

## REACTIONS AND YIELD ANSWERS

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



From the equation, 16 moles of  $\text{CO}_2$  (a greenhouse gas resulting in global warming) are produced for every 2 moles of octane burned (by combustion).

Estimate the mass of  $\text{CO}_2$  produced in 2004 by the combustion of  $3.4 \times 10^{15}$  g gasoline.

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

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

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

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

**excess reactants** = reactants not completely consumed

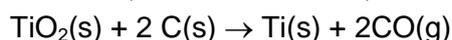
**theoretical yield** = the amount of product that can be made from the limiting reactant

**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

$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$
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When 28.6 kg of C reacts with 88.2 kg of  $\text{TiO}_2$ , 42.8 kg of Ti are obtained. Find the Limiting Reactant, Theoretical Yield, and Percent Yield.



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

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

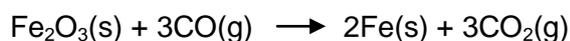
$$\text{Number of moles of TiO}_2 = 2.38 \times 10^3 \text{ moles} \Leftrightarrow 1.104 \times 10^3 \text{ moles Ti}$$

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

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

### PRACTICE EXAMPLE ONE

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



The reaction of 167g of  $\text{Fe}_2\text{O}_3$  with 85.8g of CO produces 72.3g Fe. Find the limiting reagent, theoretical yield and percent yield.

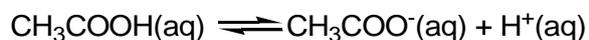
CO is the limiting reagent since it only produces 114 g Fe (work out number of moles  
 $\text{CO} = 85.8/28.01 = 3.063$  so  $\text{max Fe} = 2/3 \times 3.063 = 2 = 2 \times 55.85 = 114 \text{ g Fe}$ )  
 Therefore 114 g Fe is the theoretical yield; percentage yield =  $100\% \times 72.3 / 114 = 63.4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

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

**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  $\text{NaNO}_3$  solution should you use to make 0.525 L of a 1.20 M  $\text{NaNO}_3$  solution?

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

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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

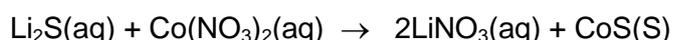
Number of moles of  $\text{Pb}(\text{NO}_3)_2$  used = molarity x volume = 0.0263 moles

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  $\text{Li}_2\text{S}$  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  $\text{Li}_2\text{S} = 0.031/0.150 = 0.208 \text{ L} = 208 \text{ mL}$

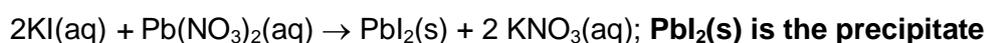
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**soluble** = compounds that dissolve in a solvent (e.g. NaCl dissolves in water)

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

Most Group I compounds are soluble in water (e.g. NaOH,  $\text{Na}_2\text{SO}_4$ ). Some Group II compounds are soluble (e.g.  $\text{CaCl}_2$ ), slightly soluble e.g.  $\text{Ca}(\text{OH})_2$ , others are not (e.g.  $\text{CaSO}_4$ ). Many ammonium  $\text{NH}_4^+$  compounds are soluble (e.g.  $\text{NH}_4\text{Cl}$ ) and ALL NITRATES (e.g.  $\text{Ca}(\text{NO}_3)_2$ ) are soluble.

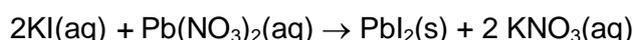
**Precipitation reactions** = reactions between aqueous solutions of ionic compounds that produce an ionic compound that is insoluble (**precipitate**) in water:



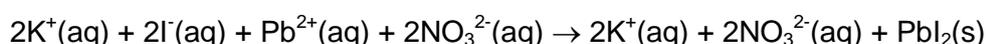
Discussed in more detail (**book p. 153**).

*Writing and simplifying equations*

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

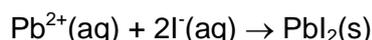


complete ionic equations = equations which describe the actual dissolved species:

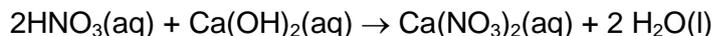


- ions that are both reactants and products are called **spectator ions (they do not react and remain in the solution)**

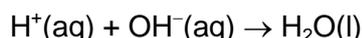
net ionic equation = an ionic equation in which the spectator ions are removed:



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

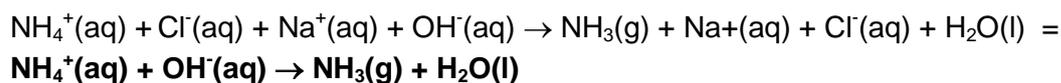
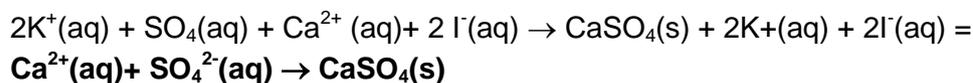
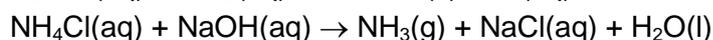
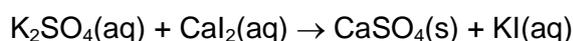


the net ionic equation for an acid-base reaction is:



#### PRACTICE EXAMPLE FOUR

Write complete ionic and net ionic equations for each reaction:



**acids** ionize in water to form  $\text{H}^{+}$  ions

more precisely, the H from the acid molecule is donated to a water molecule to form **hydronium ion,  $\text{H}_3\text{O}^{+}$** . A proton ( $\text{H}^{+}$ ) cannot exist on its own in water!

**bases** dissociate in water to form  $\text{OH}^{-}$  ions

bases, such as  $\text{NH}_3$ , that do not contain  $\text{OH}^{-}$  ions, produce  $\text{OH}^{-}$  by pulling H off water molecules

**acid + base** → **salt + water**

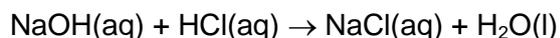
**titration** - a solution's concentration is determined by reacting it with another material and using stoichiometry (involving a chemical equation and calculation). An indicator permanently changes colour at the **end-point** (just finished).

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

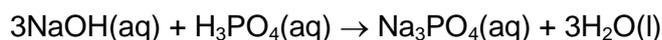
$$C_2 = 0.100 \times 12.54 \times 1 / 1 \times 10.00 = \mathbf{0.125 \text{ M HCl(aq)}}$$

Some neutralization reactions evolve gases; the intermediate carbonic acid ( $\text{H}_2\text{CO}_3$ ) thus decomposes, e.g.



### PRACTICE EXAMPLE FIVE

A 30.00 mL sample of an unknown  $\text{H}_3\text{PO}_4$  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  $\text{H}_3\text{PO}_4$  solution?



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

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

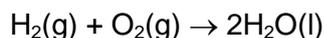

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### REDOX REACTIONS

REDOX (oxidation-reduction) reactions = transferring electrons from one atom to another  
Many involve the use of  $\text{O}_2$ ,  $\text{Cl}_2$  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  $\text{O}_2$ )** 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 ( $\Sigma$ ) of the oxidation states = 0.
- An ion is equal to the charge of the ion.

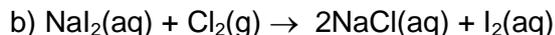
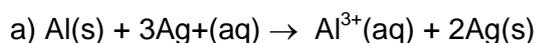
$SO_4^{2-}$ , and  $H_2SO_4$  sulphur has oxidation state of +6.

### 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  $2000 \text{ kJ mol}^{-1}$ ) 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).**