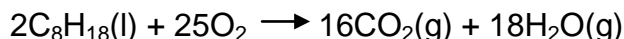


REACTIONS AND YIELD ANSWERS

stoichiometry = the numerical relationships between chemical amounts in a reaction.



From the equation, 16 moles of CO_2 (a greenhouse gas resulting in global warming) are produced for every 2 moles of octane burned (by combustion).

Estimate the mass of CO_2 produced in 2004 by the combustion of 3.4×10^{15} g gasoline.

Solution: number of moles = mass/molar mass; molar mass = molecular mass
molecular masses: $\text{C}_8\text{H}_{18} = 114.22$; $\text{CO}_2 = 44.01 \text{ gmol}^{-1}$

$$3.4 \times 10^{15} / 114.22 = 2.977 \times 10^{13} \text{ moles C}_8\text{H}_{18}.$$

$$\text{stoichiometric ratio } \text{CO}_2: \text{C}_8\text{H}_{18} = 16:2 \Leftrightarrow 2.977 \times 10^{13} \times (16/2) = 2.382 \times 10^{14} \text{ moles } \text{CO}_2 \\ = 1.05 \times 10^{16} \text{ g } \text{CO}_2.$$

limiting reactant (or reagent) = the reactant that limits the amount of product

excess reactants = reactants not completely consumed

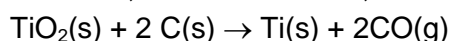
theoretical yield = the amount of product that can be made from the limiting reactant

actual yield = the amount of product that is made in a reaction; actual yield
generally less than the theoretical yield, never more!

percent yield = the efficiency of product recovery

$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$
--

When 28.6 kg of C reacts with 88.2 kg of TiO_2 , 42.8 kg of Ti are obtained. Find the Limiting Reactant, Theoretical Yield, and Percent Yield.



Solution: 1 mole TiO_2 gives 1 mole Ti; 2 moles C gives 1 mole Ti.
molar masses C = 12.01; $\text{TiO}_2 = 79.87$; Ti = 47.87 gmol^{-1} .

$$\text{Number of moles of C} = 2.38 \times 10^3 \text{ moles} \Leftrightarrow 1.191 \times 10^3 \text{ moles Ti}$$

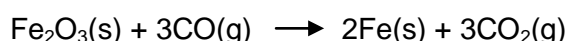
$$\text{Number of moles of } \text{TiO}_2 = 2.38 \times 10^3 \text{ moles} \Leftrightarrow 1.104 \times 10^3 \text{ moles Ti}$$

$$\text{TiO}_2 = \text{limiting reactant} \Leftrightarrow 1.104 \times 10^3 \times 47.87 = 52.9 \text{ kg Ti} = \text{theoretical yield}$$

$$\text{Percent yield} = 100\% \times 42.8 / 52.9 = 80.9 \%$$

PRACTICE EXAMPLE ONE

Mining companies use the following reaction to obtain iron from iron ore:



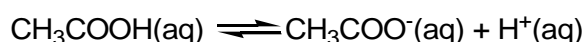
The reaction of 167g of Fe_2O_3 with 85.8g of CO produces 72.3g Fe. Find the limiting reagent, theoretical yield and percent yield.

CO is the limiting reagent since it only produces 114 g Fe (work out number of moles
 $\text{CO} = 85.8/28.01 = 3.063$ so $\text{max Fe} = 2/3 \times 3.063 = 2 = 2 \times 55.85 = 114 \text{ g Fe}$)
 Therefore 114 g Fe is the theoretical yield; percentage yield = $100\% \times 72.3 / 114 = 63.4 \%$ _____

Making aqueous solutions

When an ionic compound **dissolves** in water, the water molecules surround the cations and anions; e.g. $\text{KCl(s)} \rightarrow \text{K}^+(\text{aq}) + \text{Cl}^-(\text{aq})$. Heating the solution increases solubility.

Substances such as potassium chloride or hydrogen chloride (strong acid) that completely dissociate into ions are **strong electrolytes** (solution conducts electrical current). **Weak electrolytes** dissolve mostly as molecules but partly dissociate into ions:



dilute solutions = small amount of solute compared to solvent

concentrated solutions = large amount of solute compared to solvent

$$\text{Molarity} = \frac{\text{amount of solute (in moles)}}{\text{amount of solution (in L or dm}^3\text{)}}$$

Find the **molarity, M** (or **concentration, C**) of a solution that has 25.5 g KBr dissolved in 1.75 L of solution

Solution: molar mass KBr = 119.00 gmol^{-1}
 number of moles = $25.5 / 119 = 0.214 \text{ mol KBr}$
 molarity = $0.214 / 1.75 = \mathbf{0.122 \text{ M}}$ (or mol dm^{-3}).

Dilution = make a solution less concentrated

moles solute in solution 1 = moles solute in solution 2

$$M_1V_1 = M_2V_2$$

To what volume should you dilute 0.200 L of 15.0 M NaOH to make 3.00 M NaOH?

Solution: $V_2 = M_1V_1 / M_2 = 15 \times 0.2 / 3 = \mathbf{1 \text{ L}}$

PRACTICE EXAMPLE TWO

What volume of a 6.00 M NaNO_3 solution should you use to make 0.525 L of a 1.20 M NaNO_3 solution?

$$\text{_____ } V_2 = M_1V_1/M_2 = 1.20 \times 0.525 / 6 = 0.105 \text{ L} \text{_____}$$

What volume of 0.150 M KCl is required to completely react with 0.150 L of 0.175 M $\text{Pb}(\text{NO}_3)_2$ in the reaction $2\text{KCl}(\text{aq}) + \text{Pb}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) + 2\text{KNO}_3(\text{aq})$?

1 L $\text{Pb}(\text{NO}_3)_2 = 0.175 \text{ mol}$, 1 L KCl = 0.150 mol,
stoichiometry: 1 mol $\text{Pb}(\text{NO}_3)_2$ reacts with 2 mol KCl

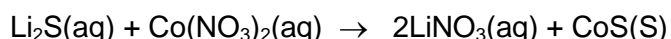
Number of moles of $\text{Pb}(\text{NO}_3)_2$ used = molarity x volume = 0.0263 moles

Number of moles of KCl used = 0.02625 moles x 2 = 0.0525 moles

volume KCl = moles / molarity = 0.0525 / 0.150 = **0.35 L**

PRACTICE EXAMPLE THREE

Consider the following reaction:



What volume of 0.150 M Li_2S solution is required to completely react with 125 mL of 0.250 M $\text{Co}(\text{NO}_3)_2$?

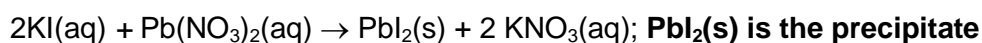
Stoichiometry: 1:1 so number of moles $\text{Co}(\text{NO}_3)_2 = 0.250 \times 0.125 = 0.031 \text{ moles} =$
volume $\text{Li}_2\text{S} = 0.031/0.150 = 0.208 \text{ L} = 208 \text{ mL}$

soluble = compounds that dissolve in a solvent (e.g. NaCl dissolves in water)

insoluble = compounds that do not dissolve (e.g. AgCl does not)

Most Group I compounds are soluble in water (e.g. NaOH, Na_2SO_4). Some Group II compounds are soluble (e.g. CaCl_2), slightly soluble e.g. $\text{Ca}(\text{OH})_2$, others are not (e.g. CaSO_4). Many ammonium NH_4^+ compounds are soluble (e.g. NH_4Cl) and ALL NITRATES (e.g. $\text{Ca}(\text{NO}_3)_2$) are soluble.

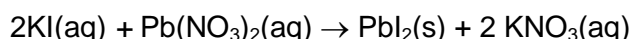
Precipitation reactions = reactions between aqueous solutions of ionic compounds that produce an ionic compound that is insoluble (**precipitate**) in water:



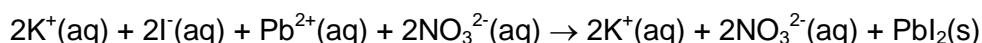
Discussed in more detail (**book p. 153**).

Writing and simplifying equations

molecular equations = equations which describe the chemicals put into the water and the product molecules molecular equations, e.g.

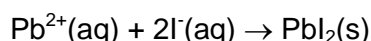


complete ionic equations = equations which describe the actual dissolved species:

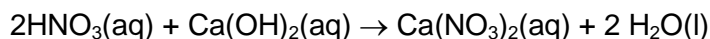


- ions that are both reactants and products are called **spectator ions (they do not react and remain in the solution)**

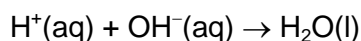
net ionic equation = an ionic equation in which the spectator ions are removed:



neutralization reactions = the acid and base neutralize each other's properties

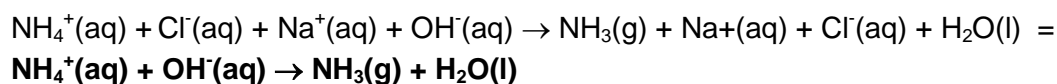
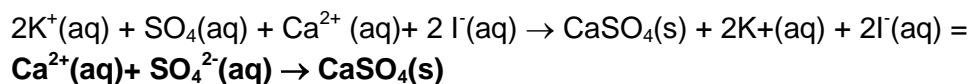
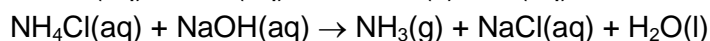
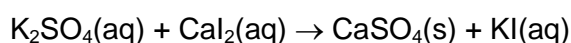


the net ionic equation for an acid-base reaction is:



PRACTICE EXAMPLE FOUR

Write complete ionic and net ionic equations for each reaction:



acids ionize in water to form H^{+} ions

more precisely, the H from the acid molecule is donated to a water molecule to form **hydronium ion, H_3O^{+}** . A proton (H^{+}) cannot exist on its own in water!

bases dissociate in water to form OH^{-} ions

bases, such as NH_3 , that do not contain OH^{-} ions, produce OH^{-} by pulling H off water molecules

acid + base → **salt + water**

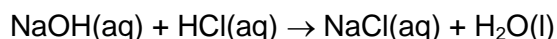
titration - a solution's concentration is determined by reacting it with another material and using stoichiometry (involving a chemical equation and calculation). An indicator permanently changes colour at the **end-point** (just finished).

For example, if we know the concentration and precise volume of the alkali but only the volume of acid of unknown concentration, we can determine its concentration. Titration can be used for many ionic reactions.



$$\boxed{\frac{C_1 \times V_1}{C_2 \times V_2} = \frac{n_1}{n_2}} \quad \text{or} \quad C_2 = \frac{C_1 \times V_1 \times n_2}{n_1 \times V_2}$$

12.54 mL of 0.100 M NaOH(aq) reacts 10.00 mL of HCl(aq) solution. What is the concentration of the acid? The formula can be used with mL (for both solutions) or converted to L for both solutions.



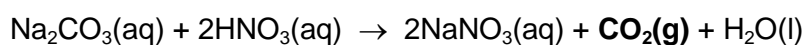
C_1 concentration = 0.100 M NaOH(aq); $C_2 = ?$

V_1 volume = 12.54 mL NaOH(aq); $V_2 = 10.00$ mL

n_1 stoichiometric value = 1 for NaOH; $n_2 = 1$ for HCl

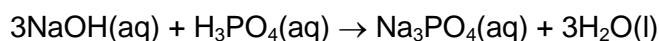
$$C_2 = 0.100 \times 12.54 \times 1 / 1 \times 10.00 = \mathbf{0.125 \text{ M HCl(aq)}}$$

Some neutralization reactions evolve gases; the intermediate carbonic acid (H_2CO_3) thus decomposes, e.g.



PRACTICE EXAMPLE FIVE

A 30.00 mL sample of an unknown H_3PO_4 solution is titrated with a 0.100 M NaOH solution. The end-point (equivalence point) is reached when 26.38 mL of NaOH solution is added. What is the concentration of the unknown H_3PO_4 solution?



Stoichiometry: $n_1 = 3$; $n_2 = 1$

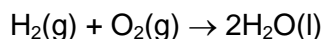
$$\text{so } C_2 = 0.100 \times 26.38 \times 1 / 30.00 \times 3 = 0.029 \text{ M H}_3\text{PO}_4$$

REDOX REACTIONS

REDOX (oxidation-reduction) reactions = transferring electrons from one atom to another
Many involve the use of O_2 , Cl_2 etc.

$4 \text{Fe(s)} + 3\text{O}_2(\text{g}) \rightarrow 2 \text{Fe}_2\text{O}_3(\text{s})$; the iron has transferred its electrons to the oxygen atoms

Combustion (molecular - involve the use of O_2) is also REDOX, e.g.



oxidation occurs when an atom's oxidation state **increases** during a reaction (Fe(0) to Fe +3). **LOSS OF ELECTRONS**

reduction occurs when an atom's oxidation state **decreases** during a reaction (O (0) to O (-2)). **GAIN OF ELECTRONS**

In a REDOX reaction, the oxidizing agent is always reduced and the reducing agent always oxidized.

$C + 2S \rightarrow CS_2$ carbon is the **reducing agent** (0 to +4) and sulphur the **oxidizing agent** (0 to -2).

Oxidation state and ionic charges are not to be confused; **see the table in book p.164.**

- for single atoms with a charge (e.g. Fe^{3+} , O^{2-}) the oxidation state is often the charge!
- oxygen is nearly always -2 (exception KO_2).
- Group 1 = +1; Group 2 = +2
- Transition metals have variable oxidation states (e.g. chromium from +2 to +6)
 $K_2Cr_2O_7$: Cr = +6 = potassium chromate(VI)
- elements (e.g. O_2 oxidation state = zero (0)).
- Metals have positive oxidation states.
- In a neutral atom or formula the sum (Σ) of the oxidation states = 0.
- An ion is equal to the charge of the ion.

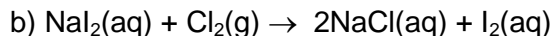
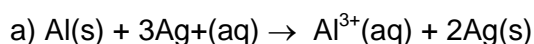
SO_4^{2-} , and H_2SO_4 sulphur has oxidation state of +6.

PRACTICE EXAMPLE SIX

What is the oxidation state of Cl in each ion?

1. a) ClO^- +1 b) ClO_2^- +3 c) ClO_3^- +5 d) ClO_4^- +7

2. Determine the *oxidizing* and *reducing* agent:



a) oxidizing agent = reduced (gain in electrons); reducing agent = oxidized (loss of electrons) so reducing agent = Al; oxidizing agent = Ag^+ . b) I^- = reducing agent; Cl_2 = oxidizing agent.

COMBUSTION REACTIONS

We have encountered several combustion reactions so far. Common ones include burning elemental carbon, hydrogen, sulphur, phosphorous, hydrocarbons (e.g. methane) or alcohol in air or oxygen. They give off heat (exothermic, $\Delta H < 0$ typically about 2000 kJ mol^{-1}) and light. See **book, p. 168-169 for equations. Equations must be balanced in an exam (number of atoms on the left = number of atoms on the right).**