SOLIDS

- RIGID, DENSE, and INCOMPRESSIBLE

    - Properties of interest:

        1. MELTING POINT

            - Temperature at which the bulk phase change from solid to liquid occurs

        2. HARDNESS and BRITTLENESS

            - hardness: resistance of a solid to deformation (shape change) caused by the application of a force

            - brittleness: tendency of a material to fracture or break rather than to deform.

        3. CONDUCTIVITY

            - ability of a material to conduct an electric current

... these properties will be influenced by the KINDS OF FORCES holding the solid together!

CLASSIFICATION OF SOLIDS: By attractive forces

- Solids may be classified either by the type of forces holding the solid together or by structure. We'll discuss forces first.

- Some solids are held together by the same sorts of forces found in liquids. But there are moe options for solids!

- There are four kinds of solids when classified by forces.

1 MOLECULAR SOLIDS

- held together by the same kinds of forces that hold liquids together:

  A van der Waals forces: London dispersion forces and dipole-dipole interactions

  B hydrogen bonds

... generally, these forces are the weakest.

Examples: candle wax, water ice

Generally, molecular solids:
- have LOW MELTING POINTS
- are SOFT
- are NONCONDUCTORS

2 METALLIC SOLIDS

- held together by METALLIC BONDS, which involve electron sharing throughout the body of the metal..

... strength of these metallic bonds is variable.

Examples: iron, gold, copper, zinc, other metals

Generally, metallic solids:
- have a wide range of MELTING POINTS, though almost all melt above room temperature.
- range from SOFT to HARD. Many are MALLEABLE, meaning they deform before breaking.
- are good CONDUCTORS of both heat and electricity

3 IONIC SOLIDS

- held together by IONIC BONDS:

... generally, these forces are much stronger than the ones in molecular solids.

Examples: sodium chloride, any ionic compound

Generally, ionic solids:
- have HIGH MELTING POINTS, well over room temperature
- are HARD
- are NONCONDUCTORS of electricity in the solid phase,
but CONDUCT when melted or dissolved into a liquid solution.

4 COVALENT NETWORK SOLIDS

- held together by COVALENT BONDS.
- are, in essence, giant molecules where the entire solid (not simply individual molecules WITHIN the solid) are held together by covalent bonds.

... these are the strongest kind of forces holding solids together.

Example: diamond

Generally, covalent network solids:
- have EXTREMELY HIGH MELTING POINTS. Most don't melt at all - they decompose before melting.
- are EXTREMELY HARD. The hardest materials known are covalent network solids.
- are NONCONDUCTORS

Relative strengths of the forces holding solids together:

INTERMOLECULAR FORCES	$<$	IONIC BONDS	$<$	COVALENT BONDS
molecular solids		ionic solids		covalent network solids

... the stronger the forces, the:
- HARDER a material
- HIGHER the melting point of the material

Metallic bonds vary considerably, so they have been left out of the comparison!

CLASSIFICATION OF SOLIDS: By structure

- Solids may also be classified by structure. A more in-depth look at solids is something you would find in a materials science class, but we'll discuss two broad categories of solid materials.

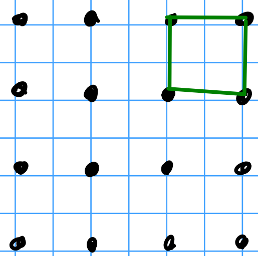
1. AMORPHOUS SOLIDS

   - have a disordered structure at the microscopic level.
   - a very small amount of solids are completely amorphous, but quite a few plastics are at least partially amorphous.

2. CRYSTALLINE SOLIDS

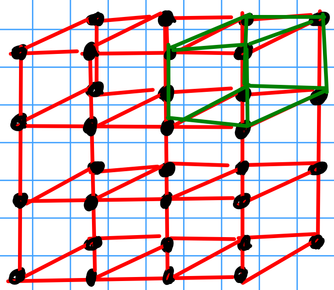
   - have a well-defined three dimensional structure at the microscopic level.
   - structure is made up of a regular, repeating arrangement of points in space - a CRYSTAL LATTICE

MORE ON CRYSTALS



Here's a simple CRYSTAL LATTICE in 2D. The points represent atoms occupying LATTICS POINTS

The simplest repeating pattern that describes the entire crystal is called the UNIT CELL. It's outlined in GREEN here.



Here's a crystal lattice in three dimensions. This one is called a SIMPLE CUBIC lattica. This simple structure can be found in some solid metals like polonium. A polonium atom occupies each lattice point.

The unit cell, again, is highlighted in GREEN.

CRYSTAL DEFECTS

- Natural crystals almost always have some DEFECTS in their structure.

  - Holes in the crystal lattice, where an atom should be but isn't
  - Misaligned planes in the crystal
  - Substitutions of one atom for another in the crystal lattice

- Often defects are undesirable, but not always:

Alumina: Al_2O_3

- clear / white in color
- usually used as the "grit" in cleaners like Comet and Soft Scrub!

ruby: Al_2O_3 with some Al replaced with Cr

- red in color
- valuable gemstone!

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.

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

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

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

$$= \frac{\text{moles component}}{\text{moles solution}} \qquad X$$

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

need to calculate

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

6.50%

455 g

I know everything here except the mass of the solute. So, I can plug in to the definition of mass percent and solve for the unknown.

$$6.50 = \frac{\text{mass solute}}{455} \times 100$$

$\div 100$, $\times 455$

$$\frac{6.50 \times 455}{100} = \text{mass solute} = \boxed{29.6 \text{ g } Na_2SO_4}$$

mix with... 425.4 g water

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

29.6 g Na_2SO_4, 425.4 g water ← previous solution

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

$$X_{Na_2SO_4} = \frac{\text{moles } Na_2SO_4}{\text{moles total}}$$

1. Convert mass of sodium sulfate to moles! (Use FW)
2. Convert grams of water to kg

1. Convert mass of sodium sulfate to moles! (Use FW)
2. Convert mass water to moles (using FW), then add to moles sodium sulfate to get total

$$29.6\text{ g } Na_2SO_4 \times \frac{\text{mol}}{142.04\text{ g}} = 0.208392 \text{ mol } Na_2SO_4$$

$$425.4\text{ g } H_2O \times \frac{\text{kg}}{10^3\text{ g}} = 0.4254 \text{ kg } H_2O$$

$$425.4\text{ g } H_2O \times \frac{\text{mol}}{18.02\text{ g}} = 23.60710 \text{ mol } H_2O, \text{ total} = 23.815495 \text{ mol}$$

$$m = \frac{0.208392 \text{ mol } Na_2SO_4}{0.4254 \text{ kg } H_2O} = \boxed{0.490\text{ m } Na_2SO_4}$$

$$X_{Na_2SO_4} = \frac{0.208392 \text{ mol } Na_2SO_4}{23.815495 \text{ mol}} = \boxed{0.00875}$$

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!

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

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: $\underbrace{5.00 \text{ mol } Na_2SO_4}_{\text{constant when heated}}$ in $\underbrace{1 \text{ L solution}}_{\text{increases. (thermal expansion)}}$

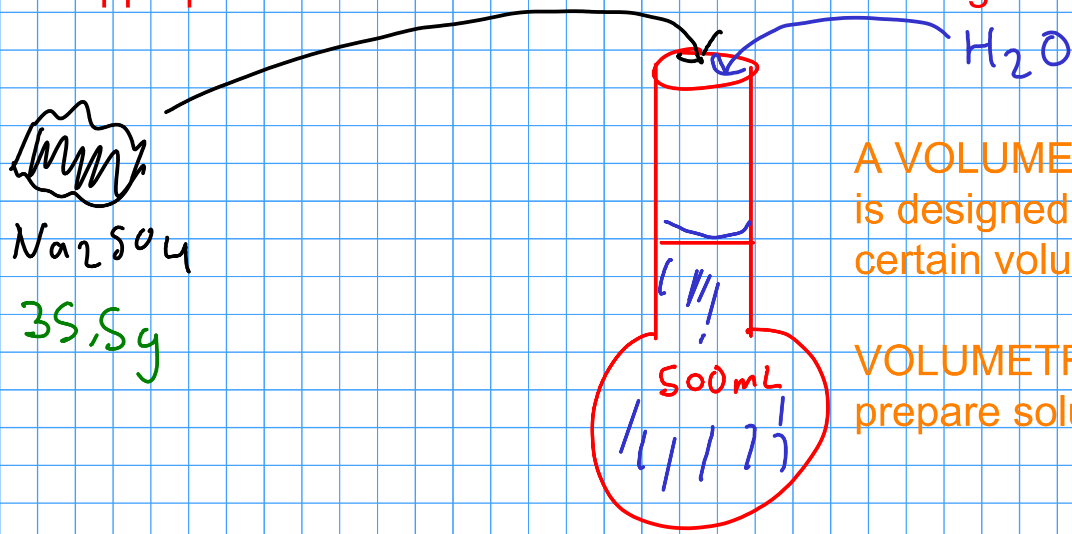
... the MOLAR CONCENTRATION decreases. (But the concentration in the other three units we discussed stays the same.)

- 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.

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.



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 in 500. mL (0.500 L) solution, then convert to grams.

$0.500\ \underline{M}\ Na_2SO_4$ has $0.500\ mol\ Na_2SO_4 = 1\ L$ solution

$$0.500\,L \times \frac{0.500\ mol\ Na_2SO_4}{1\,L} = 0.250\ mol\ Na_2SO_4$$

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

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)

$$M_1 V_1 = M_2 V_2$$

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

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

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?

$M_1V_1 = M_2V_2$

$M_1 = 0.500\ \underline{M}$

$V_1 = ?$

$M_2 = 0.333\ \underline{M}$

$V_2 = 150.\ mL$

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

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