TEMPERATURE!

- If you HEAT a solution, what happens to CONCENTRATION?

ex: 5.00 mol Na_2SO_4 in 1 L solution

5.00 mol Na_2SO_4 — constant when heated

1 L solution — increases, (thermal expansion)

... the MOLAR CONCENTRATION decreases. (But the concentration in the other three units we discussed stays the same.)

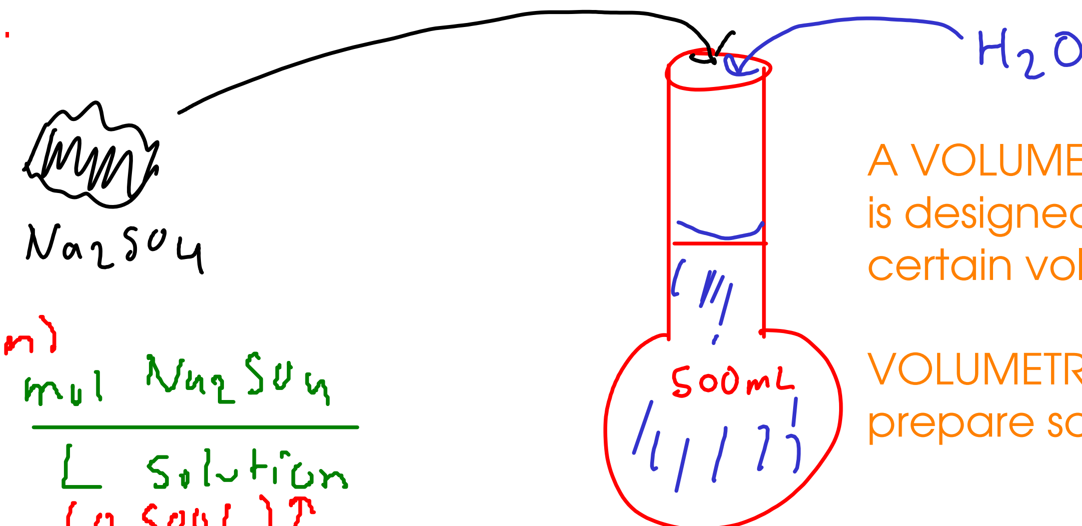
- If you COOL a solution, the MOLAR CONCENTRATION increases. (The other three units stay the same!)

... 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.

Na_2SO_4: (142.05 g/mol)

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

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



volumetric flask

$$M = \frac{\text{mol } Na_2SO_4}{\text{L solution}}$$

(0.500 M) → M; (0.500 L) → L 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.

* 500 mL = 0.500 L

From the definition of molarity, we see that we can calculate the moles of sodium sulfate required (since we already know everything else in the definition.

$$0.500\ M = \frac{\text{mol } Na_2SO_4}{.500\ L} \; ; \; 0.250 \text{ mol } Na_2SO_4 \text{ needed.}$$

Use the formula weight of sodium sulfate to convert moles to mass...

$$0.250 \text{ mol } Na_2SO_4 \times \frac{142.05\ g\ Na_2SO_4}{\text{mol } Na_2SO_4} = 35.5\ g\ Na_2SO_4$$

So, weigh out 35.5 grams sodium sulfate into a 500 mL volumetric flask, then add water to the mark.

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)

$$M_1 V_1 = M_2 V_2$$

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 ← (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?

$M_1V_1 = M_2V_2$

$M_1 = 0.500\,M \qquad M_2 = 0.333\,M$

$V_1 = ? \qquad V_2 = 150.\,mL$

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

$V_1 = 99.9\,mL$ of $0.500\,M\ Na_2SO_4$

Measure out 99.9 mL of 0.500 M sodium sulfate, then add water to get a total volume of 150 mL (can be done quickly in a large graduated cylinder!)

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.