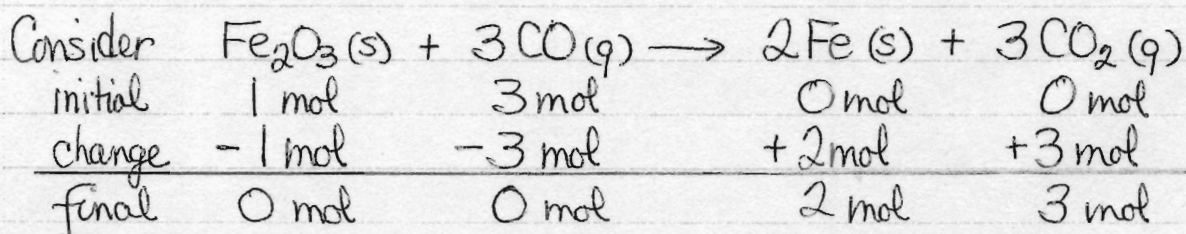


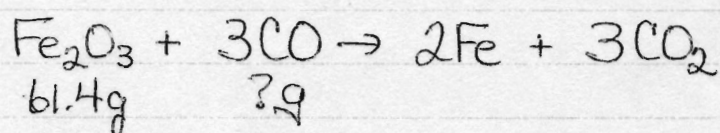
Chapter 4: Stoichiometry: Quantitative Information About Chemical Reactions

A balanced reaction shows the quantitative relationships between reactants and products in a chemical reaction.



Note: (1) the stoich. coefficients do not tell us how much we have.
 (2) they tell us the ratios in which reactants are lost and products are formed (more later)

Example: What mass of CO is required to react with 61.4 g of Fe_2O_3 ?



plan of action: $\text{g Fe}_2\text{O}_3 \xrightarrow{\text{FW}} \text{moles Fe}_2\text{O}_3 \xrightarrow{\text{coeff.}} \text{moles CO} \xrightarrow{\text{FW}} \text{g CO}$

$$\textcircled{1} ? \text{ mol Fe}_2\text{O}_3 = \frac{\text{g Fe}_2\text{O}_3}{\text{FW}} = \frac{61.4\text{g}}{159.7\text{g/mol}} = 0.384\text{ mol}$$

$$\textcircled{2} ? \text{ mol CO} = 0.384\text{ mol Fe}_2\text{O}_3 \times \frac{3\text{ mol CO}}{1\text{ mol Fe}_2\text{O}_3} = 1.15\text{ mol}$$

$$\textcircled{3} ? \text{ g CO} = \text{FW} \times 1.15\text{ mol CO} = 32.2\text{ g CO}$$

HINT: don't round until the end!

FW = $\frac{\#g}{\#\text{mol}}$
FW Fe_2O_3
$2 \times 55.85 = 111.70$
$3 \times 16.00 = 48.00$
$\hline 159.7\text{ g/mol}$
FW CO
$12.0 + 16.0 = 28.0\text{ g/mol}$
NOT $3 \times 28 = 84$

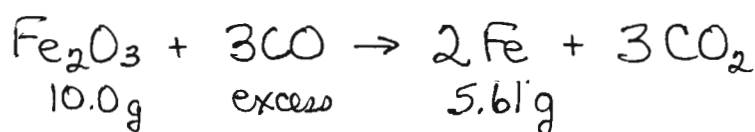
By dimensional analysis: STEP 1

$$\begin{aligned} ? \text{ g CO} &= 61.4\text{ g Fe}_2\text{O}_3 \times \left(\frac{1\text{ mol Fe}_2\text{O}_3}{159.7\text{ g Fe}_2\text{O}_3} \right) \times \left(\frac{3\text{ mol CO}}{1\text{ mol Fe}_2\text{O}_3} \right) \times \left(\frac{28.0\text{ g CO}}{1\text{ mol CO}} \right) \\ &= 32.2\text{ g} \end{aligned}$$

FW
stoich. coeff
relationship

% yield problem types: (1) calculate % yield
 (2) calculate actual yield, given % yield & reactant
 more difficult (3) calculate amount of reactant necessary to
 give a certain amt. of product knowing % yield.

Example: Calculate the % yield of Fe if a 10.0 g sample of Fe_2O_3 reacted with excess CO to produce 5.61 g of Fe.



We know $\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$

(1) Find theoretical yield of Fe (do straightforward stoich. problem)

Plan: g $\text{Fe}_2\text{O}_3 \rightarrow$ mol $\text{Fe}_2\text{O}_3 \rightarrow$ mol Fe \rightarrow g Fe

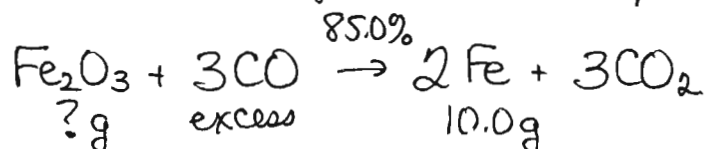
$$\begin{aligned} ? \text{ g Fe} &= 10\text{g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \\ &= 6.99 \text{ g} \end{aligned}$$

But only 5.61 g of Fe were produced!!

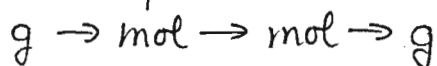
$$\begin{aligned} (2) \quad \% \text{ yield} &= \frac{5.61 \text{ g} \leftarrow \text{what we actually got!}}{6.99 \leftarrow \text{what we should have produced!}} \times 100 \\ &= 80.3\% \end{aligned}$$

So, in this reaction, we expected to make 6.99 g of Fe, but we only produced 80.3% of that amount or 5.61 g of Fe.

Example: Using the same reaction, how many grams of Fe_2O_3 are required to produce 10.0 g of Fe assuming a percent yield of 85.0%?



Plan: (1) Calculate theoretical yield from mass of Fe and % yield
 (2) Using theoretical yield, calculate mass of Fe_2O_3 .



(1) we know $\% \text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100$

So, theoretical yield = $\frac{\text{actual}}{\% \text{yield}} \times 100 = \frac{10.0 \text{ g}}{85.0\%} \times 100 = 11.8 \text{ g Fe}$

(2) ? g $\text{Fe}_2\text{O}_3 = 11.8 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}} \times \frac{159.7 \text{ g}}{1 \text{ mol Fe}_2\text{O}_3}$
 $= 16.9 \text{ g}$

So with this reaction with a yield of 85.0%, you need more Fe_2O_3 than you would think to make 10.0 g of Fe.

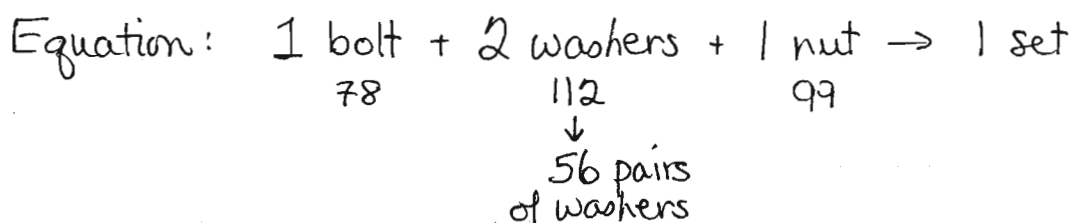
The Limiting Reactant Concept

Up until now, we have had either stoichiometric amounts of reactants or we were told which reactant was in excess?

But how do we decide which, if any, of the reactants is in excess and which reactant is limiting, i.e. controlling how much product we can make?

Example: NON-CHEMICAL

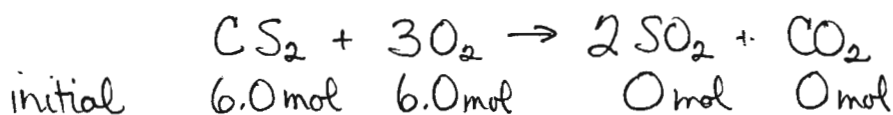
Suppose we have 78 bolts, 112 washers, 99 nuts. How many sets, each consisting of 1 bolt, 2 washers and 1 nut, can we construct?



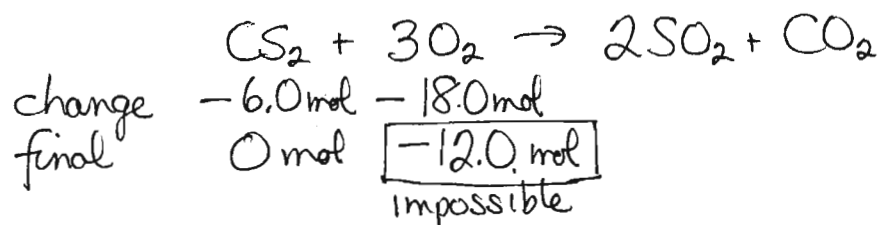
- ∴ we can only put together 56 total sets
- the number of washers limited the number of sets we can make.
- ∴ Washers is the limiting reactant!

Example: CHEMICAL

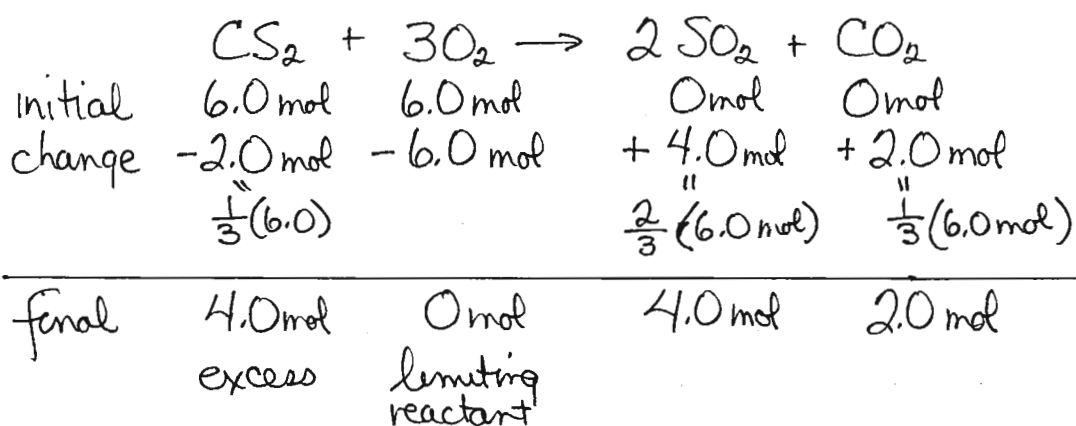
Suppose we have 6.0 mol of CS_2 and 6.0 moles of O_2 in our container. How much SO_2 and CO_2 can we make?



Now you have to make a choice. Is CS_2 going to be used up first or is O_2 ? Let's make an arbitrary decision - CS_2 will be used up first and its number of moles will go to 0.



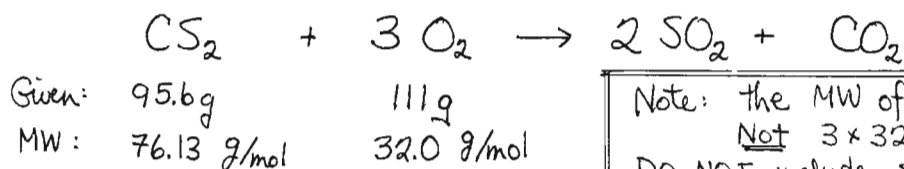
So, CS_2 cannot be the limiting reactant; O_2 must be it!
From the beginning,



It's the change that relates to the stoichiometric coefficient

Now let's do another problem!

Example: What is the maximum mass of sulfur dioxide, SO_2 , that can be produced by the reaction of 95.6 g of carbon disulfide, CS_2 , with 111 g of oxygen, O_2 ?



Note: the MW of O_2 is 32.0 g/mol
 Not $3 \times 32 = 96$.
 DO NOT include stoichiometric coefficients when determining MW.

Step ① decide which reactant is limiting - MUST work in moles!!

∴ determine number of moles of each reactant

$$\text{moles CS}_2 = \frac{g}{\text{MW}} = \frac{95.6g}{76.13 \text{ g/mol}} = 1.26 \text{ mol}$$

$$\text{moles O}_2 = \frac{g}{\text{MW}} = \frac{111g}{32.0 \text{ g/mol}} = 3.47 \text{ mol}$$

Let us assume that CS_2 is limiting (50/50 chance)

∴ this means that CS_2 is all used up in the reaction and O_2 is excess

∴ to use up all CS_2 we need at least $3 \times 1.26 \text{ mol} = 3.78 \text{ moles O}_2$

But we only have 3.47 moles of O_2 ∴ O_2 must be limiting
 (CS_2 is in excess)

✱ Mathematically, therefore, to determine the limiting reagent:

(1) change mass → moles

(2) divide the number of moles by stoichiometric coefficient

$$\text{CS}_2 : 1.26 \div 1 = 1.26$$

$$\text{O}_2 : 3.47 \div 3 = 1.16$$

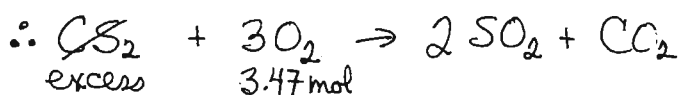
} this is a normalization procedure like nuts + bolts example earlier.

(3) the reactant with the smallest number is

the limiting reagent and controls how much product is formed eg O_2 in this example.

All the other reactants are in excess and can be ignored.

All calculations are based on the amount of the limiting reagent!



Step ② do the problem like any other stoichiometric problem

Plan: moles $\text{O}_2 \rightarrow$ moles $\text{SO}_2 \rightarrow$ g SO_2

$$? \text{ moles SO}_2 = 3.47 \text{ moles O}_2 \times \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} = 2.31 \text{ moles}$$

* Hint: use dimensional analysis to decide if the ratio is $\frac{2}{3}$ or $\frac{3}{2}$

$$? \text{ g SO}_2 = 2.31 \text{ moles SO}_2 \times 64.1 \text{ g/mol} = 148 \text{ g SO}_2$$

Note: What is in the container when the reaction is done? $\text{SO}_2, \text{CO}_2, \text{CS}_2$
How much of the excess reactant remains? in moles? in grams?

$$? \text{ mol CS}_2 \text{ used up} = 3.47 \text{ mol O}_2 \times \frac{1 \text{ mol CS}_2}{3 \text{ mol O}_2} = 1.16 \text{ mol}$$

$$? \text{ mol CS}_2 \text{ remains} = \underset{\text{starting}}{1.26 \text{ mol}} - \underset{\text{ending}}{1.16 \text{ mol}} = 0.10 \text{ mol}$$

$$? \text{ g CS}_2 \text{ remains} = 0.10 \text{ mol} \times 76.13 \text{ g/mol} = 7.6 \text{ g CS}_2$$

Analysis of Mixtures:

A sample of limestone and other soil materials is heated, and the limestone decomposes to give calcium oxide and carbon dioxide



A 1.506 g sample of limestone-containing material gives 0.558 g of CO_2 in addition to CaO after being heated to a high temperature. What is the mass percent of CaCO_3 in the original sample?

Plan: g $\text{CO}_2 \rightarrow$ mol $\text{CO}_2 \rightarrow$ mol $\text{CaCO}_3 \rightarrow$ g $\text{CaCO}_3 \rightarrow$ % CaCO_3 in sample

$$\begin{aligned} ? \text{ g CaCO}_3 \text{ in sample} &= 0.558 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CO}_2} \times \frac{100.1 \text{ g}}{1 \text{ mol CaCO}_3} \\ &= 1.27 \text{ g} \end{aligned}$$

$$\begin{aligned} \therefore \% \text{ CaCO}_3 \text{ in sample} &= \frac{\text{g CaCO}_3}{\text{g sample}} \times 100 \\ \text{by mass} & \\ &= \frac{1.27 \text{ g CaCO}_3}{1.506 \text{ g sample}} \times 100 \\ &= 84.3 \% \text{ by mass} \end{aligned}$$

Concentrations of Solutions

solution: a homogeneous mixture of 2 or more substances

A solution is composed of two parts:

- (1) solute: dissolved substance (substance in lesser amount)
- (2) solvent: dissolving substance (substance in greater amount)

Example: NaCl(aq) solution of NaCl in water, H₂O
↑ solute ↑ solvent

$$\text{Concentration} = \frac{\text{amount of solute}}{\text{mass or volume of solution}}$$

relative terms { dilute solution: small amount of solute in large amount solvent
 { concentrated solution: large amount of solute in solvent

We will discuss 2 concentration units

(1) percent by mass

(do not confuse with % by mass of element in compound)

(2) Molarity

(3) ppm (parts per million) — see lab

∴ lots of concentration units

1. Percent by mass:

$$\begin{aligned} \% \text{ solute} &= \frac{\text{mass solute}}{\text{mass solution}} \times 100 \\ &= \frac{\text{mass solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100 \end{aligned}$$

Note: if the question says the solution is aqueous or does not specify the solvent, the solvent is water, H_2O .

Example: Calculate the mass of potassium nitrate, KNO_3 required to prepare 250.0 g of solution that is 20.0% KNO_3 by mass. What is the mass of water in the solution?

$$\begin{aligned} \text{(a) } \% \text{ by mass} &= \frac{\text{g } \text{KNO}_3}{\text{g solution}} \times 100 \\ 20.0\% &= \frac{\text{g } \text{KNO}_3}{250.0 \text{ g}} \times 100 \\ \text{g } \text{KNO}_3 &= \frac{20.0\% \times 250.0 \text{ g}}{100} \\ &= 50.0 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{(b) mass of solution} &= \text{mass of } \text{KNO}_3 + \text{mass } \text{H}_2\text{O} \\ \text{mass } \text{H}_2\text{O} &= \text{mass of solution} - \text{mass of } \text{KNO}_3 \\ &= 250.0 - 50.0 \text{ g} \\ &= 200.0 \text{ g} \end{aligned}$$

Example: What volume of 24.0% lead(II) nitrate, $\text{Pb}(\text{NO}_3)_2$ solution contains 40.0 g of $\text{Pb}(\text{NO}_3)_2$? The density of the solution is 1.25 g/mL.

$$\begin{aligned} \text{Step 1 } \% \text{ by mass} &= \frac{\text{g } \text{Pb}(\text{NO}_3)_2}{\text{g solution}} \times 100 \\ 24.0\% &= \frac{40.0 \text{ g}}{\text{g solution}} \times 100 \\ \text{g solution} &= \frac{40.0 \text{ g}}{24.0} \times 100 \\ &= 167 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Step 2: } D \left(\frac{\text{g}}{\text{mL}} \right) &= \frac{\# \text{ g solute}}{\# \text{ mL}} \\ \# \text{ mL} &= \frac{\# \text{ g}}{D} \\ &= \frac{167 \text{ g}}{1.25 \text{ g/mL}} \\ &= 134 \text{ mL} \end{aligned}$$

2. Molarity :

$$\text{molarity (M)} = \frac{\text{moles solute}}{\text{liters of solution}} \quad \text{units: } \left(\frac{\text{mol}}{\text{L}}\right) \text{ or } \left(\frac{\text{mmol}}{\text{mL}}\right)$$

(~~Show Fig. 3-2~~ : preparation of a solution)

Example: Calculate the mass of ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, required to prepare 100. mL of 0.300 M $(\text{NH}_4)_2\text{SO}_4$ solution

$$M \left(\frac{\text{moles}}{\text{L}}\right) = \frac{\# \text{ moles}}{\# \text{ L}}$$

$$\text{Plan: } M, \# \text{ L} \xrightarrow{\textcircled{1}} \text{ moles} \xrightarrow{\textcircled{2}} \text{ g}$$

$$\begin{aligned} \text{Step } \textcircled{1} \quad \# \text{ moles} &= M \times \# \text{ L} \\ &= 0.300 \frac{\text{mol}}{\text{L}} \times 0.100 \text{ L} \\ &= 0.0300 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Step } \textcircled{2} \quad \# \text{ g} &= \# \text{ moles} \times \text{MW} && \text{since } \text{MW} = \frac{\# \text{ g}}{\# \text{ moles}} \\ &= 0.0300 \text{ mol} \times 132.13 \frac{\text{g}}{\text{mol}} \\ &= 3.96 \text{ g} \end{aligned}$$

Example: The specific gravity of commercial nitric acid is 1.44 and it is 70.4 % HNO_3 by mass. Calculate the molarity of HNO_3 .

Assume we have 100 g of solution; therefore the solution contains 70.4 g HNO_3 :

$$\therefore \text{concentration is } \frac{70.4 \text{ g } \text{HNO}_3}{100 \text{ g solution}} \xrightarrow[\text{mL} = \frac{\text{g}}{\text{D}}]{\text{mol} = \frac{\text{g}}{\text{MW}}} \frac{\# \text{ moles } \text{HNO}_3}{\# \text{ L solution}}$$

Plan: change numerator (g \rightarrow moles); change denominator (g \rightarrow # L)
then divide. since $M = \frac{\# \text{ moles}}{\# \text{ L}}$

$$\text{numerator: moles HNO}_3 = \frac{\text{g HNO}_3}{\text{MW}} = \frac{70.4 \text{ g}}{63.01 \text{ g/mol}} = 1.12 \text{ moles}$$

$$\text{denominator: mL solution} = \frac{\text{g solution}}{\text{sp. gr. or density}} = \frac{100. \text{ g}}{1.44 \text{ g/mL}} = 69.4 \text{ mL}$$

$$\# \text{ L solution} = 69.4 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0694 \text{ L}$$

$$\therefore M = \frac{\# \text{ mol}}{\# \text{ L}} = \frac{1.12 \text{ mol}}{0.0694 \text{ L}} = \underline{\underline{16.1 \text{ M}}}$$

Example: What volume of 18.0 M H_2SO_4 , sulfuric acid, is required to prepare 2.00 L of 1.40 M H_2SO_4 ?

This is a DILUTION problem, in which a solution is made less concentrated by adding more solvent (in this case, H_2O).

In a dilution problem, the number of moles of solute before dilution equals the number of moles of solute after dilution, since all we are adding is solvent (H_2O)

$$\therefore \# \text{ moles before} = \# \text{ moles after}$$

$$M_{\text{before}} \times \#L_{\text{before}} = M_{\text{after}} \times \#L_{\text{after}} \quad \begin{array}{l} M = \frac{\# \text{ moles}}{\# L} \\ \therefore \text{moles} = M \times \#L \end{array}$$

or more generally

$$M_{\text{before}} V_{\text{before}} = M_{\text{after}} V_{\text{after}} \quad \text{where } V = \text{mL, L} \\ \text{any volume}$$

$$18.0 \text{ M} \times V = 1.40 \text{ M} \times 2.00 \text{ L}$$

$$V = \underline{\underline{0.156 \text{ L}}} \quad \text{or} \quad \underline{\underline{156 \text{ mL}}}$$

Stoichiometry and Solutions:

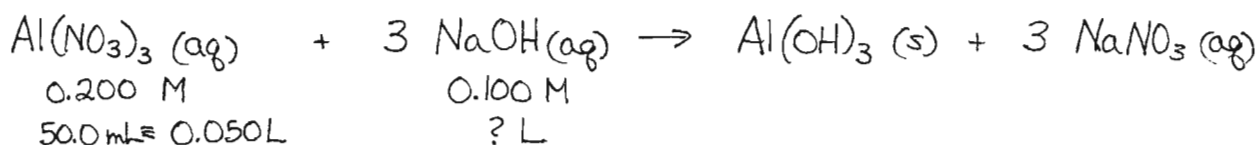
In stoichiometric calculations, we must work in moles of substance.

Up until now, our substances could be weighed out on a balance to get mass, then we converted mass (g) to moles

We now have another way to handle substances in the laboratory - in solutions. Therefore if we know concentration (eg. M) and amount of solution (eg. #L), we can calculate amount of solute (eg. moles = $M \times \#L$) and use this in an equation.

Example: What volume of 0.100 M NaOH solution is required to react with 50.0 mL of 0.200 M $\text{Al}(\text{NO}_3)_3$?

Note: (aq) means aqueous ie. dissolved in H_2O
(s): solid



Plan: $M, \#L \xrightarrow{\textcircled{1}} \text{moles} \xrightarrow{\textcircled{2}} \text{moles} \xrightarrow{\textcircled{3}} \#L$
 $\text{Al}(\text{NO}_3)_3 \quad \text{Al}(\text{NO}_3)_3 \quad \text{NaOH} \quad \text{NaOH}$

$$\begin{aligned} \textcircled{1} \text{ moles } \text{Al}(\text{NO}_3)_3 &= M \times \#L && \text{since } M = \frac{\# \text{ moles}}{\# \text{ L solution}} \\ &= 0.200 \text{ M} \times 0.050 \text{ L} \\ &= 0.0100 \text{ moles} \end{aligned}$$

$$\begin{aligned} \textcircled{2} \text{ moles } \text{NaOH} &= 3 \times \text{moles } \text{Al}(\text{NO}_3)_3 \\ &= 0.0300 \text{ moles} \end{aligned}$$

$$\begin{aligned} \textcircled{3} \#L \text{ NaOH} &= \frac{\# \text{ moles NaOH}}{M} \\ &= \frac{0.0300 \text{ mol}}{0.100 \text{ M}} = 0.300 \text{ L} \end{aligned}$$

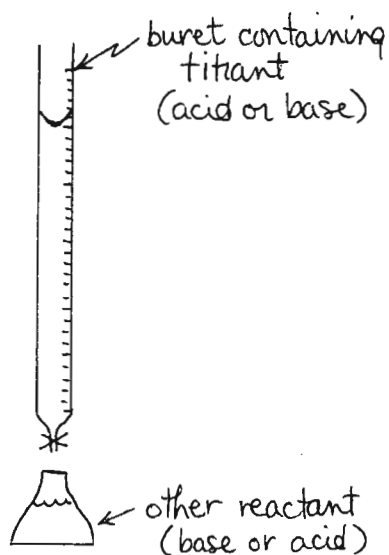
Example (continued): What mass of aluminum hydroxide, $\text{Al}(\text{OH})_3$, precipitates (becomes a solid)?

From the equation, we know that for every mole of $\text{Al}(\text{NO}_3)_3$ used up, one mole of $\text{Al}(\text{OH})_3 (\text{s})$ is formed.

$$\begin{aligned} \therefore \# \text{ moles } \text{Al}(\text{OH})_3 &= \# \text{ moles } \text{Al}(\text{NO}_3)_3 = 0.0100 \text{ moles} \\ \therefore \text{mass } \text{Al}(\text{OH})_3 &= \text{moles} \times \text{MW} = 0.0100 \text{ mol} \times 78.0 \frac{\text{g}}{\text{mol}} = \underline{\underline{0.780 \text{ g}}} \end{aligned}$$

In the laboratory, acid-base reactions are carried out as a titration:

Titration: process by which a solution of one reactant (the titrant) is added to the other reactant, and the volume of titrant required for complete reaction is measured.



The number of moles of one reactant is $(\text{moles} = M \times V)$ known either by knowing M and V or by weighing out a known solid.

The volume of the other is known and the unknown concentration is calculated when the reaction is just completed. It is known when the reaction is finished by adding an indicator to the flask - it changes color when reaction is over.

Standard solution: a solution of known concentration.

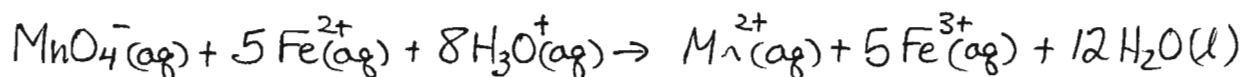
Standardization: process by which the concentration of a ^{standard} solution is determined by carefully measuring the volume of solution required to react with an exactly known amount of a PRIMARY STANDARD. The standardized solution is known as the SECONDARY STANDARD and is used in the analysis of an unknown.

Equivalence Point: the point in a titration at which chemically equivalent amounts of reactants have reacted.

Endpoint: the point in a titration at which the indicator changes color and the titration is stopped.

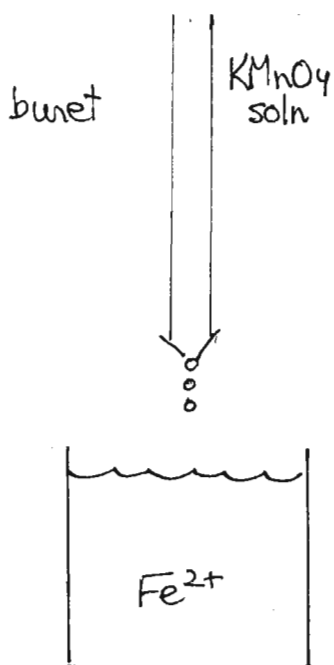
Redox titration

To analyze an iron-containing compound, you convert all the iron to Fe^{2+} in solution then titrate the solution with standardized KMnO_4 . The balanced net ionic equation is

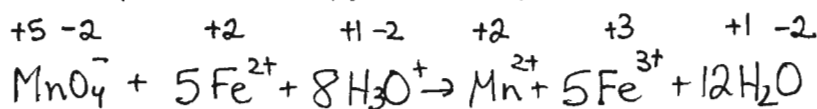


A 0.598 g sample of the iron containing compound requires 22.25 mL of 0.0123 M KMnO_4 for titration to the equivalence point. What is the mass % of iron in the sample? (What happened to the K in KMnO_4 ?)*

Plan: $M, V \text{ KMnO}_4 \rightarrow \text{mol KMnO}_4 \rightarrow \text{mol Fe}^{2+} = \text{mol Fe} \rightarrow \text{g Fe (pure)}$
 \downarrow
 % Fe in sample



Oxidation - Reduction reaction



Mn is element being reduced $\text{Mn}^{+5} \rightarrow \text{Mn}^{+2}$

$\therefore \text{MnO}_4^-$ is the oxidizing agent

Fe is element being oxidized $\text{Fe}^{+2} \rightarrow \text{Fe}^{+3}$

$\therefore \text{Fe}^{2+}$ is the reducing agent.

$$\begin{aligned} ? \text{ g Fe (pure)} &= 0.02225 \text{ L MnO}_4^- \times \frac{0.0123 \text{ mol MnO}_4^-}{1 \text{ L MnO}_4^-} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}^{2+}} \\ &= 0.0764 \text{ g Fe} \end{aligned}$$

$$\% \text{ Fe in ore} = \frac{0.0764 \text{ g Fe}}{0.598 \text{ g ore}} \times 100 = 12.8 \% \text{ Fe}$$

*its a spectator ion!

pH - measure of the molar concentration of H^+ ion

In aqueous solution, the concentration of H^+ can vary widely:
from $> 1 M$ to $< 1 \times 10^{-14} M$. To manage these numbers,
it is easier to work in logarithms

Review: logs are exponents!

$$\log_{10} 1000 = 3$$

$$\Downarrow$$

$$10^3 = 1000$$

$$\log_{10} (1 \times 10^{-6}) = -6$$

$$\Downarrow$$

$$10^{-6} = 1 \times 10^{-6}$$

Definition of pH: $pH = -\log_{10} (H^+ \text{ concentration in molarity})$
 $= -\log_{10} [H^+]$ where $[]$ means M

$[H^+]$	$\log [H^+]$	$pH = -\log [H^+]$	
1.00 M	0	0	acidic
0.10 M = 10^{-1}	-1	+1	acidic
$1 \times 10^{-5} M$	-5	+5	acidic
$1 \times 10^{-7} M$	-7	7	neutral
$1 \times 10^{-9} M$	-9	9	basic
$1 \times 10^{-14} M$	-14	14	basic

What is the pH of a $2.5 \times 10^{-4} HNO_3$ soln?

HNO_3 is a strong acid so $[HNO_3] = [H^+]$ $HNO_3 \rightarrow H^+ + NO_3^-$

$$pH = -\log_{10} (2.5 \times 10^{-4}) = 3.60$$

What is the concentration of an HCl solution with $pH = 1.28$?

↑ another strong acid with $[H^+] = [HCl]$

$$pH = -\log [H^+] = 1.28$$

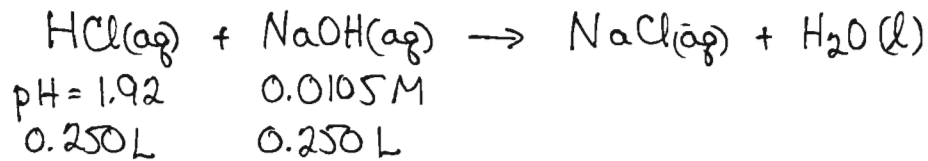
$$\log [H^+] = -1.28$$

take inverse log (also called 10^x) of both sides

$$[H^+] = 10^{-1.28} = 0.053$$

Acid Base Titration

A solution of hydrochloric acid has a volume of 250. mL and pH of 1.92. Exactly 250. mL of 0.0105 M NaOH is added. What is the pH of the resulting solution?



$$\begin{aligned} \text{HCl: } \text{pH} = 1.92 & \therefore -\log_{10} [\text{H}^+] = 1.92 \\ & \log_{10} [\text{H}^+] = -1.92 \\ & [\text{H}^+] = 10^{-1.92} = 0.0120 \text{ M} \\ ? \text{ mol HCl} & = 0.0120 \text{ M} \times 0.250 \text{ L} = 0.00300 \text{ mol HCl} \end{aligned}$$

$$\text{NaOH: } ? \text{ mol NaOH} = 0.0105 \text{ M} \times 0.250 \text{ L} = 0.002625 \text{ mol NaOH}$$

What kind of a problem is this? Limiting reagent!! But easy one!

What is limiting reactant? What gets used up first? NaOH
When the reaction is over, all NaOH is gone; been converted to NaCl + H₂O

How much HCl remains?

since stoichiometric ratio is 1:1, we can just subtract! not significant

$$\begin{aligned} ? \text{ mol HCl in excess} & = 0.00300 \text{ mol HCl} - 0.002625 \text{ mol HCl} \\ & = 0.000375 \text{ mol HCl} \end{aligned}$$

↑ not significant

To find the pH of the solution, we first need to find [HCl]

$$? \text{ M HCl} = \frac{0.000375 \text{ mol HCl}}{(0.250 \text{ L} + 0.250 \text{ L})} = 0.00075 \text{ M} = [\text{H}^+]$$

total V

since HCl is strong acid

$$\text{pH} = -\log_{10} [\text{H}^+] = -\log_{10} (0.00075) = 3.12$$