Example:

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

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

1) Convert 25.0 grams sodium carbonate to moles. Use FORMULA WEIGHT.
2) Convert moles sodium carbonate to moles HCl. Use CHEMICAL EQUATION.
3) Convert moles HCl to volume HCl solution. Use MOLARITY (6.00 M HCl)

③ 6.00 mol HCl = L

$$0.4717426172\,mol\,HCl \times \frac{L}{6.00\,mol\,HCl} = 0.0786L$$

Since the problem specifies to use milliliters, so do a quick unit conversion!

$$mL = 10^{-3}L$$

$$0.0786L \times \frac{mL}{10^{-3}L} = \boxed{78.6\,mL\text{ of }6.00\,M\,HCl}$$

$$4\,\underset{\text{propylene}}{\overset{42.081\text{ g/mol}}{C_3H_6}} + 6NO \longrightarrow 4\,\underset{\text{acrylonitrile}}{\overset{53.064\text{ g/mol}}{C_3H_3N}} + 6H_2O + N_2$$

Calculate how many grams of acrylonitrile could be obtained from 651 g of propylene, assuming there is excess NO present.

---

1) Convert 651 g propylene to moles. Use FORMULA WEIGHT.
2) Convert moles propylene to acrylonitrile. Use CHEMICAL EQUATION.
3) Convert moles acrylonitrile to mass. Use FORMULA WEIGHT.

---

① $42.081\text{ g }C_3H_6 = \text{mol }C_3H_6$ ② $4\text{ mol }C_3H_6 = 4\text{ mol }C_3H_3N$

③ $53.064\text{ g }C_3H_3N = \text{mol }C_3H_3N$

$$651\text{ g }C_3H_6 \times \underset{①}{\frac{\text{mol }C_3H_6}{42.081\text{ g }C_3H_6}} \times \underset{②}{\frac{4\text{ mol }C_3H_3N}{4\text{ mol }C_3H_6}} \times \underset{③}{\frac{53.064\text{ g }C_3H_3N}{\text{mol }C_3H_3N}} =$$

$$= \boxed{821\text{ g }C_3H_3N}$$

151.90 g/mol

$$10\,FeSO_4 + 2\,KMnO_4 + 8\,H_2SO_4 \rightarrow 5\,Fe_2(SO_4)_3 + 2\,MnSO_4 + K_2SO_4 + 8\,H_2O$$

How many mL of 0.250M potassium permangenate are needed to react with 3.36 g of iron(II) sulfate?

1) Convert 3.36 g iron(II) sulfate to moles. Use FORMULA WEIGHT.
2) Convert moles iron(II) sulfate to moles potassium permangenate. Use CHEMICAL EQUATION.
3) Convert moles potassium permangenate to volume. Use MOLARITY (0.250 M)

① $151.90\,g\,FeSO_4 = mol\,FeSO_4$ ② $10\,mol\,FeSO_4 = 2\,mol\,KMnO_4$ ③ $0.250\,mol\,KMnO_4 = L$

$$3.36\,g\,FeSO_4 \times \underset{\textcircled{1}}{\frac{mol\,FeSO_4}{151.90\,g\,FeSO_4}} \times \underset{\textcircled{2}}{\frac{2\,mol\,KMnO_4}{10\,mol\,FeSO_4}} \times \underset{\textcircled{3}}{\frac{L}{0.250\,mol\,KMnO_4}} = 0.0177\,L$$

As before, this problem asks for the final answer in mL. Convert ...

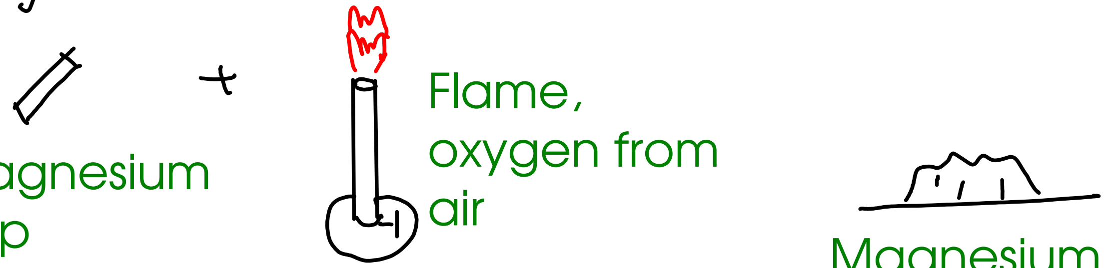
$mL = 10^{-3}L$

$$0.0177\,L \times \frac{mL}{10^{-3}\,L} = \boxed{17.7\,mL\ of\ 0.250M\ KMnO_4}$$

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?

---

① 56.08 g CaO = mol CaO ② mol CaO = mol CaC_2 ③ 64.10 g CaC_2 = mol CaC_2

$$100.\text{ g CaO} \times \frac{\text{mol CaO}}{56.08\text{ g CaO}} \times \frac{\text{mol CaC}_2}{\text{mol CaO}} \times \frac{64.10\text{ g CaC}_2}{\text{mol CaC}_2} = \boxed{114\text{ g CaC}_2}$$

---

① 12.01 g C = mol C ② 3 mol C = mol CaC_2 ③ 64.10 g CaC_2 = mol CaC_2

$$100.\text{ g C} \times \frac{\text{mol C}}{12.01\text{ g C}} \times \frac{\text{mol CaC}_2}{3\text{ mol C}} \times \frac{64.10\text{ g CaC}_2}{\text{mol CaC}_2} = 178\text{ g CaC}_2$$

The reaction produces 114 grams of calcium carbide. Once that amount is produced, the reaction must stop because there is no more CaO. There's enough carbon to make 178 grams, but that can't happen because there isn't any more CaO for that excess carbon to react with! We say that CaO is LIMITING, and C 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:

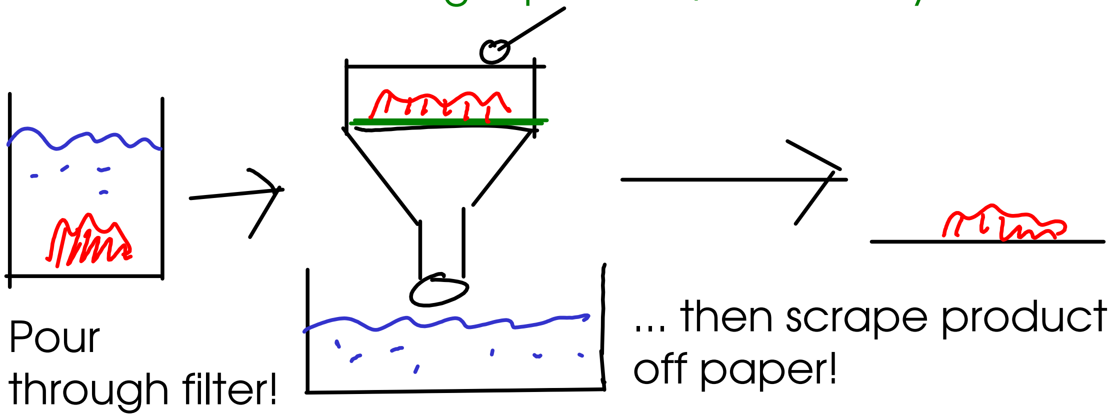
$C + O_2 \longrightarrow CO_2$ | This reaction occurs when there is a large amount of oxygen available

$2C + O_2 \longrightarrow 2CO$ | ... 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!

78.114 g|mol

22.4 g

123.111 g|mol <- Formula weights

31.6 g ACTUAL

$$C_6H_6 + HNO_3 \longrightarrow C_6H_5NO_2 + H_2O$$

benzene

nitric acid

nitrobenzene

22.4 grams of benzene are reacted with excess nitric acid. If 31.6 grams of nitrobenzene are collected from the reaction, what is the percent yield?

To find the percent yield, we need to CALCULATE the THEORETICAL YIELD of nitrobenzene based on 22.4 g of the starting material (regular benzene). Then compare the theoretical yield with the ACTUAL YIELD of 31.6 grams nitrobezene.

① $78.114g\ C_6H_6 = mol\ C_6H_6$ ② $mol\ C_6H_6 = mol\ C_6H_5NO_2$

③ $123.111g\ C_6H_5NO_2 = mol\ C_6H_5NO_2$

$$22.4g\ C_6H_6 \times \frac{mol\ C_6H_6}{78.114g\ C_6H_6} \times \frac{mol\ C_6H_5NO_2}{mol\ C_6H_6} \times \frac{123.111g\ C_6H_5NO_2}{mol\ C_6H_5NO_2} = 35.3g\ C_6H_5NO_2$$

(theoretical yield)

$$\%yield = \frac{actual}{theor.} \times 100 \qquad \frac{31.6g}{35.3g} \times 100 = \boxed{89.5\%}$$

Electrolytes and Ionic Theory

- electrolytes: substances that dissolve in water to form charge-carrying solutions

* Electrolytes form ions in solution - (ions that are mobile are able to carry charge!). These IONS can interact with one another and undergo certain kinds of chemistry!

IONIC THEORY

- the idea that certain compounds DISSOCIATE in water to form free IONS

Strong vs weak?

- If an electrolyte COMPLETELY IONIZES in water, it's said to be STRONG

- If an electrolyte only PARTIALLY IONIZES in water, it's said to be WEAK

- Both kinds of electrolyte undergo similar kinds of chemistry.

What kinds of compounds are electrolytes?

MOLECULAR COMPOUNDS

- Most molecular compounds are NONELECTROLYTES - they don't ionize in water

-ACIDS and BASES will ionize in water. Most of these are WEAK ELECTROLYTES, but there are a few STRONG ACIDS and STRONG BASES.

$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$ ... nonelectroiyte

$HC_2H_3O_2(l) \xrightarrow{H_2O} H^+(aq) + C_2H_3O_2^-(aq)$

... acetic acid (electrolyte)

IONIC COMPOUNDS

- SOLUBLE ionic compounds are STRONG ELECTROLYTES - they completely ionize in water.

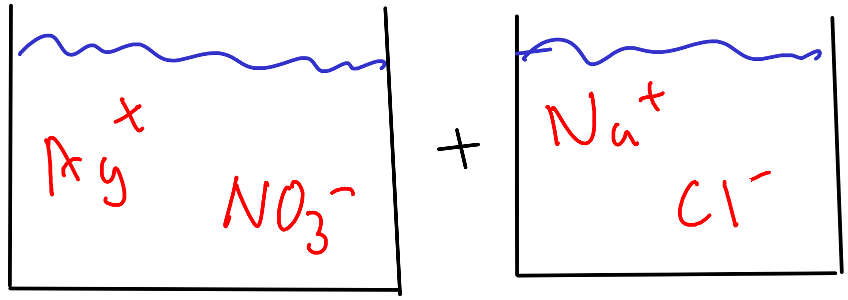
- Not all ionic compounds are water soluble, however!

$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$

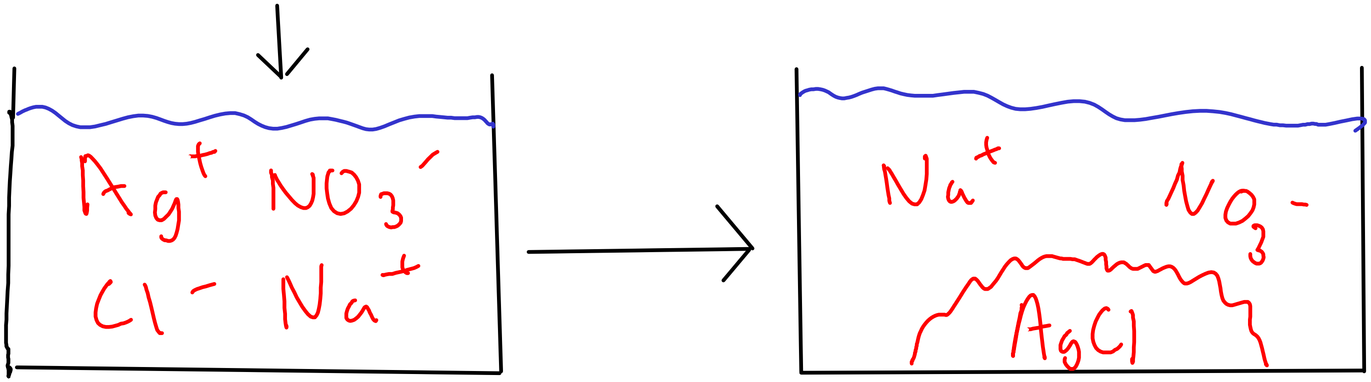
- What good is ionic theory?

- provides an easy-to-understand MECHANISM for certain kinds of chemical reactions.
  - "Exchange" reactions. (a.k.a "double replacement" reactions)

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$



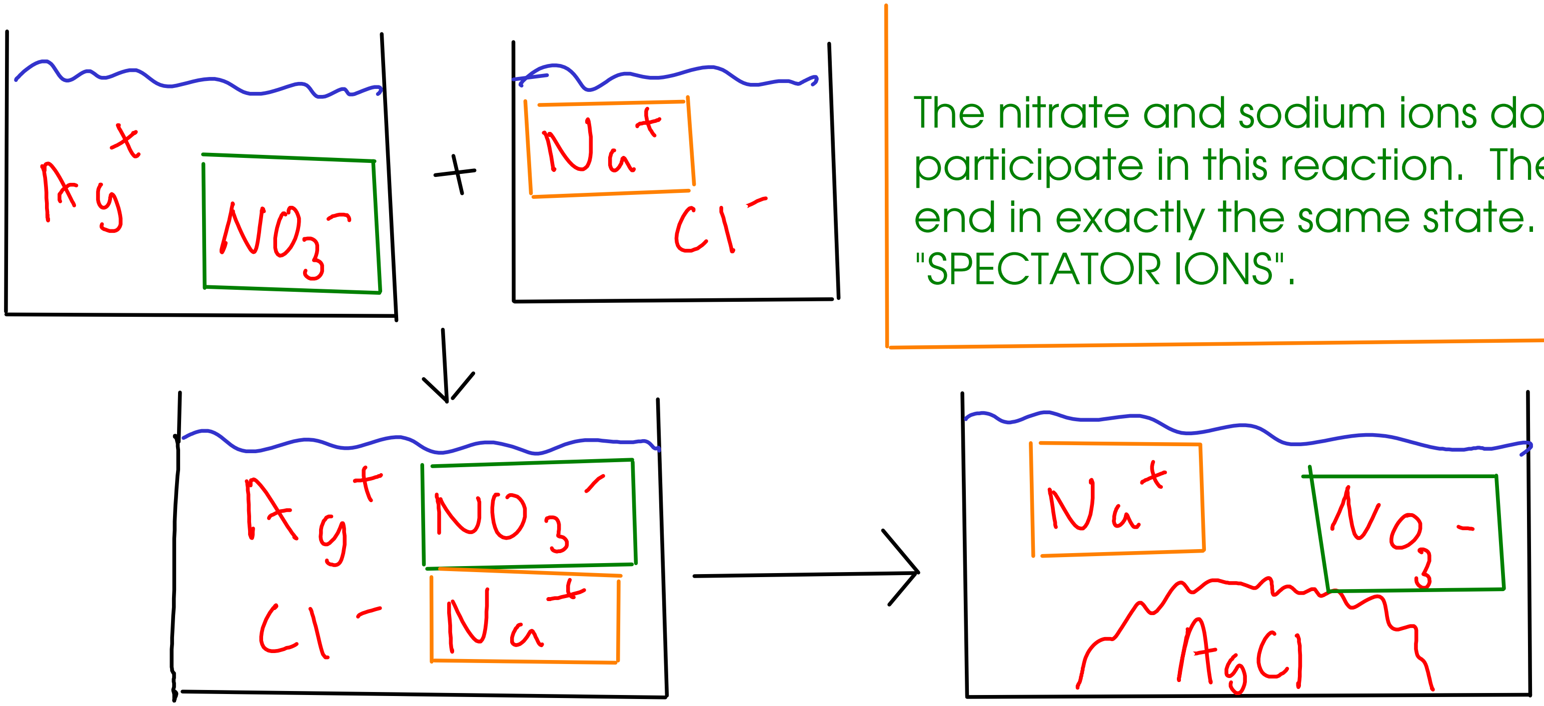
These free ions mix and can interact with each other!



"ion soup"!

Insoluble AgCl falls out of solution as it is formed - "precipitation"

Looking a bit more closely...



The nitrate and sodium ions do not really participate in this reaction. They start and end in exactly the same state. We call them "SPECTATOR IONS".

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

"Molecular equation"

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

"Net ionic equation"

(The net ionic equation shows only ions and substances that change during the course of the reaction!)

- The net ionic equation tells us that any source of aqueous silver and chloride ions will exhibit this same chemistry, not just silver nitrate and sodium chloride!

A bit more about molecular, ionic, and net ionic equations

- molecular equations: Represent all substances (even ionic substances) as if they were molecules. Include spectator ions, and do not show charges on ions. Traditional chemical equations.

- ionic equations: Show all free ions - including spectators - in a chemical reaction. Molecules and WEAK electrolytes are shown as molecules. STRONG electrolytes (like HCl) are shown as ions. Ions that are part of undissolved ionic compounds are shown as molecules.

- NET ionic equation: An ionic equation that leaves out spectator ions. Intended to show only things that actually change in a reaction.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow \underline{AgCl(s)} + Na^+(aq) + NO_3^-(aq)\ ^*$$

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

* You can get from the complete ionic equation to the net ionic equation by crossing out the spectator ions on both sides.

"Undissolved ionic compounds":

How can I tell if an ionic compound dissolves in water?

- consult experimental data: "solubility rules"!

A few of the "rules"...

- Compounds that contain a Group IA cation (or ammonium) are soluble
- Nitrates and acetates are soluble
- Carbonates, phosphates, and hydroxides tend to be insoluble

See p 129 9th edition (10th ed: p131)

... or see the web site for a solubility chart.

$Fe(OH)_3$

#8 - hydroxides generally insoluble, except Group IA, ammonium, calcium strontium, barium

Conclusion: iron(III) hydroxide is insoluble.

$Ca(C_2H_3O_2)_2$

#2 - acetates are soluble, no common exceptions.

Conclusion: calcium acetate is soluble.

AgI

#3 - Iodides usually dissolve, exceptions are silver, mercury, lead

Conclusion: silver(I) iodide is INSOLUBLE

$BaCO_3$

#5 - Most carbonates are insoluble

Conclusion - barium carbonate is insoluble.