MOLAR CONCENTRATION *

- unit: MOLARITY (M): moles of dissolved substance per LITER of solution

dissolved substance

$$M = \text{molarity} = \frac{\text{moles of SOLUTE}}{\text{L SOLUTION}}$$

6.0 M HCl solution: $\frac{6.0 \text{ mol HCl}}{\text{L}}$

If you have 0.250 L (250 mL) of 6.0 M HCl, how many moles of HCl do you have?

6.0 mol HCl = L

$$0.250 \text{ L} \times \frac{6.0 \text{ mol HCl}}{\text{L}} = 1.5 \text{ mol HCl}$$

*See SECTIONS 4.7 - 4.10 for more information about MOLARITY and solution calculations (p 154 - 162)

If you need 0.657 moles of hydrochloric acid, how many liters of 0.0555 M HCl do you need to measure out?

$0.0555 \text{ mol HCl} = \text{L}$

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{0.0555 \text{ mol HCl}} = \boxed{11.8 \text{ L}}$$

11800 mL

This is too large a volume for typical lab-scale work. To get a more reasonable volume, pick a solution with higher concentration, like ...

What if we used 6.00 M HCl?

$6.00 \text{ mol HCl} = \text{L}$

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} = \boxed{0.110}$$

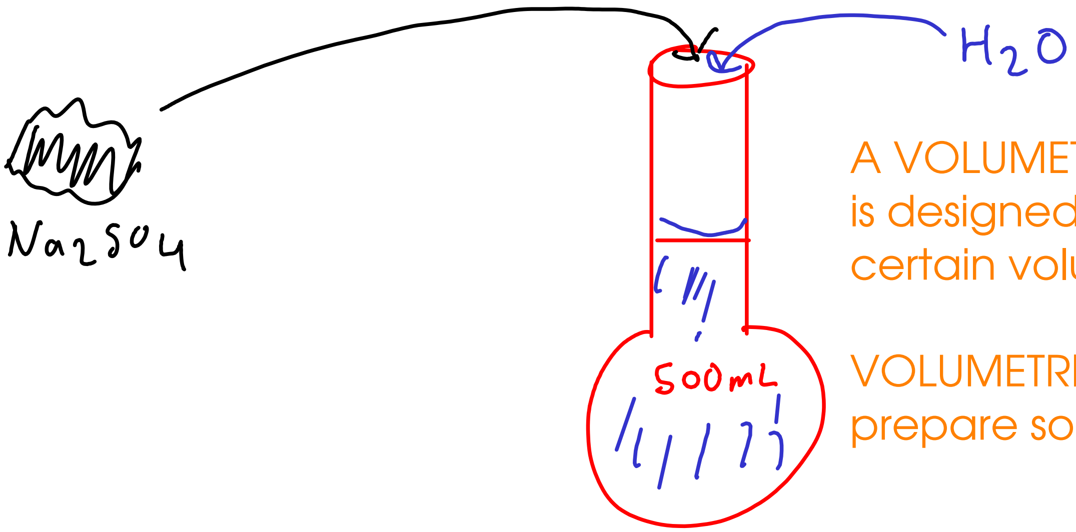
$11\bar{0}$ mL

110 mL is a much more convenient volume, considering our glassware!

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

Na_2SO_4 : 142.05 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

We know we need 500. mL of solution, AND we know the concentration we want is 0.500 M. From that, we can calculate the moles of sodium sulfate, and then convert that to mass to see how much we need to weigh.

$0.500 \text{ mol } Na_2SO_4 = L$ | $mL = 10^{-3} L$ | $142.05 \text{ g } Na_2SO_4 = \text{mol } Na_2SO_4$

$$500. \text{ mL} \times \frac{10^{-3} \text{ L}}{\text{mL}} \times \frac{0.500 \text{ mol } Na_2SO_4}{\text{L}} \times \frac{142.05 \text{ g } Na_2SO_4}{\text{mol } Na_2SO_4} = \boxed{35.5 \text{ g } Na_2SO_4}$$

So, to prepare this solution, we weigh out 35.5 grams of sodium sulfate into a 500 mL volumetric flask, then add water until the water level gets to the fill line.

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)

$$M_1V_1 = M_2V_2$$

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

✻ Remember ... V2 represents the TOTAL VOLUME of solution, not the volume of added water.

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?

$M_1V_1 = M_2V_2$

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

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

So, to make this new solution we'll measure out 99.9 mL of 0.500 M sodium sulfate, then dilute to 150. mL with water. (Ideally, do this in a 150 mL volumetric flask)

CHEMICAL CALCULATIONS CONTINUED: REACTIONS

- Chemical reactions proceed on an ATOMIC basis, NOT a mass basis!

- To calculate with chemical reactions (i.e. use chemical equations), we need everything in terms of ATOMS ... which means MOLES of atoms

$$\underline{2}\,Al(s) + \underline{3}\,Br_2(l) \longrightarrow \underline{2}\,AlBr_3(s)$$

coefficients are in terms of atoms and molecules!

2 atoms Al = 3 molecules Br_2 = 2 formula units $AlBr_3$

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2 mol Al = 3 mol Br_2 = 2 mol $AlBr_3$

- To do chemical calculations, we need to:
  - Relate the amount of substance we know (mass or volume) to a number of moles
  - Relate the moles of one substance to the moles of another using the equation
  - Convert the moles of the new substance to mass or volume as desired

$$\underline{2}Al(s) + \underline{3}Br_2(l) \longrightarrow 2AlBr_3(s)$$

* Given that we have 25.0 g of liquid bromine, how many grams of aluminum would we need to react away all of the bromine? How many grams of aluminum bromide would be produced?

① Convert grams of bromine to moles: Need formula weight $Br_2: 2 \times 79.90$

$\overline{159.80}$

$159.80\,g\,Br_2 = 1\,mol\,Br_2$

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} = 0.15645\,mol\,Br_2$$

② Use the chemical equation to relate moles of bromine to moles of aluminum

$2\,mol\,Al = 3\,mol\,Br_2$

$$0.15645\,mol\,Br_2 \times \frac{2\,mol\,Al}{3\,mol\,Br_2} = 0.10430\,mol\,Al$$

③ Convert moles aluminum to mass: Need formula weight $Al: 26.98$

$26.98\,g\,Al = 1\,mol\,Al$

$$0.10430\,mol\,Al \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$

You can combine all three steps on one line if you like!

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{154.80\,g\,Br_2} \times \frac{2\,mol\,Al}{3\,mol\,Br_2} \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$

① ② ③

You can solve the second part of the question using CONSERVATION OF MASS - since there's only a single product and you already know the mass of all reactants.

$$\begin{array}{rr} & 25.0\,g\,Br_2 \\ + & 2.81\,g\,Al \\ \hline & 27.8\,g\,AlBr_3 \end{array}$$

But ...

...what would you have done to calculate the mass of aluminum bromide IF you had NOT been asked to calculate the mass of aluminum FIRST?

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{154.80\,g\,Br_2} \times \frac{2\,mol\,AlBr_3}{3\,mol\,Br_2} \times \frac{266.694\,g\,AlBr_3}{1\,mol\,AlBr_3} = 27.8\,g\,AlBr_3$$

① convert mass bromine to moles

② convert moles bromine to moles aluminum bromide

③ convert moles aluminum bromide to mass

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$\underline{2}HCl(aq) + \underline{\ \ }Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

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1 - Convert 25.0 g of sodium carbonate to moles. Use formula weight.
2 - Convert moles sodium carbonate to moles HCl using chemical equation.
3 - Convert moles HCl to volume using molarity.

① Na_2CO_3: Na : 2 x 22.99
C : 1 x 12.01
O : 3 x 16.00

105.99 g Na_2CO_3 = mol Na_2CO_3

Formula weight of sodium carbonate

$$25.0\text{ g } Na_2CO_3 \times \frac{\text{mol } Na_2CO_3}{105.99\text{ g } Na_2CO_3} = 0.2358713086\text{ mol } Na_2CO_3$$

② 2 mol HCl = mol Na_2CO_3

$$0.2358713086\text{ mol } Na_2CO_3 \times \frac{2\text{ mol HCl}}{\text{mol } Na_2CO_3} = 0.4717426172\text{ mol HCl}$$

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

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1 - Convert 25.0 g of sodium carbonate to moles. Use formula weight.
2 - Convert moles sodium carbonate to moles HCl using chemical equation.
3 - Convert moles HCl to volume using molarity.

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③ 6.00 mol HCl = L | mL = 10^{-3} L

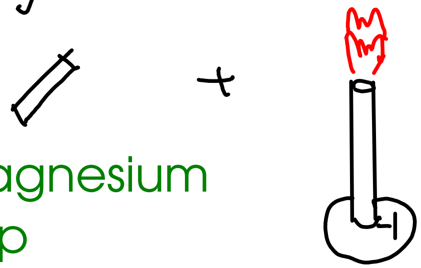
$$0.4717426172 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} \times \underbrace{\frac{\text{mL}}{10^{-3}\text{ L}}} = \boxed{78.6 \text{ mL of } 6.00 \text{ M HCl}}$$

The problem asks us for an answer in mL, so we just tacked the unit conversion on to the end of the calculation!

CONCEPT OF LIMITING REACTANT

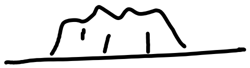
- When does a chemical reaction STOP?

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$$



Magnesium strip

Flame, oxygen from air



Magnesium oxide powder

- When does this reaction stop? When burned in open air, this reaction stops when all the MAGNESIUM STRIP is gone. We say that the magnesium is LIMITING.

- This reaction is controlled by the amount of available magnesium

- At the end of a chemical reaction, the LIMITING REACTANT will be completely consumed, but there may be amount of OTHER reactants remaining. We do chemical calculations in part to minimize these "leftovers".

These are often called "excess" reactants, or reactants present "in excess"

LIMITING REACTANT CALCULATIONS

- To find the limiting reactant, calculate how much product would be produced from ALL given reactants. Whichever produces the SMALLEST amouht of product is the limiting reactant, and the smallest anount of product is the actual amount of product produced.

Example:

$$\overset{56.08}{\underline{CaO}(s)} + 3\,\overset{12.01}{\underline{C}(s)} \xrightarrow{\Delta} \overset{64.10}{CaC_2(s)} + CO(g)$$

<- Formula weights

If you start with 100. g of each reactant, how much calcium carbide would be produced?

CaO: $56.08\,g\ CaO = mol\ CaO$ | $mol\ CaO = mol\ CaC_2$ | $64.10\,g\ CaC_2 = mol\ CaC_2$

$$100.\,g\ CaO \times \frac{mol\ CaO}{56.08\,g\ CaO} \times \frac{mol\ CaC_2}{mol\ CaO} \times \frac{64.10\,g\ CaC_2}{mol\ CaC_2} = \boxed{114\,g\ CaC_2}$$

C: $12.01\,g\ C = mol\ C$ | $3\ mol\ C = mol\ CaC_2$ | $64.10\,g\ CaC_2 = mol\ CaC_2$

$$100.\,g\ C \times \frac{mol\ C}{12.01\,g\ C} \times \frac{mol\ CaC_2}{3\ mol\ C} \times \frac{64.10\,g\ CaC_2}{mol\ CaC_2} = 178\,g\ CaC_2$$

114 g of calcium carbide should be produced. Caclium oxide runs out after we've made 114 g of calcium carbide, so there's nothing left for the remaining carbon to react with.

We say that calcium oxide is LIMITING, and carbon is present IN EXCESS.

PERCENT YIELD

- Chemical reactions do not always go to completion! Things may happen that prevent the conversion of reactants to the desired/expected product!

① SIDE REACTIONS:

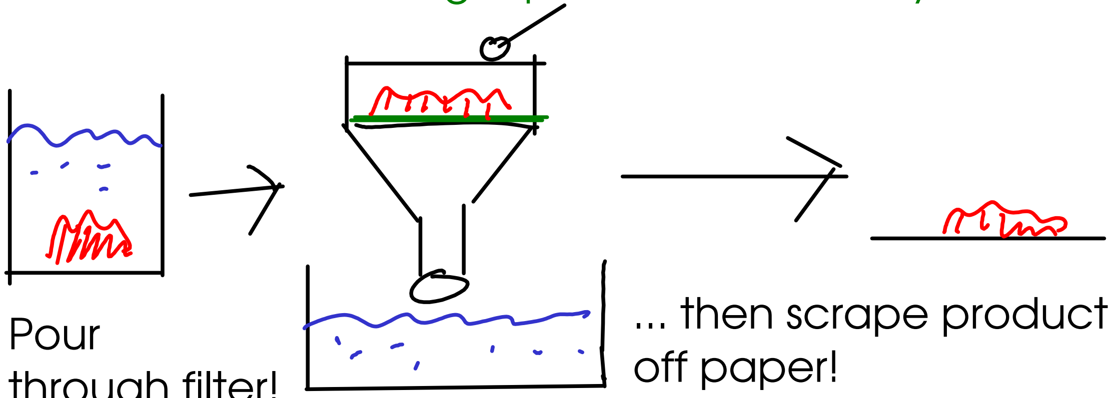
$C + O_2 \longrightarrow CO_2$ | This reaction occurs when there is a large amount of oxygen available

$2C + O_2 \longrightarrow 2CO$ | ... while this reaction is more favorable in low-oxygen environments!

... so in a low-oxygen environment, you may produce less carbon dioxide than expected!

② TRANSFER AND OTHER LOSSES

- When isolating a product, losses may occur in the process. Example: filtering



Pour through filter!

... then scrape product off paper!

During each step of this process, some amount of product loss will occur!

③ EQUILIBRIUM

- Reactions may reach an equilbrium between prodcuts and reactants. We'll talk more about this in CHM 111. The net results is that the reaction will appear to stop before all reactants have been consumed!

- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

Determined EXPERIMENTALLY!

$$\text{PERCENT YIELD} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100\,\%$$

Calculated based on the limiting reactant. (The chemical calculations you've done up to now have been theoretical yields!)

... the percent yield of a reaction can never be greater than 100% due to conservation of mass! If you determine that a percent yield is greater than 100%, then you've made a mistake somewhere - either in a calculation or in the experiment itself!