

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

"stock solutions"

- Use DILUTION EQUATION

The dilution equation is easy to derive with simple algebra.

$$M \times V$$

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

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

$$M_1 V_1 = M_2 V_2$$

before  
dilution

after  
dilution

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

$$M_1 V_1 = M_2 V_2 \quad \dots \text{the "DILUTION EQUATION"}$$

$M_1$  = 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?

$$M_1 V_1 = M_2 V_2 \quad \left| \quad \begin{array}{ll} M_1 = 0.500 \text{ M} & M_2 = 0.333 \text{ M} \\ V_1 = ??? & V_2 = 150. \text{ mL} \end{array} \right.$$

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$$(0.500 \text{ M}) V_1 = (0.333 \text{ M})(150. \text{ mL})$$

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

Take 99.9 mL of 0.500 M stock solution, and add water until the total volume of the mixture is 150. mL

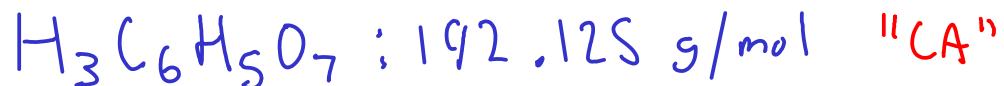
## <sup>62</sup> MOLARITY and the other concentration units

- To convert between molarity and the other three concentration units we've studied, you have to know more about the solution. For example:

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

- \* 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.

<sup>63</sup> 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



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

molality (m)  molarity (M)

- 1 - Assume we have 1 kg of solvent. This makes the number of moles of CA = 0.688 mol
- 2 - We need to FIND the VOLUME of the solution. We can use density to convert mass to volume, but we have to be careful. The mass we know right now is not the mass of the SOLUTION; it's the mass of the SOLVENT. We must find the MASS OF SOLUTE to find the mass of the solution.

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

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

Now find the volume (using density)

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

Molarity = moles / L solution

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

64 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.

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$$\frac{\text{molality}}{\text{mass percent}} \quad \frac{8.50 \text{ g NH}_4\text{Cl}}{100 \text{ g solution}} \longrightarrow \frac{\text{mol NH}_4\text{Cl}}{\text{kg H}_2\text{O}} \quad \text{molality}$$

Find moles of ammonium chloride using formula weight:

$$8.50 \text{ g NH}_4\text{Cl} \times \frac{\text{mol NH}_4\text{Cl}}{53.491 \text{ g NH}_4\text{Cl}} = 0.15891 \text{ mol NH}_4\text{Cl}$$

Find mass of water by subtracting out the solute mass!

$$100 \text{ g solution} - 8.50 \text{ g NH}_4\text{Cl} = 91.50 \text{ g H}_2\text{O} = 0.09150 \text{ kg H}_2\text{O}$$

Find molality:

$$m_{\text{NH}_4\text{Cl}} = \frac{0.15891 \text{ mol NH}_4\text{Cl}}{0.09150 \text{ kg H}_2\text{O}} = \boxed{1.74 \text{ m NH}_4\text{Cl}}$$

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$$\frac{\text{mole fraction}}{\text{mass percent}} \quad \frac{8.50 \text{ g NH}_4\text{Cl}}{100 \text{ g solution}} \longrightarrow \frac{\text{mol NH}_4\text{Cl}}{\text{mol NH}_4\text{Cl} + \text{mol H}_2\text{O}} \quad \text{mole fraction of ammonium chloride}$$

65 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.

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Find moles of water. Take the mass of water (found before) and convert it to moles using the formula weight:

$$91.50 \text{ g H}_2\text{O} \times \frac{\text{mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 5.07881 \text{ mol H}_2\text{O}$$

$$X_{\text{NH}_4\text{Cl}} = \frac{0.15891 \text{ mol NH}_4\text{Cl}}{0.15891 \text{ mol NH}_4\text{Cl} + 5.07881 \text{ mol H}_2\text{O}} = \boxed{0.0303 = X_{\text{NH}_4\text{Cl}}}$$

... if we need mole fraction of water, it equals 1-(mole fraction of ammonium chloride)

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$$\begin{array}{ccc} \text{molarity} & & \\ \hline & \frac{8.50 \text{ g NH}_4\text{Cl}}{100 \text{ g solution}} & \longrightarrow & \frac{\text{mol NH}_4\text{Cl}}{\text{L solution}} \\ & \text{mass percent} & & \text{molarity} \end{array}$$

Find the volume of 100 grams of SOLUTION using the density:

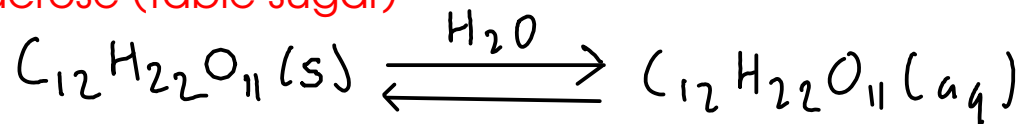
$$100 \text{ g solution} \times \frac{\text{mL}}{1.024 \text{ g}} \times \frac{10^{-3} \text{ L}}{\text{mL}} = 0.09765625 \text{ L}$$

$$\underline{M} = \frac{0.15891 \text{ mol NH}_4\text{Cl}}{0.09765625 \text{ L}} = \boxed{1.63 \underline{M} \text{ NH}_4\text{Cl}}$$

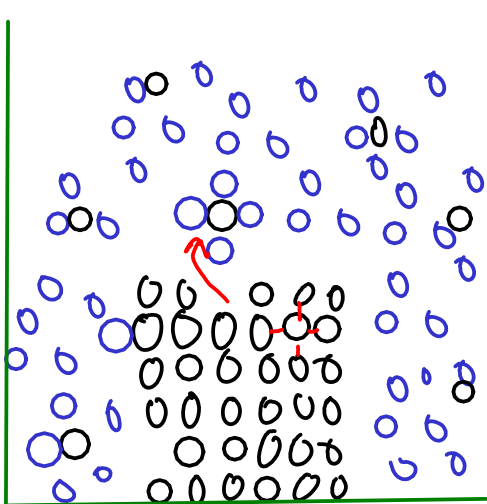
## 66 HOW THINGS DISSOLVE

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

sucrose (table sugar)



... 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)

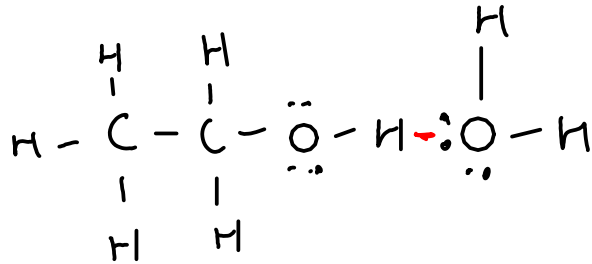
- "like dissolves like": Substances held together by similar (or at least compatible) kinds of attractive forces can dissolve in each other. Substances that are held together by very different kinds of attractive forces will not dissolve in one another!

## Consider WATER:

HYDROGEN BONDS



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



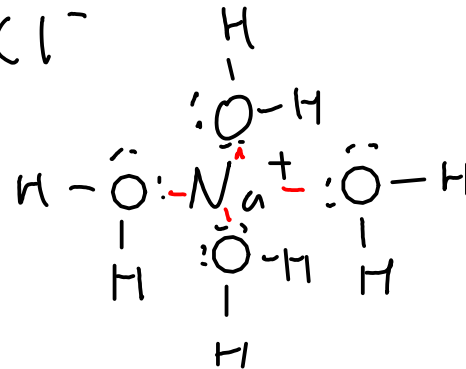
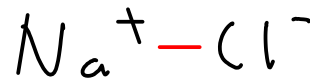
POLAR



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!



"ion-dipole" interactions

SMALL (little London force)



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

(example: OILS and WAXES)



- 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

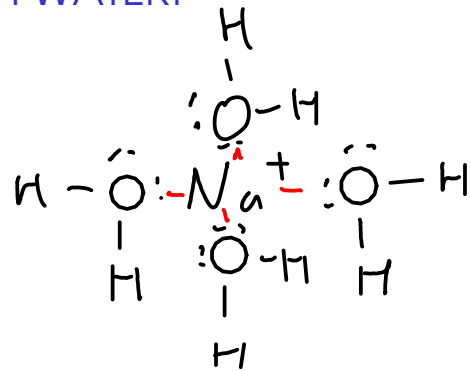
③ - Solids and liquids

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

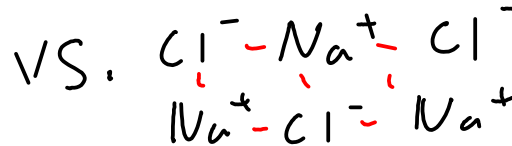
- IONIC SOLIDS will sometimes dissolve in POLAR liquids, but not in nonpolar liquids

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

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



The charged ends of the water molecule HYDRATE the ions.



- The solubility of an ionic compound depends on whether HYDRATION (attraction of water molecules for an ion) is greater than LATTICE ENERGY - the attraction 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!

70 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 solvent 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

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

└ 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?

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

$$\Delta T_f = K_f \times C_m = \frac{\text{moles unknown}}{\text{kg benzene}}$$

$\downarrow$  5.065  $^\circ\text{C}/m$  constant  
 $\uparrow$  5.455  $^\circ\text{C}$  - 4.880  $^\circ\text{C}$  = 0.575  $^\circ\text{C}$

We can calculate the MOLAL concentration based on the freezing point depression.

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

$$C_m = 0.1135241856 \text{ m} = \frac{\text{moles unknown}}{\text{kg benzene}}$$

So, we can find moles of unknown from molal concentration, since we know the amount of benzene used! (100.0g = 0.1000kg)

$$0.1135241856 \text{ m} = \frac{\text{moles unknown}}{0.1000 \text{ kg}} \quad \dots \text{ solve for moles unknown!}$$

$$0.0113524186 \text{ mol} = \text{moles unknown}$$

Molecular weight is mass per mole, so

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