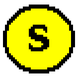
STOICHIOMETRY AND MOLE CONCEPT

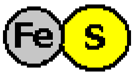
The subject of **stoichiometry** involves quantitative calculations based on chemical formulas and chemical equations.

Interpreting Chemical Formulae

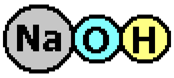
1. 

or **Fe:** 1 atom of iron, 

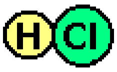
or **S:** 1 atom of sulphur (2Fe would mean two atoms, 5S would mean five sulphur atoms etc.)

2. 

or the formula **FeS** means one atom of iron is chemically combined with 1 atom of sulphur to form the **compound** called **iron sulphide.**

3. 

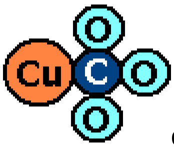
or the formula **NaOH** means 1 atom of sodium is combined with 1 atom of oxygen and 1 atom of hydrogen to form the **compound** called **sodium hydroxide.**

4. 

or the formula **HCl** means 1 atom of hydrogen is combined with 1 atom of chlorine to form 1 molecule of the **compound** called **hydrochloric acid.**

5. 

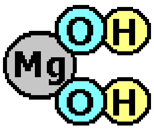
or the formula **$MgCl_2$** means **1 molecule** of the **compound** called **magnesium chloride**, made of one atom of magnesium and two atoms of chlorine.

6. 

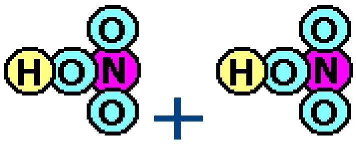
or the formula **$CuCO_3$** means **one molecule** of the **compound** called **copper carbonate**, made up of one atom of copper is combined with one atom of carbon and three atoms of oxygen to form the compound copper carbonate.

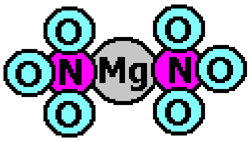
7. 

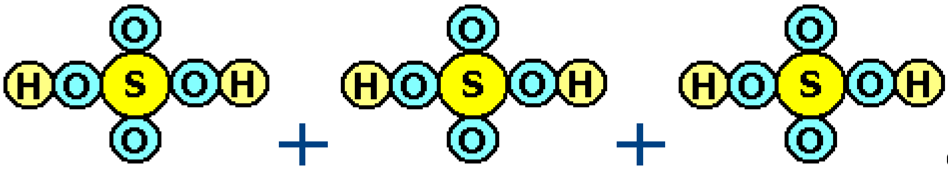
or **$2O_2$** means **two molecules** of the **element** called **oxygen**, and each oxygen molecule consists of two atoms of oxygen.

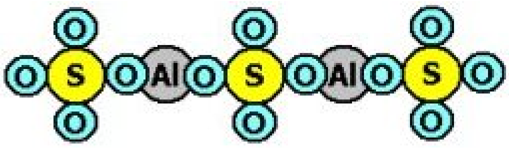
8. 

or the formula **$Mg(OH)_2$** is one molecule of the **compound** magnesium hydroxide made up of one magnesium, two oxygen and two hydrogen atoms.

9.  or **$2HNO_3$** means **two separate molecules** of the **compound nitric acid**, each molecule is made up of one hydrogen atom, one nitrogen atom and three oxygen atoms.

10.  or the formula **$Mg(NO_3)_2$** is one molecule of the **compound magnesium nitrate**, it consists of one magnesium atom, two nitrogen atoms and six oxygen atoms.

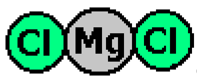
11.  or **$3H_2SO_4$** meaning **three molecules** of the **compound** called **sulphuric acid or hydrogen sulphate** containing 2 atoms of hydrogen, one atom of sulphur and four atoms of oxygen.

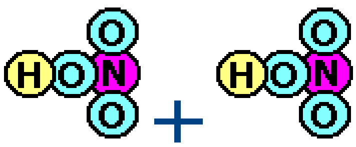
12.  or the formula **$Al_2(SO_4)_3$** means **one molecule** of the **compound** called **aluminium sulphate**, it consists of two aluminium, three sulphur and twelve oxygen atoms.

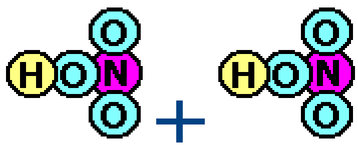
Exercise

Interpret the following formulae

1.  or the formula **CO_2** means

2.  or the formula **$MgCl_2$**

3.  or **$2HNO_3$** means

4.  or **$2HNO_3$**

5. $4Mg(OH)_2$

6. $3(NH_4)_3PO_4$

7. $5Ca(NO_3)_2$

Chemical Equations

A **chemical equation** is a symbolic representation of what happens when chemicals or substances come in contact with one another. In short, a chemical equation is a symbolic representation of a chemical reaction.

For any reaction, what you start with are called the **reactants**, and what you form are called the **products**.

So any **chemical equation** shows in some way the overall chemical change of:

REACTANTS ==> **PRODUCTS**

> means the direction of change from **reactants** to **products**

State symbols used in chemical equations

i. **(g)** means **gas**. Examples of gases include carbon dioxide (CO_2), carbon carbon monoxide (CO), ammonia (NH_3), nitrogen (N_2), hydrogen (H_2), oxygen (O_2), nitrogen dioxide (NO_2), sulphur dioxide (SO_2), etc.

(ii) **(l)** means **liquid**, e.g. water (H_2O).

(iii) **(s)** means **solid.** Examples include all **metals** (e.g. Na, Ca, K, Mg, Pb, Cu, etc.); **metal oxides** (e.g. K_2O, CaO, Na_2O MgO CuO, etc.)

(iv) **(aq)** means **aqueous solution** or **dissolved in water.** Examples include all **acids**[e.g. sulphuric acid (H_2SO_4), hydrochloric acid (HCl), carbonic acid (H_2CO_3), nitric acid (HNO_3), ethanoic acid (CH_3COOH), phosphoric acid (H_3PO_4), etc.]; **metal hydroxides** [e.g. KOH, $Mg(OH)_2$, $NaOH$, etc.]

Balancing Chemical Equations

Balancing chemical equations mean making the number of atoms of each element equal on both sides (reactants and products) of the equation.

Rules for Balancing Chemical Equations

1. The correct reaction must be represented.

2. The correct symbols and valences must be known and from these, the correct formulae of the reaction must be written in the correct places.

3. The number of each kind of atoms of elements be counted and made the same on both sides by changing, where necessary the numbers of molecules taking part or being produced by using large numbers before a chemical symbol or chemical formula.

4. Make sure you do not change any chemical formulae.

Examples

Balance the following equations. Include state symbols.

1. magnesium + hydrochloric acid ==> magnesium chloride + hydrogen

Ans: $Mg_{(s)} + HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$

Number of atoms on the: Reactants: [Mg=1, H=1, Cl=1] products: [Mg=1, Cl=2, H=2]

As can be seen, H and Cl atoms are not equal on both sides of the equation. There is 1 atom of H on the reactants but 2 of them on the products; there is also 1 atom of Cl on the reactants but 2 of them on the products. So there is a need to put 2 in front of the reactant HCl. The equation becomes:

Final answer: $\mathbf{Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}}$

Number of atoms: Reactants: Mg=1, H=2, Cl=2 Products: Mg=1, Cl=2, H=2

Since the numbers of atoms of each element are equal on both sides of the equation, the equation is now balanced.

2. Hydrogen + oxygen $\longrightarrow$ water

Ans: $H_{2(g)} + O_{2(g)} \longrightarrow H_2O(l)$

Atoms: Rectants: H=2, O=2 Products: H=2, O=1

Oxygen atoms on the reactant are 2 and 1 atom on the product. the equation is not balanced. there is need to put a 2 front of H_2O. The equation becomes:

$$H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$$

Now check again; H=2, O=2 $\longrightarrow$ H=4, O=2.

So there is need to put a 2 in front of $H_{2.}$ The balanced equation becomes:

Final answer: $\mathbf{2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}}$

Atoms: Reactants: H=4, O=2 Products: H=4, O=2

3. Aluminium oxide + sulphuric acid ==> aluminium sulphate + water

Ans: $Al_2O_{3(s)} + H_2SO_{4(aq)} \longrightarrow Al_2(SO_4)_{3(aq)} + H_2O_{(l)}$

Atoms: Reactants: Al=2, O=3+4=7; H=2, S=1 Products: Al=2, O=(4x3+1=13), H=2; S=3

The equation is not balanced. We need 3 S on the reactants as we have 3 on the products. Put 3 front of H_2SO_4. The equation becomes:

$$Al_2O_{3(s)} + 3H_2SO_{4(aq)} \longrightarrow Al_2(SO_4)_{3(aq)} + H_2O_{(l)}$$

Check out:Reactants: Al=2, O=3+(3x4)=15, H=6, S=3. Prdcts:Al=2,O=(4x3+1=13),H=2,S=3

6 H on reactants and 2 of them on the product. There is need to put 3 in front of H_2O so as to equal the number of H atoms on both sides of an equation. The balanced chemical equation becomes:

Final answer: $\mathbf{Al_2O_{3(s)} + 3H_2SO_{4(aq)} \longrightarrow Al_2(SO_4)_{3(aq)} + 3H_2O_{(l)}}$

Reactants Al=2, O=15, H=6, S=3 Products Al=2, O=15, H=6, S=3

4. methane + oxygen ==> carbon monoxide + water: $CH_4 + O_2$ ==> $CO + H_2O$

Answer: $\mathbf{2CH_{4(g)} + 3O_{2(g)}}$ **==>** $\mathbf{2CO_{(g)} + 4H_2O_{(l)}}$

5. $NH_3 + O_2 \longrightarrow N_2 + H_2O$

Answer: $\mathbf{4NH_{3(g)} + 3O_{2(g)} \longrightarrow 2N_{2(g)} + 6H_2O(l)}$

6. CuFeS2 + O2 $\longrightarrow$ Cu2S + SO2 + FeO

Answer: $\mathbf{2CuFeS_{2(s)} + 4O_{2(g)} \longrightarrow Cu_2S_{(s)} + 3SO_{2(g)} + 2FeO_{(s)}}$

Exercise

Balance the following equations. Include state symbols.

1. $C_2H_4 + O_2 \longrightarrow CO_2 + H_2O$
2. $Al + O_2 \longrightarrow Al_2O_3$
3. $N_2 + H_2 \longrightarrow NH_3$
4. $Pb(NO_3)_2 \longrightarrow PbO + NO_2 + O_2$
5. $Fe_2O_3 + CO \longrightarrow Fe + CO_2$
6. $Fe + H_2O \longrightarrow Fe_3O_4 + H_2$

Ionic Equations

Ionic equations and net ionic equations are usually written only for reactions that occur in solution and are an attempt to show how the ions present are reacting. While ionic equations show all of the substances present in solution, a net ionic equation shows only those that are changed during the course of the reaction.

The ions that do not take part in a chemical reaction are called '**spectator ions**'**.**

The ionic equation represents the **'actual'** chemical change and **omits the spectator ions.**

Rules for Writing Spectator Ions

i. Balance a chemical equation.
ii. Indicate the correct state symbols.

iii. Split (separate) only those substances that are in aqueous solutions into individual elements and / or radicals. The substances that are in other states, i.e. in solids, liquids and gases are **NEVER** separated.

iv. Re-write the remaining / uncanceled part of an equation and this becomes an ionic equation.

Examples

1. $HCl_{(aq)} + NaOH_{(aq)} ==> NaCl_{(aq)} + H_2O_{(l)}$

Cancel spectator ions as follows.

$H^+{}_{(aq)} + \cancel{Cl^-}_{(aq)} + \cancel{Na^+}_{(aq)} + OH^-{}_{(aq)} ==> \cancel{Na^+}_{(aq)} + \cancel{Cl^-}_{(aq)} + H_2O_{(l)}$

$\mathbf{H^+{}_{(aq)} + OH^-{}_{(aq)} ==> H_2O_{(l)}}$ which is the ionic equation for neutralisation.

The **spectator ions** are chloride (**Cl^-**) and sodium (**Na^+**).

2. $AgNO_{3(aq)} + NaCl_{(aq)} ==> AgCl_{(s)} + NaNO_{3(aq)}$

$Ag^+{}_{(aq)} + \cancel{NO_3^-}_{(aq)} + \cancel{Na^+}_{(aq)} + Cl^-{}_{(aq)} ==> AgCl_{(s)} + \cancel{Na^+}_{(aq)} + \cancel{NO_3^-}_{(aq)}$

so the ionic equation is simply: $\mathbf{Ag^+{}_{(aq)} + Cl^-{}_{(aq)}}$ ==> $\mathbf{AgCl_{(s)}}$

3. $Pb(NO_3)_{2(aq)} + 2KI_{(aq)} ==> PbI_{2(s)} + 2KNO_{3(aq)}$

$Pb^{2+}{}_{(aq)} + 2\cancel{NO_3^-}_{(aq)} + \cancel{2K^+}_{(aq)} + 2I^-{}_{(aq)} ==> PbI_{2(s)} + \cancel{2K^+}_{(aq)} + 2\cancel{NO_3^-}_{(aq)}$

The ionic equation is: $\mathbf{Pb^{2+}{}_{(aq)} + 2I^-{}_{(aq)}}$ ==> $\mathbf{PbI_{2(s)}}$

4. $CaSO_{4(aq)} + 2C_{17}H_{35}COONa_{(aq)} ==> (C_{17}H_{35}COO)_2Ca_{(s)} + Na_2SO_{4(aq)}$

$Ca^{2+}{}_{(aq)} + \cancel{SO_4^{2-}}_{(aq)} + 2C_{17}H_{35}COO^-{}_{(aq)} + 2\cancel{Na^+}_{(aq)} ==> (C_{17}H_{35}COO)_2Ca_{(s)} + 2\cancel{Na^+}_{(aq)} + \cancel{SO_4^{2-}}_{(aq)}$

ionically: $\mathbf{Ca^{2+}{}_{(aq)} + 2C_{17}H_{35}COO^-{}_{(aq)}}$ ==> $\mathbf{(C_{17}H_{35}COO^-)_2Ca^{2+}{}_{(s)}}$

Exercise

Write down the ionic equation of the following.

1. $AgNO_{3(aq)} + NaCl_{(aq)} ==> AgCl_{(s)} + NaNO_{3(aq}$

2. $BaCl_{2(aq)} + AgNO_{3(aq)} ==> AgCl_{(s)} + Ba(NO_3)_{2(aq)}$

3. $Pb(NO_3)_{2(aq)} + KI_{(aq)} ==> PbI_{2(s)} + KNO_{3(aq)}$

4. $Pb(NO_3)_{2(aq)} + 2HCl_{(aq)} ==> PbCl_{2(s)} + 2HNO_{3(aq)}$

Relative Atomic Mass (Ar) and Relative Molecular Mass (Mr)

Relative Atomic Mass (Ar)

This is the number of times that the mass of an atom of an element is greater than $\frac{1}{12}$ the mass of an atom of carbon-12. Note that when we took of relative atomic mass we consider the atomic mass of an element.

The table below shows the relative atomic masses of some elements.

Name of element	**Relative Atomic Mass (Ar)**
Hydrogen	1
Helium	4
Lithium	7
Beryllium	9
Boron	11
Carbon	12
Nitrogen	14
Oxygen	16
Fluorine	19
Neon	20
Sodium	23
Magnesium	24
Aluminium	27
Silicon	28
Phosphorus	31

Sulphur	32
Chlorine	35.5
Argon	40
Potassium	39
Calcium	40

Relative Molecular Mass (Mr)

This is the mass of one molecule of an element or compound compared with $\frac{1}{12}$ the mass of an atom of carbon-12. It is easy to work out the relative molecular mass of an element and compound if their formulae are known. The term relative **molecular mass** is actually the sum of the atomic masses of the atoms of elements in a substance.

Examples

Work out the relative molecular masses of the following.

1. Sodium chloride 2. Carbon dioxide 3. 2H2O 4. $3Al_2(SO_4)_3$

Answers

1. NaCl: Na: 1molecule x 1atom x 23(Ar) = 23

Cl: 1molecule x 1atom x 35.5(Ar) = 35.5+

Mr = 58.5

2. CO_2: C: 1x 1x12 =12

O_2: 1x2x16=32+

Mr=44

3. $2H_2O$: $2H_2$: 2molecules x 2 atoms x 1(Ar) = 04

$2O$: 2 molecules x 1 atom x 16(Ar) =16+

Mr = 18

3. $3Al_2(SO_4)_3$: $3Al_2$: 3 molecules x 2 atoms x 27(Ar) = 162

$3S_3$: 3 molecules x 3 atoms x 32(Ar) = 288

$3O_{(4x3=12)}$:3 molecules x 12 atoms x 16(Ar) = 576+

Mr = 1026

Exercise

Calculate the relative molecular masses of the following.

1. Ammonia 2. Sodium carbonate 3. $4(NH_4)_2SO_4$ 4. $2Ca_3(PO4)_2$

Calculating Masses from Equations and Reactions/Reacting Masses

Steps

i. Write down a balanced chemical equation.
ii. From the balanced chemical equation, pick and relate:
- **a)** the substance being asked; and
- **b)** the substance whose information or mass is given.
- **c)** Relate their reacting masses.

iii. The final answer should have the unit given in the equation.

Examples

1. When carbon is burned in air, carbon dioxide is produced. How much carbon is needed to produce 11g of carbon dioxide?

Answer: $C + O_2 \longrightarrow \mathbf{CO_2}$

C : CO_2

12 : 12 + (2x16)=44

12g : 44g

x : 11 $\frac{\cancel{44}x}{\cancel{44}} = \frac{12X11}{44}$, $x = \mathbf{3g}$

2. Work out the mass of sodium chloride produced when 5.3Kg of sodium carbonate reacts with dilute hydrochloric acid.

Answer: $\mathbf{Na_2CO_3}$+ 2HCl $\longrightarrow$ $\mathbf{2NaCl}$ +CO_2 + H_2O

Na_2CO_3 : 2NaCl

46+12+48 : 46+71

106Kg : 117Kg

5.3Kg : x

$106x = 5.3\ X\ 117$

$\frac{\cancel{106}x}{\cancel{106}} = \frac{620.1}{106}$

x = **5.85Kg**

3. 12.2 tonnes of magnesium reacts completely with oxygen gas. Calculate the mass of oxygen consumed during the reaction and the mass of magnesium oxide produced.

Answer: 2Mg + O_2 $\longrightarrow$ 2MgO

Mass of O_2 consumed

2Mg : O_2

2x24 : 2x16

48 tonnes : 32 tonnes

12.2 tonnes : x

$48x = 12.2 \text{ x } 32$

$\frac{\cancel{48}\text{x}}{\cancel{48}} = \frac{390.4}{48}$ $\text{x} = \mathbf{8.13\ tonnes}$

Mass of MgO produced

2Mg : 2MgO

2x24 : 2x24 + 2x16

48 tonnes : 48 + 32=80 tonnes

12.2 tonnes : x

$48x = 80\text{x } 12.2$

$\frac{\cancel{48}\text{x}}{\cancel{48}} = \frac{80\text{x}12.2}{48}$ $\text{x} = \mathbf{20.33\ tonnes}$

4. Iron(III) oxide reacts with carbon monoxide to form iron and carbon dioxide according to the following chemical equation.

$\mathbf{Fe_2O_3}$ +3CO $\longrightarrow$ **2Fe** + $3CO_2$

Calculate the mass of iron that can be obtained from 80 tonnes of iron (III) oxide.

Answer: Fe_2O_3 : 2Fe

2x56+ 3x16: 2x56

112+ 48 : 112

160 tonnes : 112 tonnes

80 tonnes : x

$$160x = 80X112$$

$$\frac{\cancel{160}x}{\cancel{160}} = \frac{8960}{160} \quad x = \mathbf{56\ tonnes}$$

Exercise

1. Potassium carbonate reacts with dilute sulphuric acid as follows:

$$K_2CO_3 + H_2SO_4 \longrightarrow K_2SO_4 + CO_2 + H_2O$$

Work out the mass of dilute sulphuric acid needed to react with potassium carbonate to form 34 g of potassium sulphate**.**

2. Upon heating, calcium carbonate decomposes to calcium oxide and carbon dioxide. Calculate the weight of calcium oxide produced from the decomposition of 560Kg of calcium carbonate.

3. What mass of magnesium metal reacts completely with 92 tonnes of water vapour according to the reaction below?

$$Mg + 2H_2O \longrightarrow Mg(OH)_2 + H_2$$

4. The equation for the decomposition of hydrogen peroxide is shown below.

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

What mass of oxygen is produced when 17 g of hydrogen peroxide decomposes?

5. Germanium, Ge, is extracted from germanium(IV) oxide by heating with hydrogen. This is the unbalanced chemical equation for the reaction.

$$GeO_2 + \ldots\ldots H_2 \longrightarrow Ge + \ldots\ldots H_2O$$

(a) Balance the above equation.

(b) Calculate the smallest mass of germanium(IV) oxide needed to produce 300 g of germanium by this reaction.

Calculating Moles/Molecules from Equations and Reactions

Steps

i. Write down a balanced chemical equation.

ii. From the balanced chemical equation, pick and relate:

- **d)** the substance being asked; and
- **e)** the substance whose information or mass is given.
- **f)** Relate their mole/ molecule ratios.

Examples

1. Calculate the number of moles of hydrogen needed to combine with 3 moles of oxygen to form water.

Answer: $2H_2 + O_2 \longrightarrow 2H_2O$

$2H_2$: O_2

2 moles : 1 mole

x : 3 moles $x = 2X\,3,\ x = \mathbf{6\ moles\ of\ hydrogen}$

2. Copper (II) nitrate decomposes to copper oxide, nitrite and oxygen according to the following chemical equation.

$$2Cu\,(NO_3)_2 \longrightarrow 2CuO + 4NO_2 + O_2$$

Find the number of molecules of nitrite that could be formed when 0.5 molecules of copper (II) nitrate decomposes.

Answer: $2Cu\,(NO_3)_2$: $4NO_2$

2 moles : 4 moles

x : 0.5 moles $4x = 2X0.5$ $\frac{\cancel{4}x}{\cancel{4}} = \frac{1}{4}$ $x =$ **0. 25 molecules**

3. Calculate oxide reacts with hydrochloric acid to produce calcium chloride and water. Work out the number of moles of hydrochloric acid required react with calcium oxide to produce 4 molecules of water.

Answer: $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$

$2HCl$: H_2O

2 moles : 1 mole

x : 4 moles $x = \mathbf{8}\ \boldsymbol{moles\ of}$ **hydrogen**

4. The decomposition of lead(II) nitrate produces lead (II) oxide, nitrite and oxygen according to the following chemical equation.

$$2Pb(NO_3)_2 \longrightarrow 2PbO + 4NO_2 + O_2$$

Find the number of moles of nitrite that could be formed when 3 moles of lead (II) nitrate decompose.

Answer: $2Pb(NO_3)_2$: $4NO_2$

2 moles : 4 moles

3 moles : x $2x = 3 X 4$ $\frac{\cancel{2}x}{\cancel{2}} = \frac{12}{2}$ x = **6 moles of nitrite**

Exercise

1. Calcium reacts with hydrochloric acid to form calcium chloride and hydrogen gas according to the following chemical equation.

$Ca_{(s)} + 2HCl_{(aq)} ==> CaCl_{2(aq)} + H_{2(g)}$

Calculate the moles of calcium chloride that could be formed if 4 moles of calcium reacted with hydrochloric acid.

2. When iron (III) reacts with sulphuric acid, iron (III) sulphate and hydrogen gas are formed. Work out the number of molecules of sulphuric acid needed to react with 3 moles of iron to form iron (III) sulphate and water.

3. The combination of aluminium and oxygen produces aluminium oxide. The equation for the reaction is shown below.

$4Al_{(s)} + 3O_{2(g)} ==> 2Al_2O_{3(s)}$

Find the number of moles of aluminium oxide produced when 0.25 moles of oxygen combines with aluminium.

4. What could be the number of molecules of sulphur trioxide required to decompose to form 2 molecules of oxygen and sulphur dioxide.

Percentage Composition / Percentage by Mass

This is found by using the following formulae:

1. $\frac{Ar\ of\ an\ element}{Mr\ of\ a\ compound} X\ 100\%$

2. $\frac{Mr\ of\ a\ radical\ or\ molecule\ within\ a\ comppound}{Mr\ of\ the\ whole\ compound} X 100\%$

Examples

1. Find the percentage by mass of oxygen in calcium oxide (CaO).

Answer: $\frac{Ar\ of\ O}{Mr\ of\ CaO}$X100% , $\frac{16}{56}$X100% = **28.57%**

2. Work out the percentage composition of sulphate in magnesium sulphate ($MgSO_4$)

Answer: $\frac{Mr\ of\ SO4}{Mr\ of\ MgSO4}X100\%$ $\frac{96}{120}X100\%$ = **80%**

3. Calculate the % by mass of nitrogen in 2 moles of ammonia ($2NH_3$)

Answer: $\frac{Ar\ of\ 2N}{Mr\ of\ NH3}$X100% , $\frac{28}{34}X100\%$ = **82.35%**

Exercise

Calculate the percentages of the following

1. Na in Na_2SO_4
2. Hydrogen in 3 molecules of water
3. Water in the formula $MgSO4.7H_2O$
4. Nitrate in aluminium nitrate

The Mole

The **mole** is the amount of substance which contains as many elementary entities as there are carbon atoms in 12.000g of the carbon 12 isotope.

The formula used is**: Mole =** $\boldsymbol{\frac{Mass}{Mr\ or\ MM}}$. MM= molar mass. Molar mass is Mr with units. The unit for mole is **mol**.

Examples

1. How many moles of iron are in 20g of iron?

Answer: mole = $\frac{Mass}{Ar}$; mole = $\frac{20\cancel{g}}{56\cancel{g}}$; mole = **0.36 mol**

2. Work out the number of moles of 25g of sodium hydroxide.

Answer: Mole = $\frac{Mass}{Mr\ of\ NaOH}$; Mole = $\frac{25}{40}$; Mole = **0.63 mol**

3. Find the moles contained in 84g of ammonia.

Answer: **Mole** = $\frac{Mass}{Mr\ of\ NH3}$; Mole = $\frac{84g}{17g}$; Mole = **4.94 mol**

- From the formula for mole: mole = $\frac{Mass}{Mr}$, the formula for **MASS** is derived by making mass the subject of the formula as follows:

Mass = mole x Mr

Examples

1. Calculate the mass of 2 moles of carbon dioxide.

Answer: mass = mole x Mr of CO_2; mole = 2 x (12+32); mole =**88g**

2. How many grams of propane (C_3H_8) are there in 0.21 moles of it?

Answer: Mass = mole x Mr; M_r of propane = (3 x 12) + (1 x 8) = 44,

so g of propane = moles x M_r = 0.21 x 44 = **9.24g**

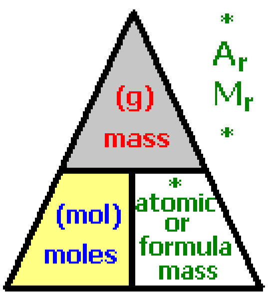
3. Work out the weight of 0.5 moles of calcium carbonate.

Answer. Mass = mole x Mr of $CaCO_3$; Mr of $CaCO_3$: 40+ 12 +48 = 100

Mass = 0.5 x 100; mass = **50g**

- Given Mass and Moles, Mr of a substance can be calculated by making Mr the subject of the formula: Mr = $\frac{Mass}{Mole}$

The triangle below summarises the formula for mole, mass and Mr:



Exercise

1. How many moles of potassium bromide are in 0.25g of potassium bromide?

2. Calculate the moles of calcium chloride contained in 2.5g of calcium chloride?

3. How many grams are in 2 moles of aluminium oxide?

4. Find the mass contained in 0.5 moles of sodium carbonate.

5. Work out the Mr of each one of the following:

- **a)** 128g of 4 moles of substance X
- **b)** 32g of 2 moles of substance Y

Avogadro's Constant Number

Avogadro's Law states that equal volumes of gases (or one mole of gases) under the same conditions of temperature and pressure contain the same number of particles. These particles include molecules atoms and ions.

The Avogadro Constant number = 6.023×10^{23} particles or simply 6×10^{23} particles.

The expression is: **1 mole = 6×10^{23} particles**

Examples

1. How many particles are in 2 moles sulphur dioxide?

Answer: 1 mole = 6 x10^{23} particles

2 moles= x $x = 2 \times 6 \times 10^{23}$ $x =$ **$12\text{x}10^{23}$ particles** or **1.2×10^{24} particles**

2. Find the number of molecules in 3g of hydrogen gas?

Answer: First find the number of moles: mole = $\frac{Mass}{Ar\ of\ H2}$, mole $=\frac{3}{2}$, mole= 1.5 mol

1 mole = 6 x10^{23} molecules

1.5 moles = x $x = 1.5 \times 6 \times 10^{23}$, x = **9×10^{23} molecules**

3. How many atoms are in 3 moles of neon?

Answer: 1mole = 6×10^{23} atoms

2 moles = x $x = 2 \times 6 \times 10^{23}$, $x =$ **12×10^{23} atoms**

Exercise

1. How many molecules are in 4 moles of carbon dioxide?

2. Calculate the number of particles in 3g of chlorine gas?

3. Work out the number of atoms in 0.023 moles of hydrogen chloride.

Molar Volume of Gases

This is the volume occupied by one (1) mole of a given sample of gas at standard conditions of temperature and pressure. These conditions are:

1. **Room temperature and pressure (r.t.p):** At r.t.p 1 mole of a gas occupies the volume of $24dm^3$ ($24,000cm^3$). The expression is **1 mole = $24dm^3$.**
2. **Standard temperature and pressure (s.t.p):** At s.t.p 1 mole of a gas occupies the volume of $22.4dm^3$ ($22,400cm^3$). The expression is: 1 mole = **$22.4dm^3$.**

Examples

1. Find the volume occupied by 4 moles of oxygen gas at r.t.p.

Answer: 1 mole = $24dm^3$

4 moles = x $\quad x = 24 \times 4 \quad x =$ **$96dm^3$**

2. What volume could 6g of hydrogen occupy at s.t.p?

Answer: First find mole: mole = $\frac{Mass}{Ar\ of\ H2}$, mole = $\frac{6}{2}$ = moles

1 mole = $22.4dm^3$

3 moles = x $\quad x = 3 \times 22.4 \quad x =$ **$67.20dm^3$**

3. Work out the number of moles occupied by $34dm^3$ of sulphur dioxide at r.t.p.

Answer: 1 mole = $24dm^3$

$x =$ $34dm^3$ $\quad 24x = 34 \quad \frac{\cancel{24x}}{\cancel{24x}} = \frac{34}{24} \quad x =$ **1.17 moles**

Exercise

1. Calculate the volume of 2.0 moles of nitrite gas at r.t.p.

2. What is the volume occupied by 45g of ammonia at s.t.p?

3. Find the number of moles of $90dm^3$ of nitrogen gas at r.t.p.

4. Work out the moles of 23g of carbon monoxide at s.t.p.

Concentration

Concentration is mole or mass of a solvent in a solution divided by the volume of solute dissolved in a solution. Concentration can be calculated by using the formulae:

1. **Concentration =** $\frac{\boldsymbol{Mole(mol)}}{\boldsymbol{Volume(dm3)}}$ **;** **units are mol/dm^3**

2. **Concentration =** $\frac{\boldsymbol{Mass(g)}}{\boldsymbol{Volume(dm3)}}$ **;** **units are g/dm^3**

NOTE: 1dm^3 = 1000cm^3

Examples

1. What is the concentration of a salt solution if you dissolve 10g of sodium chloride in 250 cm^3 of water?

Answer: Conc = $\frac{\boldsymbol{Mass(g)}}{\boldsymbol{Volume(dm3)}}$, change 250cm^3 to dm^3 : $\frac{250}{1000}$ = 0.25dm^3

Conc = $\frac{10g}{0.25dm3}$ **conc = 40g/dm^3**

2. What mass of the salt is required to make 200 cm^3 of concentration 15g/dm^3?

Conc = $\frac{Mass(g)}{Volume(dm3)}$, mass = conc x vol, mass =15g/$\cancel{dm^3}$ x 0.2$\cancel{dm^3}$ mass **= 3.00g**

3. If you were given 8.0 g of salt, what volume of water,(i) in dm^3 and (ii) in cm^3 should you dissolve it in, to give a salt solution of concentration of 5g/dm^3?

Answer: (i) Conc **=** $\frac{Mass(g)}{Volume(dm3)}$, $= \frac{\cancel{conc}\ X\ Vol}{\cancel{conc}} = \frac{mass}{conc}$, Vol $=\frac{mass}{conc}$, Vol $= \frac{8.0\cancel{g}}{5\cancel{g}/dm3}$

Vol = **1.60dm^3**

(ii) Vol = 1.60 X 1000 = **1,600cm^3**

4. Calculate the concentration of 2 moles 9dm^3 of sugar solution.

Answer: Conc = $\frac{Mole(mol)}{Volume(dm3)}$, conc = $\frac{2mol}{9dm3}$, conc = **0.22mo/dm^3**

5) 5.95g of potassium bromide was dissolved in 400cm^3 of water. Calculate its molarity.

Answer: Molarity is expressed in mol/dm^3. Molarity = $\frac{Mole(mol)}{Volume(dm3)}$,

find moles = $\frac{Mass}{Mr\ of\ KBr}$, mole = $\frac{5.95}{119}$ mole = 0.05mol, therefore conc = $\frac{0.05mol}{0.40dm3}$;

conc = **0.13mol/dm3** or conc = **0.13M**

Exercise

1. What is the concentration of 5g in 1.50cm^3 of sodium chloride (NaCl):

   (a) in g/dm^3 and

   (b) in mol/dm^3

2. How many moles of H_2SO_4 are there in 250cm^3 of a 0.800 mol dm^{-3} (0.8M) sulphuric acid solution?

3. A solution of calcium sulphate ($CaSO_4$) contains 0.500g dissolved in 2.00 dm^3 of water. Calculate:

   (a) The concentration in g/dm^3

   (b) The concentration in mol/dm^3

   (c) The molarity

4. What mass of sodium hydroxide is needed to make up 0.500 dm^3 of a 0.500 mol dm^{-3} solution?

NOTE: The following formulae can also be used to calculate concentration/molarity and volume of a base and an acid:

1. $\mathbf{M_aV_a = M_bV_b}$ Used when the mole ratio of an **acid** to a **base** is **1:1**

   E.g $HCl + NaOH \longrightarrow NaCl + H_2O$

2. $\frac{MaVa}{MbMv} = \frac{1}{2}$ Used when the mole ratio of an **acid** to a **base** is **1:2**

   E.g $H_2SO_4 + KOH \longrightarrow Na_2SO_4 + H_2O$

3. $\frac{MaVa}{MbVb} = \frac{2}{1}$ Used when the mole ratio of an **acid** to a **base** is **2:1**

   E.g $2HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2H_2O$

   - Where (in all the three formulae above):**Ma** = Molarity of an acid, **Va** = Volume of an acid, **Mb** = Molarity of a base, **Vb** = Volume of a base

Examples

1. Calculate the morality of $0.50dm^3$ of hydrochloric acid required to neutralise $10mol/dm^3$ of $5dm^3$ sodiuim hydroxide?

Ans

Step 1: Write down a balanced chemical equation for the neutralisation reaction:

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

Step 2: Identify acid to base mole ratio: $\frac{\text{acid}}{1} : \frac{\text{base}}{1}$

Step 3: Chose the appropriate formula to use: **MaVa = MbVb**

Step 4: Collect data: Ma = ? Va = $0.50dm^3$ Mb = $10mol/dm^3$ Vb = $5dm^3$

Step 5: Substitute in the values/ carry out the calculation:

$0.50Ma = 10 \times 5$, $0.50Ma = 50$, $\frac{\cancel{0.50}Ma}{\cancel{0.50}} = \frac{50}{0.50}$, $Ma = \frac{50}{0.50}$, Ma = ***<u>100mol/dm³</u>***

2. What volume of 34M potassium hydroxide would be needed to react with $62cm^3$ of 97M sulphuric acid?

Ans

Follow the steps in example 1 above.

$H_2SO_4 + 2KOH \longrightarrow K_2SO_4 + 2H_2O$ $\quad \frac{Acid}{1} : \frac{Base}{2}$

$\frac{MaVa}{MbVb} = \frac{1}{2}$ $\quad$ Ma = 97M, Va = $62cm^3$, Mb = 34M, Vb?

$\frac{MaVa}{MbVb} \frac{1}{2}$ $\quad$ 2(MaVa) = MbVb, $2(97 \times 62) = 34Vb$, $2(6014) = 34Vb$, $\frac{12028}{34} = \frac{\cancel{34}Vb}{\cancel{34}}$,

Vb = ***<u>353.76cm³ or 0.35dm³</u>***

3. Work out the concentration of $500cm^3$ of calcium hydroxide solution needed to neutralise $76cm^3$ of $23mol/dm^3$ nitric acid.

Ans

$2HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2H_2O$ $\quad \frac{Acid}{2} : \frac{Base}{1}$

$M_a = 23mol/dm^3$, $V_a = 76cm^3$, $M_b = ?$ $V_b = 500cm^3$

$\frac{MaVa}{MbVb} = \frac{2}{1}$, $\frac{MaVa}{MbVb} \frac{2}{1}$, $2(M_bV_b) = M_aV_a$, $2(500M_b) = 23 \times 76$, $1000M_b = 1748$

$\frac{\cancel{1000}Mb}{\cancel{1000}} = \frac{1748}{1000}$, Mb = ***<u>1.75mol/dm</u>^{3}***

Exercise

1) 10dm^3 of hydrochloric acid (HCl) is needed to neutralise 0.34dm^3 of 8M sodium hydroxide (NaOH). Calculate the molarity of the acid.

2) What volume of 0.2 mol/dm^3 of sulphuric acid is required to neutralise 20cm^3 of 0.3 mol/dm^3 potassium hydroxide?

Empirical Formula

This is the simplest formula that shows the ratio of the number of atoms of elements in a compound.

Steps for Calculating Empirical Formula

1. Find ratios: divide the Ar of each element into its percentage (percentage by mass).

2. Find atoms: divide all the ratios by the smallest ratio.

3. Finally write down the compound with the number of atoms of elements.

Molecular Formula

This is the formula that shows the actual number of of atoms of elements present in a compound. The formula used is:

Molecular formula: (Empirical formula)n. where n = number of moles

Examples

1. A compound is found to contain 23.3% magnesium, 30.7% sulphur and 46.0% oxygen. What is the empirical formula of this compound?

Answer: Mg = $\frac{23.3}{24}$, Mg = 0.97 S = $\frac{30.7}{32}$, S = 0.96 O = $\frac{46.0}{16}$, O = 2.88

Mg = $\frac{0.97}{0.96}$, Mg = 1 S = $\frac{0.96}{0.96}$, S = 1 O = $\frac{2.88}{0.96}$, O = 3

Empirical formula is **$MgSO_3$**

2. What is the empirical formula for a compound containing 38.8% carbon, 16.2% hydrogen and 45.1% nitrogen?

Answer: C = $\frac{38.8}{12}$, C = 3.23 H = $\frac{16.2}{1}$, H = 16.20 N = $\frac{45.1}{14}$, N = 3.22

C= $\frac{3.23}{3.22}$, C = 1 H = $\frac{16.20}{3.22}$, H = 5.03 = 5 N = $\frac{3.22}{3.22}$, N = 1

Empirical formula is $\mathbf{CH_5N}$

3. A sample of an oxide of nitrogen is found to contain 30.4% nitrogen. Its molecular mass is 92.

(a) What is its empirical formula?

(b) What is its molecular formula ?

Answer: (a) N = $\frac{30.4}{14}$, N = 2.17 O = 100% - 30.4 (for N) = $\frac{69.6}{16}$, O = 4.35

N = $\frac{2.17}{2.17}$, N = 1 O = $\frac{4.35}{2.17}$, 2.00 = 2

Empirical formula is $\mathbf{NO_2}$

(b) Molecular formula = (Empirical formula)n

n = $\frac{Molecular\ mass}{Mr\ of\ NO2}$, n = $\frac{92}{46}$, n = 2

Molecular formula is $(NO_2)2$ = $\mathbf{N_2O_4}$

4. A compound contains 32% of carbon, 4% of hydrogen and 64% of oxygen. If its molecular mass is 150, calculate its:

(a) empirical formula

(b) molecular formula

Answer: (a) C= $\frac{32}{12}$, C= 2.67 H = $\frac{4}{1}$, H = 4 O = $\frac{64}{16}$, O = 4

C = $\frac{2.67}{2.67}$, C = 1 H= $\frac{4}{2.67}$, H = 1.5 = 2 O = $\frac{4}{2.67}$, O =1.5 = 2

Empirical formula: $\mathbf{CH_2O_2}$

(b) Molecular formula = (impirical formula)n

n = $\frac{mass}{MrofCH2O2}$, n = $\frac{150}{46}$, n = 3.26 = 3

Molecular formula = $(CH_2O_2)3$ = $\mathbf{C_3H_6O_6}$

NOTE THAT the empirical formula **MUST** be left in the **LOWEST TERMS**.

e.g. (a) $C_2H_4O_4$ should be CH_2O_2 (b) B_2M_2 should be BM

Molecular formula should **NEVER** be left in the lowest terms.

Exercise

1. A sample of an oxide of arsenic is found to contain 75.74% arsenic. What is its empirical formula?

2. What is the empirical formula for a compound containing 26.57% potassium, 35.36% chromium, and 38.07% oxygen?

3. What is the empirical formula of a compound comprised of 1.8% hydrogen, 56.1% sulfur and 42.1% oxygen?

4. A borane is a compound containing only boron and hydrogen. If a borane is found to contain 88.45% boron. Its molecular mass is 177. Calculate its:

(a) empirical formula

(b) molecular formula

5. A compound contains 75% carbon and 25% hydrogen. Its molecular mass is 80. Calculate its:

(a) empirical formula

(b) Molecular formula

ACIDS, BASES ABD SALTS

ACIDS

Acids are substances that when dissolved in water form hydrogen ions [$H_{(aq)}^+$] as the only positively charged ions e.g.

- hydrochloric acid **HCl** gives **$H^+_{(aq)}$ and $Cl^-_{(aq)}$ ions** in water (aqueous solution)
- sulphuric/sulphuric acid **H_2SO_4** gives **$2H^+_{(aq)}$ and $SO_4^{2-}{}_{(aq)}$ ions** in water (aqueous solution)
- nitric acid **HNO_3** gives **$H^+_{(aq)}$ and $NO_3^-{}_{(aq)}$ ions** in water (aqueous solution)
- ethanoic acid **CH_3COOH** gives **H^+**$_{(aq)}$ and **$CH_3COO^-{}_{(aq)}$ ions** (aqueous solution)

Acids can also be defined as proton donors.

The table below shows the names of common acids and their formulae.

Name of Acid	Formula of Acid
Sulphuric acid	H_2SO_4
Sulphurous acid	H_2SO_3
Hydrochloric acid	HCl
Carbonic acid	H_2CO_3
Nitric acid	HNO_3
Ethanoic acid	CH_3COOH
Phosphoric acid	H_3PO_4

Basicity of an Acid

This is the number of hydrogen ions that can be formed from one molecule of an acid. Basicities of acids are shown below.

e.g. $HNO_{3(aq)} \longrightarrow H^+_{(aq)} + NO^-_{3(aq)}$: number of hydrogen ions formed is 1. therefore, the basicity is1.

$H_2SO_{4(aq)} \longrightarrow 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$:number of hydrogen formed is 2. Therefor, basicity is 2.

Name of Acid	Formula of Acid	Basicity of Acid
Sulphuric acid	H_2SO_4	2 = di basic
Sulphurous acid	H_2SO_3	2 = di basic
Hydrochloric acid	HCl	1 = mono basic
Carbonic acid	H_2CO_3	2 = di basic
Nitric acid	HNO_3	1 = mono basic
Ethanoic acid	CH_3COOH	1 = mono basic
Phosphoric acid	H_3PO_4	3 = tri basic

Physical Properties of Acids

1. Turn blue litmus paper red.
2. Have sour taste.
3. Have a strong chocking/unpleasant smell.
4. Are corrosive, i.e. they eat away surfaces of substances and can even destroy the entire substance.
5. Are clear colourless solutions.

Chemical Properties of Acids

1. React with bases or alkalis to form salt and water only. This reaction by which acids react with alkalis to form salt and water only is called **neutralisation reaction**
   e.g $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

$H_2SO_{4(aq)} + 2KOH_{(aq)} \longrightarrow K_2SO_4 + 2H_2O_{(l)}$

$CuO_{(s)} + H_2SO_{4(aq)} \longrightarrow CuSO_{4(aq)} + H_2O_{(l)}$

NOTE THAT **Neutralisation reaction ionically is: $H^+_{(aq)} + OH^-_{(aq}$**

2. Acids react with metals to form salt and hydrogen gas.
   e.g. $Zn_{(s)} + 2HCl_{(aq)} \Longrightarrow ZnCl_{2(aq)} + H_{2(g)}$
   $Mg_{(s)} + H_2SO_{4(aq)} \Longrightarrow MgSO_{4(aq)} + H_{2(g)}$

3. Acids react with carbonates and hydrogen carbonates to form salt, carbon dioxide and water.
   e.g. $CaCO_{3(s)} + 2HCl_{(aq)} \Longrightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$
   $CuCO_{3(s)} + 2HCl_{(aq)} \Longrightarrow CuCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$
   $NaHCO_{3(s)} + HCl_{(aq)} \Longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Weak and Strong Acids

A **strong acid** is one that is ionised completely in water (or in aqueous solution).

Examples of strong acids are hydrochloric, sulphuric and nitric acids.

- nitric acid is: $HNO_{3(aq)} \Longrightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$
- hydrochloric acid is: $HCl_{(aq)} \Longrightarrow H^+_{(aq)} + Cl^-{}_{(aq}$
- and sulphuric acid is: $H_2SO_{4(aq)} \Longrightarrow 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

A **weak acid** is one that ionises partially in aqueous solution.

Examples of weak acids are ethanoic, phosphorous and carbonic acids. The ionisation (forming ions) of weak acids is reversible. This is indicated by both forward and backward arrows as shown below.

$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-{}_{(aq)} + H^+_{(aq)}$

$H_2CO_{3(aq)} \rightleftharpoons 2H^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

Indicators

Indicators are substances that show a particular colour in acidic or alkaline medium. The table below summarises the names of indicators and their colour changes in acidic and alkaline media.

Substance to be tested	Indicators			
	Blue litmus paper	Red litmus paper	Methyl orange	Phenolphthalein
Acid	Blue to red	No effect	Orange to pink	Remains colourless
Alkaline/Base	No effect	Red to blue	Orange to	Colourless to

			yellow	pink
Neutral solution	No effect	No effect	No colour change/remains orange	No colour change/remains colourless

Bases and Alkalis

Bases are substances that contain oxides (O^{2-}) and hydroxide ions (OH^-).

Examples of bases include:

i. sodium hydroxide ($NaOH$)
ii. potassium hydroxide (KOH)
iii. copper (II) oxide (CuO)
iv. magnesium hydroxide [$Mg(OH)_2$]
v. calcium oxide (CaO)
vi. Ammonium hydroxide (NH_4OH)
vii. Lithium hydroxide ($LiOH$)

Bases are actually metal hydroxides and metal oxides.

Some bases are soluble in water while others are insoluble. Soluble bases are called **alkalis.** Alkalis are substances which dissolved in water produce hydroxide ions (OH^-).

Examples of alkalis include:

i. sodium hydroxide ($NaOH$)
ii. potassium hydroxide (KOH)
iii. lithium hydroxisde ($LiOH$)
iv. ammonium hydroxide (NH_4OH)

Physical Properties Bases/Alkalis

1. Turn red litmus paper blue
2. Have bitter taste
3. Feel breezy or soapy when dissolved in water

Chemical Properties of Bases/Alkalis

1. React with acids to form salt and water only. This is neutralisation reaction. Refer to chemical property 1 of acids.
2. React with ammonium salts to form salt, water and ammonia gas.
   e.g $NH_4Cl + NaOH \longrightarrow NaCl + H_2O + NH_3$

Weak and Strong Alkalis

A **strong alkali** is one that is ionised completely in water (or in aqueous solution).

Examples of strong alkalis (soluble strong bases) are sodium hydroxide, potassium hydroxide etc. (usually Group 1 or 2 hydroxides).

- e.g. the maximum (or nearly) hydroxide ion concentration results in the highest pH ...
- potassium hydroxide is: $KOH_{(aq)}$ ==> $K^{+}_{(aq)} + OH^{-}_{(aq)}$
- sodium hydroxide is: $NaOH_{(aq)}$ ==> $Na^{+}_{(aq)} + OH\text{-}_{(aq)}$
- or strontium hydroxide is: $Sr(OH)_{2(aq)}$==> $Sr^{2+}_{(aq)} + 2OH^{-}_{(aq)}$

A **weak alkali** is one that ionises partially in aqueous solution.

An example of a weak alkali is ammonium hydroxide (or ammonia solution: NH_4OH).

Uses of some Alkaline Substances

Agricultural Uses

i. Ash and lime are used to neutralise the acid in the soil.
ii. Ash and lime are also used as fertilizers.

Industrial Uses

i. Sodium hydroxide (caustic soda) is used in reaction with fatty acids in the chemikcal industry to make soap.
ii. Sodium carbonate (washing soda) is used to soften hard water and to remove grease.
iii. Sodium carbonate is also used in the manufacture of glass and soap making.

PH Scale and Classification of Oxides

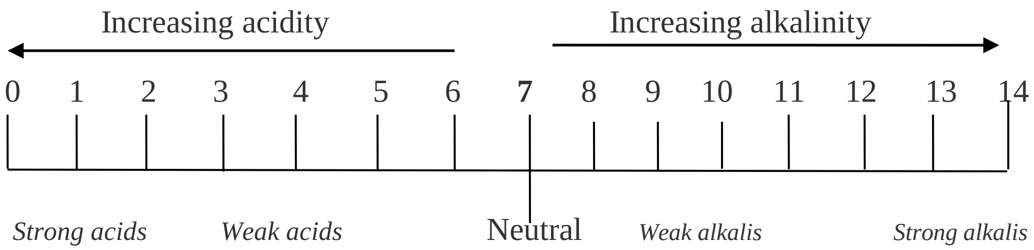
PH stands for Potential of Hydrogen.

The PH scale is the measure of the degree of acidity and alkalinity of a substance. The degree of acidity and alkalinity can be measured by the PH scale using the universal indicator. Universal indicator is a mixture of colours. It is available in paper and solution form. It undergoes various colour changes depending on the degree of acidity or alkalinity of a substance. The table below shows the colours of universal indicator at different PH values.

PH Value	Colour of Universal Indicator
0-2	**Red**
3	**Pink**
4	**Brown**
5	**Yellow**
6-8	**Green**
9-10	**Blue**
11 – 12	**Indigo**

13 -14	**Violet**

The PH scale consists of numbers ranging from 0 to 14. The figure below shows the PH scale.



These PH numbers are used to compare the strengths of acids and alkalis.

When a solution is neutral the PH value is 7. Pure water for example has a PH value of 7. When a solution is acidity, the PH value is less than 7; and when a solution is alkalinity the PH value is greater than 7. In short, all acids have got PH values less than 7 and all alkalis have got PH values greater than 7.

- The smaller the PH value of an acid the stronger it is and vice versa. E.g. an acid with PH value of 3 is stronger than the one with PH value of 6.
- The smaller the PH value of an alkali the weaker it is and vice versa. E.g. an alkali with PH value of 8 is weaker than the one with PH value of 14.

The Importance of PH Value in Agriculture

Agriculturalists and good farmers are interested in knowing which plants grow well in which type of soil. Some plants grow well in acid soil while others grow well in alkaline soil. Taking the PH measurement of the soil will show the alkalinity or acidity of the soil. If for example, a soil is discovered to be too acidic for the good growth of a plant, some lime is applied so as to neutralise the acid in the soil. Equally, if the soil is found to be too alkalinity for the good growth of a plant, then a fertilizer such as ammonium sulphate is applied to improve the condition of the soil.

Oxides

An **oxide** is a compound of oxygen and another element.

Classification of Oxides

Oxides can be classified as either **Acidic, Basic, Amphoteric or Neutral**

Nature of the oxides of the elements **across the periodic table from left to right** changes from **Basic** $\longrightarrow$ **Amphoteric** $\longrightarrow$ **Acidic**

Acidic Oxides

These are oxides of non- metals. They are known as acidic oxides because they form acids when dissolved in water.

Examples:

SO_2, SO_3, CO_2, NO_2

$SO_{2(g)} + H_2O_{(l)} \longrightarrow H_2SO_{39(aq)}$

$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$

$CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$

$NO_{2(g)} + H_2O_{(l)} \longrightarrow H_2NO_{3(aq)}$

Properties:

1. Do not react with acids.
2. React with bases and alkalis to form salt & water.
3. Dissolve in water to form acidic solutions.
4. Usually gases at room temp.

Basic Oxides

Basic oxides are oxides of metals

Examples:

Na_2O, CaO, MgO, FeO, CuO

Properties:

1. Do not react with bases.
2. React with acids to form salt & water.
3. Basic Oxides are usually insoluble in water. Those that dissolve in water form alkaline solutions.

Neutral Oxides

Examples:

CO, NO, H_2O

Properties:

Neutral pH

Amphoteric Oxides

These are oxides of semi – metals/metalloids.

Examples:

Oxides formed with metals near and non metals such as ZnO, Al_2O_3, PbO, Ga_2O_3

Properties:

1. They have the characteristic properties of acidic and basic oxides, therefore, react with both acids and bases to form salt & water.

SALTS

A **salt** is a compound consisting of positive metallic ion(s) and negative non-metallic ion(s) derived from an acid. Salts are formed when hydrogen ion(s) are replaced wholly (completely) or partially (partly) by other cations (e.g the metal ion or ammonium ion).

Salts can either be **normal** or **acid salts** depending on the displacement of hydrogen ions.

Normal Salts are salts formed when **all** the hydrogen ions are displaced by other cations. E.g

(i) sodium chloride: $NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$: $\mathbf{(Na^+Cl^-)}$

(ii) potassium nitrate: $HNO_{3(aq)} + K_{(S)} \longrightarrow KN0_{3(aq)} + H_{2(g)}$: $\mathbf{(K^+NO_3^-)}$

(iii) sodium carbonate: $Na_{(s)} + H_2CO_{3(aq)} \longrightarrow Na_2CO_{3(aq)} + H_{2(g)}$: $\mathbf{(2Na^+\ CO_2^{2-})}$

Acid salts are salts formed when one or more hydrogen ion(s) are attached to the negatively charged group of a salt. E.g (i) potassium hydrogen sulphate: $\mathbf{(K^+HSO_4^{2-})}$

(ii) calcium hydrogen carbonate: $\mathbf{(Ca^+HCO_3^{2+})}$

Preparation of salts

There are several methods of preparing salts. The method chosen for preparing a salt entirely depends on the **solubility** of that particular salt. The table below shows the classification of salts into soluble and insoluble salt:

Soluble salts	**Insoluble salts**
All Sodium, Potassium and ammonium	
All Chlorides	Except silver and lead, chlorides
All Ethanoates	
All Sulphate	Except calcium, barium and lead sulphates
All Nitrates	
Carbonates and hydroxides of Sodium, potassium and ammonium	All other carbonates and hydroxides are insoluble

Soluble salts are prepared either by **crystallization** (with an insoluble base) or **titration** (with a soluble base).

Insoluble salts are prepared by **precipitation** (also refers to as **double decomposition**).

<u>Preparation of soluble salts</u>

There are **four ways** of preparing soluble salts. These are the same as the typical reactions of acids:

(a) acid + insoluble base $\longrightarrow$ salt + water

(b) acid + metal $\longrightarrow$ salt + hydrogen

(c) acid + soluble base $\longrightarrow$ salt + water

(d) acid + carbonate $\longrightarrow$ salt + carbon dioxide + water

(a) By using a Dilute Acid and an Insoluble Base

E.g dilute sulphuric acid and copper(II) oxide

Steps

(i) Reaction

Excess copper(II) oxide powder is added to **warm** dilute sulphuric acid.

- **Excess** copper (II) oxide is added to the acid so as to neutralise all the acid and form a saturated solution required for the formation of crystals
- **Warm** acid is used so as speed up the reaction

The chemical equation for the reaction is $H_2SO_{4(aq)} + CuO_{(s)} \longrightarrow CuSO4_{(aq)} + H_2O_{(l)}$

(ii) Filtration

When the reaction is complete, the excess (undissolved) copper (II) oxide is filtered of from the copper (II) oxide solution.

(iii) Evaporation

using water bath method, evaporate s**ome** of the water. Some water should remain for crystallization. **Precaution**: do not heat the solution to dryness or else there would be no water for cyrstallisation.

(iv) Crystallisation

The solution is left (put) to cool slowly in a beaker at ordinary temperature in a clean environment. Blue crystals of copper (II) oxide start to form.

(v) Clean (rinse) the crystals

Wash the crystals two or three times with small quantities of cold distilled water so as to remove surface solution from the crystals.

(vi) Dry the crystals

Put the washed crystals in between filter papers or in a porous plate.

(b) By using a dilute acid and a metal

Note that this method is suitable for fairly reactive metals like magnesium, zinc, and iron:

- For more reactive metals, the reaction is too vigorous
- For least reactive metals, these metals can not displace hydrogen from an acid, hence no reaction will occur.

Steps (follow steps in (a) above

(c) By Using a Dilute Acid and Metal Carbonate

Metal carbonates will react with acids to form salt, carbon oxide and water.

Steps (follow steps in (a) above

(d) By Using a dilute acid and a Soluble Base/Alkali (Acid – Base Titration)

E.g Hydrochloric acid (HCl) and sodium hydroxide (NaOH)

This reaction needs special technique since both reactants are soluble, and so, an indicator must be used to show the required quantities of reactants of an acid an alkali.

Procedure/Steps

(1) Using a pipette measure 25cm^3 of sodium hydroxide solution into a conical flask. Drop about two drops of an indicator (ie, phenolphthalein). The solution becomes **pink** in colour. Universal indicator can also be used, it turns **green** in acidic solution. Note that phenolphthalein is **colourless** in neutral and acidic solutions but **pink** in alkaline solution.

(2) The acid is added to the 25cm^3 of the base little at time while swelling/shaking the flask to let the acid and base mix evenly. When the pink colour disappears (in the case of phenolphthalein) or when the universal indicator turns green, stop adding the acid. This indicates that all the acid has been neutralized by the base.

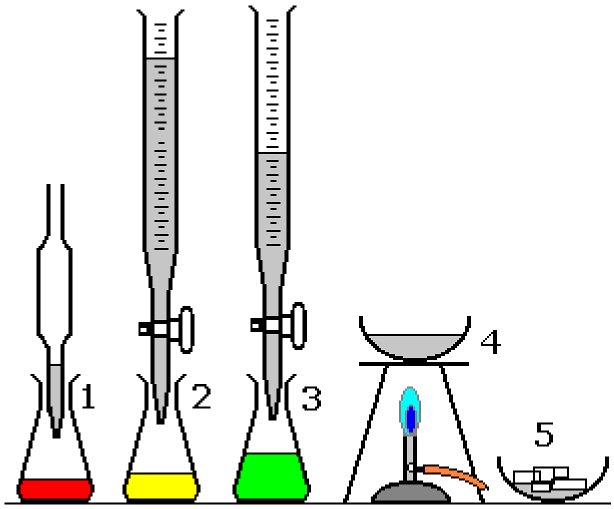
(3) Take note of the amount of the acid used. This is called the endpoint volume.

(1) to (3) are repeated with both known volumes mixed together BUT without an indicator (phenolphthalein or universal). An indicator would make the slat impure.

(4) The solution is transferred to an evaporating dish and heated to partially evaporate the water causing crystallisation or can be left to slowly evaporate - which tends to give bigger and better crystals.

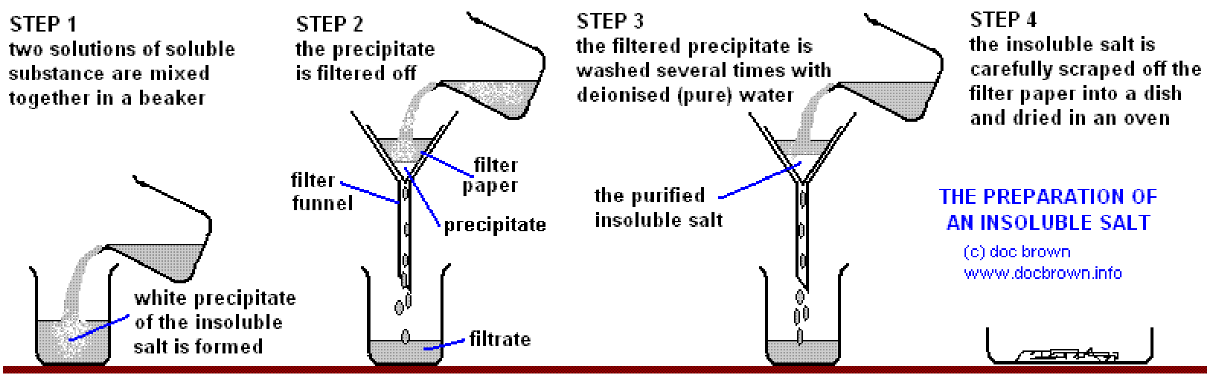
(5) The residual liquid can be decanted away and the crystals can be carefully collected and dried by 'dabbing' with a filter paper OR the crystals can be collected by filtration.

The diagrams for the steps/procedure above are shown below:



Preparation of Insoluble Salts

Insoluble salts are prepared by **precipitation or double decomposition**.



- An insoluble salt can be made by mixing two solutions of soluble salts in a process is called precipitation. A **precipitate** is an insoluble solid formed when a chemical reaction occurs between two dissolved ionic substances.

- The method is quite simple - illustrated above, assuming in this case the insoluble salt is colourless-white.
- One solution contains the 1st required ion, and the other solution contains the 2nd required ion.
- The two solutions of SOLUBLE compounds are mixed together so the INSOLUBLE salt precipitate is formed.
- The precipitated salt can then be filtered off with a filter funnel and paper.
- The collected solid is washed with distilled water to remove any remaining soluble salt impurities and carefully removed from the filter paper to be dried e.g. left out in dry room or warmed in a pre-heated oven.

- Examples ...
  - (i) Silver chloride is made by mixing solutions of solutions of silver nitrate and sodium chloride.
    - silver nitrate + sodium chloride ==> silver chloride + sodium nitrate
    - $AgNO_{3(aq)} + NaCl_{(aq)}$ ==> $AgCl_{(s)} + NaNO_{3(aq)}$
    - in terms of ions it could be written as
    - $Ag^{+}NO_3^{-}{}_{(aq)} + Na^{+}Cl^{-}{}_{(aq)}$ ==> $AgCl_{(s)} + Na^{+}NO_3^{-}{}_{(aq)}$
    - or: $Ag^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)} + Na^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$ ==> $AgCl_{(s)} + Na^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)}$
    - but the spectator ions are nitrate NO_3^- and sodium Na^+ which do not change at all,
    - so the ionic equation is simply: $Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$ ==> $AgCl_{(s)}$
      - Note that ionic equations omit ions that do not change there chemical or physical state.
      - In this case the nitrate, $NO_3^{-}{}_{(aq)}$ and sodium $Na^{+}{}_{(aq)}$ ions do not change physically or chemically and are called spectator ions,
      - BUT the aqueous silver ion, $Ag^{+}{}_{(aq)}$, combines with the aqueous chloride ion, $Cl^{-}{}_{(aq)}$, to form the insoluble salt silver chloride, $AgCl_{(s)}$, thereby changing their states both chemically and physically.
    - If you use barium chloride the word and symbol equations are ...
    - barium chloride + silver nitrate ==> silver chloride + barium nitrate
    - $BaCl_{2(aq)} + 2AgNO_{3(aq)}$ ==> $2AgCl_{(s)} + Ba(NO_3)_{2(aq)}$
    - which can be written as

- $Ba^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + 2Ag^{+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)} \Rightarrow 2AgCl_{(s)} + Ba^{2+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)}$
- the spectator ions are Ba^{2+} and NO_3^{-}
- so the ionic equation is: $Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)} \Rightarrow AgCl_{(s)}$

- (ii) Lead(II) iodide, a yellow precipitate (insoluble in water!) can be made by mixing lead(II) nitrate solution with e.g. potassium iodide solution.
  - lead(II) nitrate + potassium iodide ==> lead(II) iodide + potassium nitrate
  - $Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \Rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$
  - which can be written as
  - $Pb^{2+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)} + 2K^{+}{}_{(aq)} + 2I^{-}{}_{(aq)} \Rightarrow PbI_{2(s)} + 2K^{+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)}$
  - the ionic equation is: $Pb^{2+}{}_{(aq)} + 2I^{-}{}_{(aq)} \Rightarrow PbI_{2(s)}$
  - because the spectator ions are nitrate NO_3^{-} and potassium K^{+}.
  - In a similar way you can make lead(II) chloride by e.g. using dilute hydrochloric acid
    - lead(II) nitrate + hydrochloric acid ==> lead(II) chloride + nitric acid
    - $Pb(NO_3)_{2(aq)} + 2HCl_{(aq)} \Rightarrow PbCl_{2(s)} + 2HNO_{3(aq)}$
    - $Pb^{2+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)} + 2H^{+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} \Rightarrow PbCl_{2(s)} + 2H^{+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)}$
    - the ionic equation is: $Pb^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} \Rightarrow PbCl_{2(s)}$
    - because the spectator ions are nitrate NO_3^{-} and hydrogen H^{+}.
  - and you can make lead(II) bromide by e.g. using sodium bromide
    - lead(II) nitrate + sodium bromide ==> lead(II) bromide + sodium nitrate
    - $Pb(NO_3)_{2(aq)} + 2NaBr_{(aq)} \Rightarrow PbBr_{2(s)} + 2NaNO_{3(aq)}$
    - $Pb^{2+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)} + 2Na^{+}{}_{(aq)} + 2Br^{-}{}_{(aq)} \Rightarrow PbBr_{2(s)} + 2Na^{+}{}_{(aq)} + 2NO_3^{-}{}_{(aq)}$
    - the ionic equation is: $Pb^{2+}{}_{(aq)} + 2Br^{-}{}_{(aq)} \Rightarrow PbBr_{2(s)}$
    - because the spectator ions are nitrate NO_3^{-} and sodium Na^{+}.
- (iii) Calcium carbonate, a white precipitate, forms on e.g. mixing calcium chloride and sodium carbonate solutions ...
  - calcium chloride + sodium carbonate ==> calcium carbonate + sodium chloride

- $CaCl_{2(aq)} + Na_2CO_{3(aq)}$ ==> $CaCO_{3(s)} + 2NaCl_{(aq)}$
- $Ca^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2Na^+_{(aq)} + CO_3^{2-}_{(aq)}$ ==> $CaCO_{3(s)} + 2Na^+_{(aq)} + 2Cl^-_{(aq)}$
- ionically: $Ca^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$ ==> $CaCO_{3(s)}$
- because the spectator ions are chloride Cl^- and sodium Na^+.

- (iv) Barium sulphate, a white precipitate, forms on mixing e.g. barium chloride and dilute sulphuric acid ...
  - barium chloride + sulphuric acid ==> barium sulphate + hydrochloric acid
  - $BaCl_{2(aq)} + H_2SO_{4(aq)}$ ==> $BaSO_{4(s)} + 2HCl_{(aq)}$
  - $Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2H^+_{(aq)} + SO_4^{2-}_{(aq)}$ ==> $BaSO_{4(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)}$
  - ionic equation: $Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$ ==> $BaSO_{4(s)}$
  - because the spectator ions are chloride Cl^- and hydrogen H^+.
    - Or you can use sulphate salts like sodium sulphate, so the word and symbol equations are ..
    - barium chloride + sodium sulfate ==> barium sulfate + sodium chloride
    - $BaCl_{2(aq)} + Na_2SO_{4(aq)}$ ==> $BaSO_{4(s)} + 2NaCl_{(aq)}$
    - The ionic equation is the same: $Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$ ==> $BaSO_{4(s)}$
    - because the spectator ions are sodium Na^+ and chloride Cl^-
- (v) Lead(II) sulphate, a white precipitate, forms in a similar way e.g.
  - lead(II) nitrate + sodium sulphate ==> lead(II) sulphate + sodium nitrate
  - $Pb(NO_3)_{2\ (aq)} + Na_2SO_{4\ (aq)}$ ==> $PbSO_{4\ (s)} + 2NaNO_{3\ (aq)}$
  - ionically: $Pb^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$ ==> $PbSO_{4(s)}$
  - because the spectator ions are sodium Na^+ and nitrate NO_3^-

THE PERIODIC TABLE OF ELEMENTS

Mendeleev Dmitri (1869), the Russian chemist arranged elements in order of masses and found that elements with similar physical and chemical properties appeared at periodic intervals. Some elements appeared periodically after eight (8) elements. Sometimes after eighteen (18) elements and sometimes after 32 elements.

The modern periodic table is arranged according to the **increasing atomic number** of elements. The periodic table is the method of classifying elements in groups and periods according to using the atomic number of elements.

Periodicity is the occurrence of successive group of elements sharing strong chemicals properties.

GROUP (FAMILIES)

Groups, also referred to as families are vertical columns of elements on the periodic table. The elements within a group of the periodic table show marked physical and chemical similarities with a graduation of properties. There are eight (8) groups of elements on the periodic table. Groups are arranged according to the number of electrons in the valence shell (outermost shell). Elements with the same number of electrons in the outermost shell are in the same group. The number of electrons in the outermost shell of an element is equal to the group it belongs.

PERIODS

Periods are horizontal rows of elements on the periodic table. Elements within a period have different physical and chemical properties. There are seven (7) periods of elements on the periodic table. Periods are arranged according to the number of shells. Elements with the same number of shells are in the same period. The number of shells an element has is equal to the period it belongs. As you move from left to right across the period, you move from metals to non metals. As you move from right to left across the periodic table, you move from non-metals to metals.

Group One (1) Elements (Alkali Metals)

The alkali metals are the series of elements in group 1 of the periodic table of elements (excluding hydrogen in all but one are circumstances.) The series consists of the elements Lithium (Li), sodium (Na), potassium (K), rubidium (Rb), caesium (Cs) and francium (Fr).

Physical Properties of Group 1 Elements

- The alkali metals are silver-coloured when fleshly cut
- Soft and easy to cut with a knife
- Low-density metals (float on water)
- Low melting and boiling points
- Are good conductors of heat and electricity.

Chemical Properties of Group 1 Elements

These elements all have one valence electron which is easily lost to form an ion with a single positive charge. They have the lowest ionisation energy in their respective periods. This makes them very reactive and they are the most active metals. Due to their activity they occur naturally in ionic naturally in ionic compounds not in their elemental state.

1. They burn in air with coloured flames to form metal oxides,

E.g $4Na + O_2 \longrightarrow 2Na_2O$

$4K + O_2 \longrightarrow K_2O$

These elements all have one valence electron which is easily los to form an ion with a single positive charge. They have the lowest ionisation energies in their respective periods. This makes them very reactive and they are the most active metals. Due to their activity they occur naturally in ionic compounds not in their elemental state.

2. The alkali metals react readily with halogens to form ionic salts, such as table salt, sodium chloride (NaCI).

E.g $Na + Cl \longrightarrow NaCl$

$Fr + F \longrightarrow FrF$

3.They are famous for their vigorous reactions with water to liberate hydrogen gas. These reaction also often liberates sufficient energy to ignite the hydrogen and can be quite dangerous. As we move down the group the reactions become increasingly violent. The reaction with water is as follows:

Alkali metal + water $\longrightarrow$ Alkali metal hydroxide + hydrogen gas

With potassium as an example:

$2K_{(s)} + 2H_2O_{(l)} \longrightarrow 2KOH_{(aq)} + H_{2(g)}$

The oxide, hydrides, and hydroxide of these metals are basic (alkaline) . In particular the hydroxide resulting from the reaction from water are our most common laboratory bases (alkalis) . It is from this character that they derive their group name.

Hydrogen also has a single valence electron and is usually placed at the top of group 1, but it is not a metal (except under extreme circumstances as metallic hydroxide); rather it exists naturally as a diatomic gas. Hydrogen can form ions with a single positive charge, but removal of its single electron requires considerably more energy than removal of the outer electron from the alkali metals. Unlike the alkali metals hydrogen atoms can also gain an electron to form the negatively charged hydride ion. The hydride ion is an extremely strong base and does not usually occur except when combined with the alkali metals and some transition metals (i.e. the ionic sodium hydride, NaH). In compounds hydrogen most often forms covalent bonds.

4. Alkali metals react with acids to form salt and hydrogen gas.

E.g $2K + 2HCl \longrightarrow 2KCl + H_2$

$2Li + H_2SO4 \longrightarrow Li_2SO4 + H_2$

Group 2 Elements (Alkaline Earth Metals)

The alkaline earth metals are the series of elements in group 2 of the periodic table. The series consists of the elements beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra), (though radium is not always considered an alkaline on earth due to its radioactivity). Reactivity of this group increases as you move down the group.

Physical properties

- The alkaline earth metals are silvery coloured when fleshly cut,
- Soft and easily cut with a knife, though are a bit harder than the alkali metals
- low-density metals,
- Are good conductors of heat and electricity.

Chemical properties

These elements all have two valence electrons and tend to lose both to form ions with a two plus charge. Berylium is the least metallic element in the group and tends form covalent bond in its compounds. They are less activity than the alkali metals, but are still fairly active.

- They react readily with halogens (group 7 elements) to form ions salts,
- They can react slowly with water to form basic hydroxide. Magnesium reacts only with steam and calcium with hot water. Beryllium is an exception: It does not react with water or steam, and its halides are covalent. The oxide are basic and dissolve in acids and the hydroxides are strong bases, though not as soluble as the alkali metal hydroxides.
- They react with acids to form salt, e.g $Mg_{(s)} + 2HI_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$

Group 7 Elements (Halogens)

The halogens are the elements in group seven (7). They are fluorine (F), Chlorine (Cl), bromine (Br), iodine (I), Astatine (At).

Physical properties

- They are non –metals which form diatomic (two atoms) molecules with coloured vapours;
- they are poisonous;

- they are poor conductors of heat and electricity;
- all have one less electron in the outermost shell.

Other physical properties of the halogens.

Halogen	Melting point (^0C)	Boiling point (^oC)	State	Colour
Fluoride	-220	-188	Gas	Pale yellow
Chloride	-101	-35	Gas	Dense Green
Bromine	-7	59	Liquid	Reddish brown
Iodine	114	184	Solid	Black
Astatine	-	-	Solid	-

Note that astatine has many colours depending on the physical state in which it is.

Chemical Properties

Halogens are highly reactive, and as such can be harmful or lethal to biological organisms in sufficient quantities. Fluorine is the most reactive and the reactivity declines as we go down the group but increases as you go up the group.

Group seven elements require one more electrons to fill their outer electron shell, and so have a tendency to gain one electron to form a single-charged negative ion. These negative ions are referred to as **halide ions**, and salts containing these ions are known as **halides.**

- React with metals to form ionic compounds (neutral salts);
- React with other non-metals to form molecular compound;
- The oxide and hydrides, like those of most non-metal, of the halogens are acidic. Halide ions combined with single hydrogen atoms form the hydrohalic acids (i.e, HF, HCI, HBr, HI), a series of particularly strong acids.
- More react halogens will displace less reactive halogens from aqueous solution.
  - $Cl_{2(g)} + 2Nal_{(aq)} \longrightarrow 2NaCl_{(aq)} + I_{2(l)}$
  - $F_{2(g)} + 2NaCl_{(aq)} \longrightarrow 2NaF_{(aq)} + Cl_{2(l)}$
  - $Br_{2(g)} + 2kl_{(ag)} \longrightarrow 2KBr_{(aq)} + I_{2(S)}$

Uses of Halogens

- Small of **fluorine** is added to tap water and toothpaste to prevent tooth decay;
- Chloride is used to treat tap water and swimming pools to kill harmful germs and bacteria;
- **Iodine** is used as an antiseptic; small amounts of iodine are needed in our bodies to prevent goitre (swelling of thyroid gland);
- **Silver halides** are used in black and white photographic film.

Group 8 Elements (Noble Gases/Rare Gases/inert gases)

The **noble gases** are the elements in group 8 (group VIII or 0) of the periodic table. They are helium, neon, argon, krypton, xenon, and radon. They are sometimes called **inert gases** or **rare gases**. They are called **'noble gases'** due to their preciousness, resistance to corrosion. The word '**inert**' means **unreative**. Therefore, group 8 elements are un- reactive.

Physical Properties

- They are colourless gases;
- They exist as monatomic gases at room temperature,
- They have very low melting points and boiling point due to their very inter-atomic forces of attraction;

Chemical properties

The noble gases are all non-metals and are characterized by having completely filled shells of electrons. In general this makes them very unreactive chemically since it is difficult to add or remove electrons.

Uses of Noble Gases

- **Helium** is used in balloons and airships because it is less dense than air (second lightest gas) and is not flammable like hydrogen**.**
- **Neon** is used in advertising signs because it glows red when electricity is discharged through it.
- **Argon** is used to fill filament lamps (light bulbs). It prevents the filament inside the bulb from burning out.
- **Krypton** and **Xenon** are used in lamps in light houses, stroboscopic lamps and photographic flash units.

Transition Metals

Transition metals are found in the centre block of the periodic table.

Properties

- They form coloured ions in aqueous solution: e.g Copper (II) is blue; iron (II) is pale green and iron (III) is reddish brown (when solid) and yellow (when in solution;
- They are typically metals;
- They form positively charged ions with variable charges. E.g copper forms either Cu^{+} or Cu^{2+} and ion forms either Fe^{2+} or Fe^{3+}
- They are hard and strong metals;
- They have high melting and boiling points;
- They have high density

Uses of Transition Metals

- As catalysts in industries to speed up reactions. E .g the hydrogenation of oil to make margarine uses a nickel catalyst, the manufacture of ammonia uses iron catalyst;
- Many of them are used to make alloys. E. g steel is made by mixing iron with small amount of carbon;
- Due to their high strength and hardness, many of them are useful engineering materials.

Semi-metals or Metalloids

These are elements near the zig-zag line. Examples includes aluminium, Silicon, boron, gallium, etc. Semi –metals have the characteristics of both metals and non-metals**.**

Unique Position of Hydrogen on the Periodic Table

Hydrogen resembles both group 1 elements and group 7 elements:

- It resembles group 1 elements in that just like group 1 elements, it has 1 electron in the outermost shell; and under special condition it can lose one electron in the outermost.
- It resembles group 7 elements in that it also requires 1 electron for it to be electrically stable.

HEAT (ENERGY) CHANGES

HEAT CHANGES IN CHEMICAL REACTIONS

- **When chemical reactions occur**, as well as the formation of the products - the **chemical change**, there is **also a heat energy change** which can often be detected as a temperature changes.

This means the products have different energy content than the original reactants.

There are two types of energy (heat) changes namely **Exothermic** and **Endothermic reactions.**

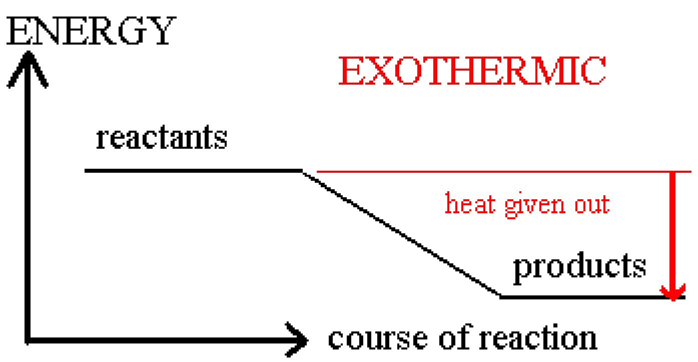
Exothermic Reaction

This is the type of energy change where heat energy is **released** (given out) to the surrounding. This results in the rise in temperature in the surroundings. In exothermic reaction, the products contain less energy than the reactants

Examples:

- the burning or combustion of hydrocarbon fuels (see **Oil Products** e.g. petrol or candle wax.
- the burning of magnesium, reaction of magnesium with acids, or the reaction of sodium with water.
- the neutralisation of acids and alkalis (see **Acids, Bases and salts**)
- using hydrogen as a fuel in fuel cells.
- condensation
- freezing
- bond formation

The energy profile diagram for exothermic reaction is shown below.



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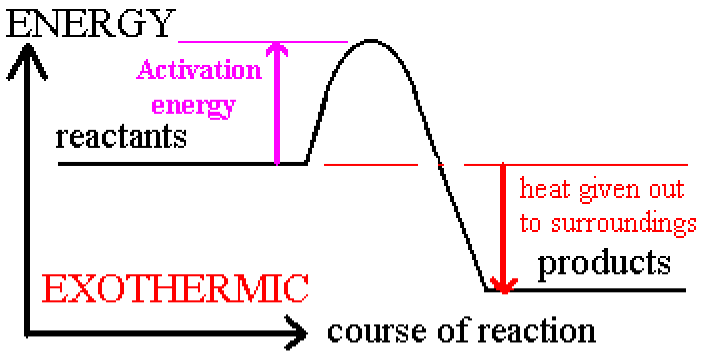


Energy change also referred to as enthalpy change can be calculated by using the formula:

Energy (enthalpy) change = Energy of products – Energy of reactants, which is:

$\Delta H = Hp - Hr$

The heat energy change (ΔH, enthalpy change) involved is expressed in kilojoules per mole (kJ/mol).

Since the products have less energy than the reactants in exothermic reaction, the enthalpy change is **ALWAYS** a **NEGATVE value**, that is **$-\Delta H$**.

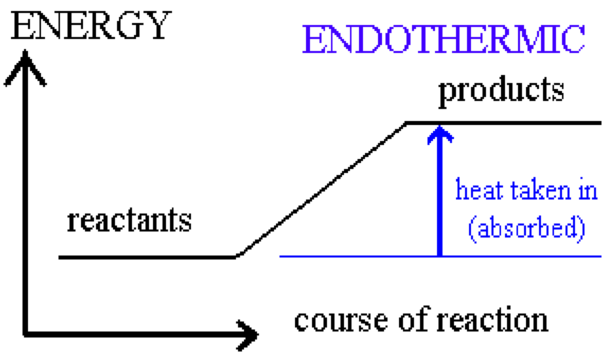
Endothermic Reaction

This type of energy change involves the absorption of heat energy from the surrounding. This results in the decrease in temperature in the surroundings. In endothermic reaction, the products contain more energy than the reactants.

Examples:

- the thermal decomposition of limestone
- the cracking of oil fractions
- boiling
- evaporation
- melting
- photosynthesis
- photography
- bond breaking

The energy profile diagram for endothermic reaction is shown below.



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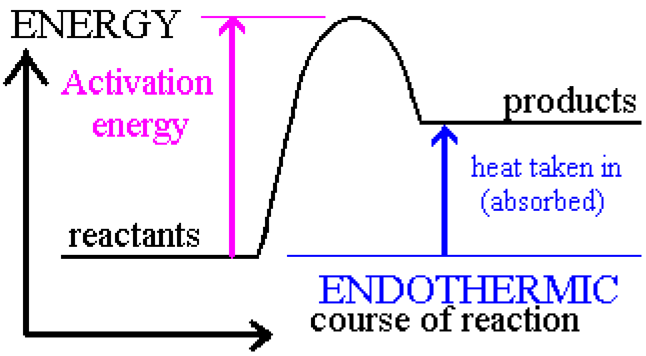


Since the products have more energy than the reactants in endothermic reaction. the enthalpy change is **ALWAYS** a **POSITVE** value, that is, +Δ**H**.

Calculations Involving Enthalpy Changes

The table below shows bond dissociation energies

Bond type	Bond energy (KJ/mol
H-O	413
N-N	346
N-H	389
H-H	436
Cl-Cl	242
H-Cl	431
Br-Br	193
O-O	339
H-Br	365
S-O	

Examples

1. Calculate the enthalpy of formation of hydrogen chloride (HCl).

Answer: Need to write a balanced chemical equation first: $H_2 + Cl_2 \longrightarrow 2HCl$

Δ Δ Δ

H= Hp –Hr H = 2(HCl) – (H_2 + Cl_2) H= 2(431) – (436+242)

ΔH = 862 – 678 ΔH = **184KJ/mol**

Since ΔH is a positive value, the heat energy is endothermic.

2. Find the enthalpy of formation of water (H_2O)

Answer: $2H_2 + O_2 \longrightarrow 2H_2O$

ΔH = Hp –Hr Δ H = 2(O-H-O) - 2(H-H) +O-O

ΔH = 2(2X413) – 2(2 X 436)+ 2 X 339

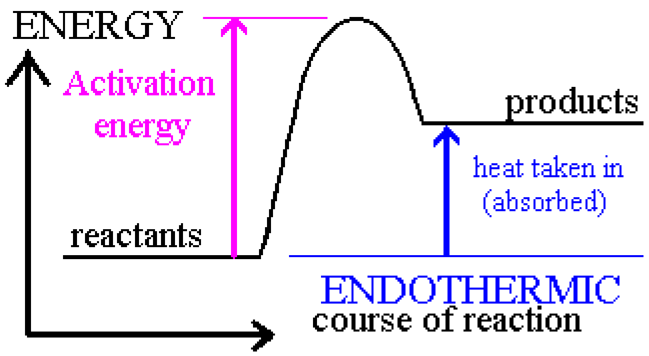
ΔH = 1652 – 1744 + 678 ΔH = 1652-2422 ΔH = **-770KJ/mol**

Since the value of ΔH is negative, the change in heat energy is exothermic.

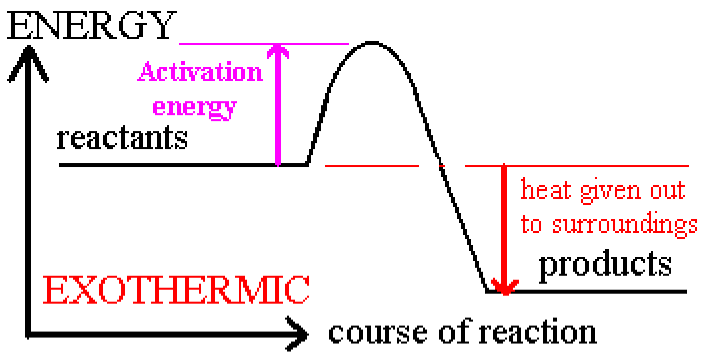
Activation Energy

This is the minimum amount of energy required to start the reaction.

- The activation energy 'hump' can be related to the process of **bond breaking and making**
- Up the hump is endothermic, representing breaking bonds (energy absorbed, needed to pull atoms apart),



- down the other side of the hump is exothermic, representing bond formation (energy released, as atoms become electronically more stable).



RATE OF REACTIONS

The phrase '**rate of reaction**' means '***how fast is the reaction***' or '***the speed of the reaction***'. It can be measured as the '**rate of formation of products**' (e.g. collecting gaseous product in a syringe) or the '**rate of reaction of reactants**'. The speeds of reactions are very varied.

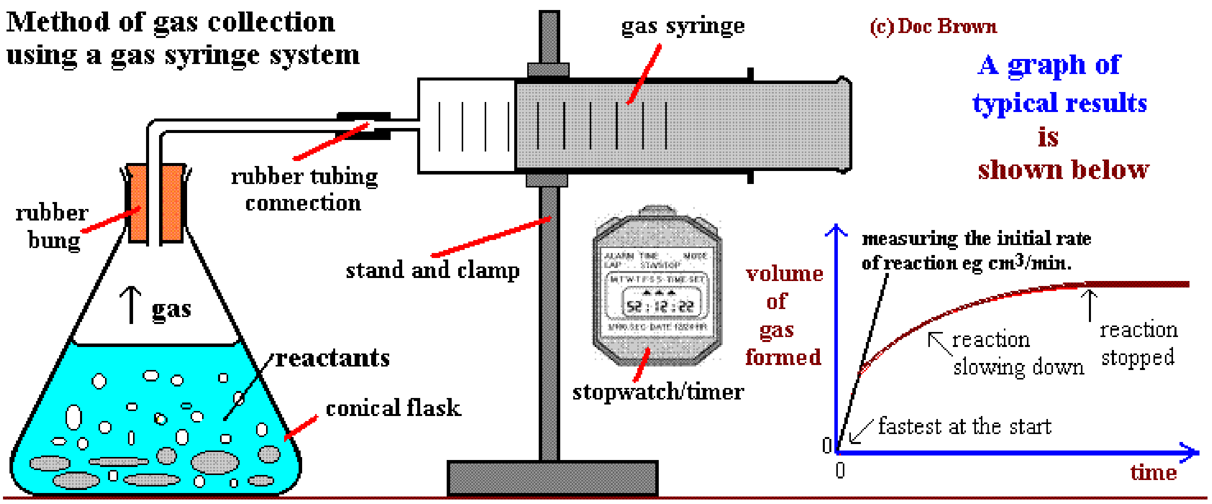
Importance of "Rates of Reaction knowledge":

Time is money in industry, the faster the reaction can be done, the more economic it is. You need to know how long reactions are likely to take. Hence the great **importance of catalysts** e.g. transition metals or enzymes which reduce time and save money. A reaction will continue until one of the reactants is used up.

Measuring the Speed of a Reaction

To measure the 'speed' or 'rate' of a reaction depends on what the reaction is, and whether what is formed can be measured as the reaction proceeds. Two examples are outlined below.

When a gas is formed from a solid reacting with a solution, it can be collected in a **gas syringe** (see **diagram** below and the **graph**).



In general, the rate or speed of a reaction can be measured by using the formula:

$$Rate\ of\ Reaction = \frac{amount\ of\ products\ produced\ or\ amount\ of\ reactants\ used\ up}{time\ taken}$$

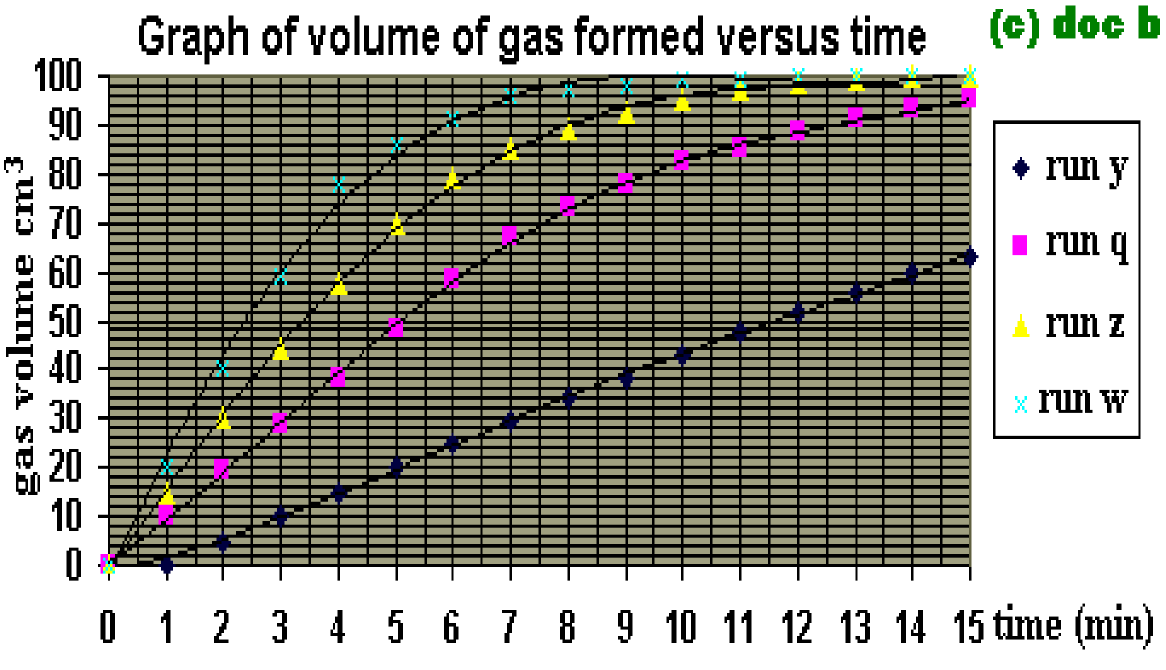
The units for products and reactants can be g, kg, cm^3, etc while the units for time taken can be minutes (min), hours (hr) or seconds (s). Therefore, the units for rate of reaction can be g/min, g/s kg/hr, cm^3/min depending on the units used for products produced or reactants used up and time taken.

The initial gradient of the graph e.g. in **cm^3/min** (speed or rate) gives an accurate measure of how fast a gaseous product is being formed in metal/carbonate - acid reaction (forming H_2/CO_2 respectively). You can measure the gas formed every e.g. 30 seconds and plot the graph and measure the initial gradient in e.g. cm^3/min or cm^3/sec.

The most accurate measurements are made early on in the reaction when the gas volume versus time is almost linear. You can take a series of measurements and draw the graph (origin **0,0**) to get the rate from the gradient (e.g. cm^3/min) or measure the time to make a fixed volume of gas (see above).

If the reaction is allowed to go on, you can measure the final maximum volume of gas and the time at which the reaction stops, though this a very poor measure of rate, because the reaction just goes slower and slower as the reactant amounts/concentrations are decreasing - so **don't** use this as a method of measuring reaction speed.

The reciprocal of the reaction time, **1/time**, can also be used as a measure of the speed of a reaction. The time can represent how long it takes to form a fixed amount of gas first few minutes of a metal/carbonate - acid reaction, or the time it takes for so much sulphur to form to obscure the X in the sodium thiosulphate - hydrochloric acid reaction. The time can be in minutes or seconds, as long as you stick to the same unit for a set of results e.g. a set of experiments varying the concentration of one of the reactants.The shape of the graph is quite characteristic (see diagram below):



- The **reaction is fastest at the start** when the reactants are at a maximum (steepest gradient in cm^3/min).
- The **gradient becomes progressively less** as reactants are used up and the reaction slows down. 8
- Finally the **graph levels out when one of the reactants is used up** and the reaction stops.

The **steeper** the graph the **faster** (the higher) the rate of a chemical reaction and vice versa. From the grapy above:

✓ Run or reaction y is faster than run q; run q is faster than run z; run z is faster than run w and vice versa.

From the graph of results you can measure the relative rate of reaction from (i) the initial gradient in cm^3/min (see on diagram above), (ii) you can estimate from the graph the volume of gas formed after a particular time e.g. 3 minutes or (iii) you can estimate the time it takes to form a particular volume of gas. (i) is the best method i.e. the best straight line covering several results at the start of the reaction.

Examples

Use the graph shown above to answer the following questions.

1. How much gas was produced in 11 minutes by run **y**?

Answer: Simply check on the graph for run **y**, on time 11 min, move up, where you meet the line for run **y**, move to the left hand side to check on the volume produced in 11 min. You will find that in 11 min the volume produce is **49cm³** or **50cm³**.

2. How long did it take to form $30cm^3$ of gas for run **z**?

Answer: on volume $30cm^3$ move to the right hand side. When you meet line **z,** move down to check on time. The time it took to for run **z** form $30cm^3$ is **2 min**.

3. Calculate the rate of reaction q in 8 min

Answer: Rate of reaction: $\frac{amount\ of\ gas\ produced(cm3}{time\ taken(\text{min}\quad)}$, Rate of reaction = $\frac{70cm3}{8min}$,

Rate of reaction = $\mathbf{8.75cm^3/min}$

Exercise

Use the graph shown above to answer the following questions.

1. How long did it take for run **w** to produce $10cm^3$ of the gas?

2. Find the amount of gas formed in 6 min for run **z.**

3. Work out the rate of reaction in 11.5 min for reaction **y.**

4. How long did it take for the reactant(s) to be used up for runs (reactions) **w** and **z** respectively?

5. Briefly explain why graphs for both run **w** and **z** started flattening from 10 min and 12 min respectively.

Examples of reactions involving gas formation

- (i) metals dissolving in acid ==> hydrogen gas, (test is lit splint => pop!),
  - e.g. magnesium + sulphuric acid ==> magnesium sulphate + hydrogen
    - $\mathbf{Mg_{(s)} + H_2SO_{4(aq)}}$ ==> $\mathbf{MgSO_{4(aq)} + H_{2(g)}}$
- (ii) carbonates dissolving in acids => carbon dioxide gas, (test is limewater => cloudy),
  - calcium carbonate *(marble chips)* + hydrochloric acid ==> calcium chloride + water + carbon dioxide
    - $\mathbf{CaCO_{3(s)} + 2HCl_{(aq)}}$ ==> $\mathbf{CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}}$
- and (iii) the manganese(IV) oxide catalysed decomposition of hydrogen peroxide (oxygen gas, test is glowing splint => relights)
  - hydrogen peroxide ==> water + oxygen
    - $\mathbf{2H_2O_{2(aq)}}$ ==> $\mathbf{2H_2O_{(l)} + O_{2(g)}}$
      - can all be followed with the gas syringe method.

You can do all sorts of investigations to look at the effects of

(a) the solution **concentration**,

(b) the **temperature** of the reactants,

(c) the size of the solid particles (**surface area** effect),

(d) the **effectiveness of a catalyst** on hydrogen peroxide decomposition.

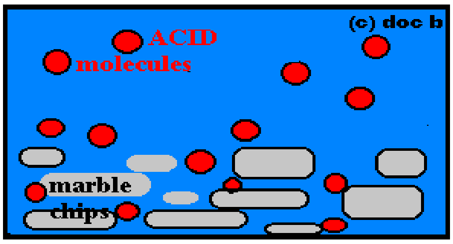
Factors that Affect Rate of Chemical Reactions

NOTE THAT anything that causes high collision to the reacting particles increases the rate (speed) of a chemical reaction and vice versa.

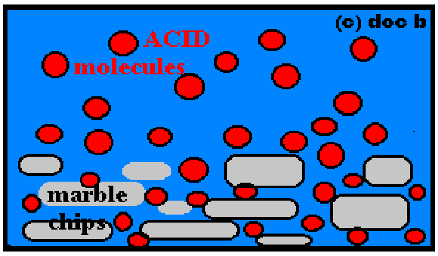
The effect of Concentration ((see also graphs 4.6, 4.7 and 4.8 below**)**

- what is the effect of changing the concentration of a reactant?
- and why is the reaction speed changed?
- Why does increase in concentration speed up a reaction?
- If the concentration of any reactant in a solution is increased, the rate of reaction is increased vice versa.

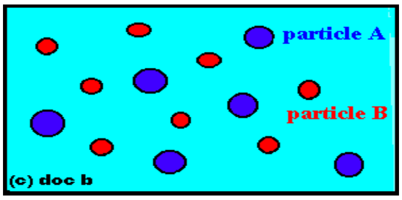
- Increasing the concentration, increases the probability of a collision between reactant particles because there are more of them in the same volume and so increases the chance of a fruitful collision forming products.
- e.g. Increasing the concentration of acid molecules increases the frequency or chance at which they hit the surface of marble chips to dissolve them (slower => faster, illustrated below)



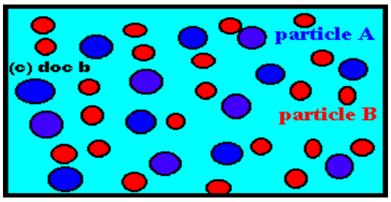
slower => faster



- In general, increasing the concentration of reactant A or B will increase the chance or frequency of a successful collision between them and increase the speed of product formation (slower => faster, illustrated below).



slower => faster



- Increasing the concentration of reactant A or B will increase the chance or frequency of collision between them and increase the speed of product formation (slower => faster).

The effect of Pressure

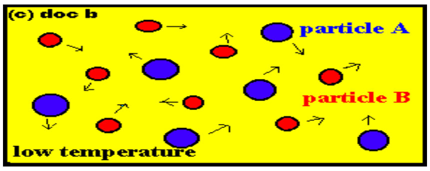
- what is the effect of changing pressure on the speed of a reaction?
- does increasing the pressure always have an effect?
- Why does an increase in pressure speed up a reaction with a gaseous reactant**?**
- If one or more of the reactants is a gas then **increasing pressure** will effectively **increase** the concentration of the reactant molecules and **speed up the reaction** and vice versa.
- The particles are, on average, closer together and collisions between the particles will occur more frequently.
- The A and B particle diagrams above could represent lower/higher pressure, resulting in lesser or greater concentration and so slower or faster reaction all because of the **increased chance of a 'fruitful' collision**.

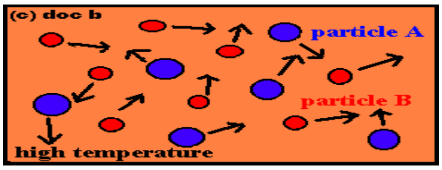
Solid reactants and solutions are **NOT** affected by change in pressure, there concentration is **unchanged**

The effect of Temperature (see also graphs 4.3, 4.4 and 4.8 below)

- does temperature affect the speed of a chemical reaction?
- if so, how and why?

- Why does a reaction go faster at a higher temperature and vice versa?
- When gases or liquids are heated the particles gain kinetic energy and move faster (see diagrams below).
- The **increased speed increases the chance (frequency) of collision** between reactant molecules and the rate increases. The other way round is true.



slower => faster


The effect of a Catalyst (see also light effect and graph 4.8)

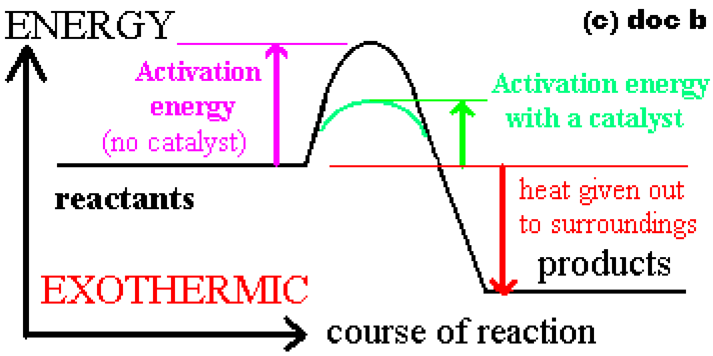
- what is a catalyst?
- how does it affect the speed of a chemical reaction?
- how does a catalyst work?
- Why does a catalyst speed up a reaction?

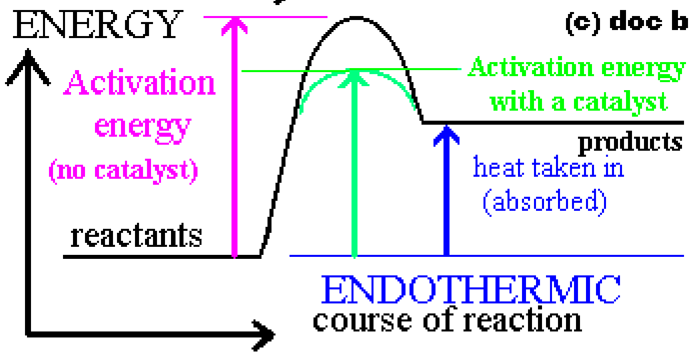
I was once asked by one pupil from Kabulonga Girls' High School ''what is the opposite of a catalyst''? There is no real opposite to a catalyst, other than the uncatalysed reaction!

The word **catalyst means an added substance, in contact with the reactants, that changes the rate of a reaction without itself being chemically changed in the end**.

- There are the two phrases you may come across:
  - a 'positive catalyst' meaning speeding up the reaction (plenty of examples in most chemistry courses)
  - OR a 'negative catalyst' slowing down a reaction (rarely mentioned at GCSE, sometimes at AS-A2 level, e.g. adding a chemical that 'mops up' free radicals or other reactive species).

Catalysts **increase the rate** of a reaction by helping break chemical bonds in reactant molecules and provide a 'different pathway' for the reaction. This **effectively means the Activation Energy is reduced**, irrespective of whether it's an exothermic or endothermic reaction (see diagrams below).





Therefore at the same temperature, more reactant molecules have enough kinetic energy to react compared to the uncatalysed situation. Different reactions need different catalysts and they are extremely important in industry: examples:

- **nickel** catalyses the hydrogenation of unsaturated fats to margarine
- **iron** catalyses the combination of unreactive nitrogen and hydrogen to form ammonia
- **enzymes** in yeast convert sugar into alcohol
- **zeolites** catalyse the cracking of big hydrocarbon molecules into smaller ones
- most **polymer making reactions require a catalyst** surface or additive in contact with or mixed with the monomer molecules.
- **Enzymes are biochemical catalysts**- enzymes and biotechnology
- They have the **advantage of bringing about reactions at normal temperatures and pressures which would otherwise need more expensive and energy-demanding equipment**.

The Effect of Light

- can light affect the speed of any reactions?
- if it does, how does change the speed of a chemical reaction?
- Why does increasing light intensity sometimes increase the speed of a reaction?
- **Light energy** (ultra violet or visible radiation) can initiate or catalyse particular chemical reactions.

As well as acting as an electromagnetic wave, light can be considered as an energy 'bullets' called photons and they have sufficient 'impact energy' to break chemical bonds, that is, enough energy to overcome the activation energy.

The greater the intensity of light (visible or ultra-violet) the more reactant molecules are likely to gain the energy react, so the reaction speed increases and vice versa.

- Examples:
  - **Silver salts** are converted to silver in the chemistry of photographic exposure of the film.
    - Silver chloride (AgCl), silver bromide (AgBr) and silver iodide (AgI) are all sensitive to light ('photosensitive'), and all three are used in the production of various types of photographic film to detect visible light and beta and gamma radiation from radioactive materials.
    - Each silver halide salt has a different sensitivity to light.
    - When radiation hits the film the silver ions in the salt are reduced by electron gain to silver
      - $\mathbf{Ag^+ + e^-}$ ==> **Ag** (X = halogen atom, Cl, Br or I)
      - and the halide ion is oxidised to the halogen molecule by electron loss
      - $\mathbf{2X^-}$ ==> $\mathbf{X_2 + 2e^-}$
      - so overall the change via light energy is: **2AgX** ==> $\mathbf{2Ag + X_2}$
    - AgI is the least sensitive and used in X-ray radiography, AgCl is the most sensitive and used in 'fast' film for cameras.
  - **Photosynthesis** in green plants:
    - The conversion of water + carbon dioxide ==> glucose + oxygen
      - $\mathbf{6H_2O_{(l)} + 6CO_{2(g)}}$ ==> $\mathbf{C_6H_{12}O_{6(aq)} + 6O_{2(g)}}$
      - requires the input of sunlight energy and the green chlorophyll molecules absorb the photon energy packets of light and initiate the chemical changes summarised above.
  - **Photochemical Smog**:
    - This is very complex chemistry involving hydrocarbons, carbon monoxide, ozone, nitrogen oxides etc.
    - Many of the reactions to produce harmful chemicals are catalysed or promoted by light energy.

More examples of interpreting graphical results ('graphing'!)

Plotting graphs - plots of graphs of data and how to interpret them

Note that:

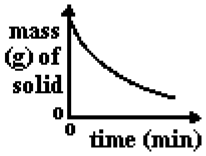
(i) rate of reaction = speed,

(ii) Graphs 4.1, 4.2 and 4.5 just show the theoretical shape of a graph for a single particular experiment. Graphs 4.3 and 4.4 (temperature), 4.6 and 4.7 (concentration) and 4.8 (several factors illustrated) shows the effect of changing a variable on the rate of the reaction and hence the relative change in the curve-shape of the graph line.

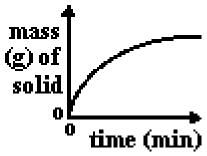
(iii) The rate of reaction may be expressed as the reciprocal of the reaction time (1/time) e.g. for the time for sulphur formation (to obscure the X) in the sodium thiosulphate -

hydrochloric acid reaction or where a fixed volume of gas is formed, though in this can also be expressed as gas volume/time too as cm^3/s or cm^3/min even though the gas volume is the same for a given set of results of changing one variable whether it be concentration or temperature.

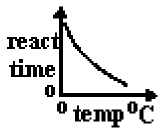
If you have detailed data e.g. multiple gas volume readings versus time, the best method for rate analysis is the initial rate method described on and below the diagram of the gas syringe gas collection system.



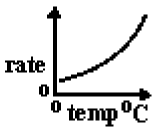
Graph 4.1 shows the decrease in the amount of a solid reactant with time. The graph is curved, becoming less steep as the gradient decreases because the reactants are being used up, so the speed decreases. Here the gradient is a measure of the rate of the reaction. In the first few minutes the graph will (i) decline less steeply for larger 'lumps' and (ii) decline more steeply with a fine powder i.e. (i) less surface area gives slower reaction and (ii) more surface area a faster reaction.



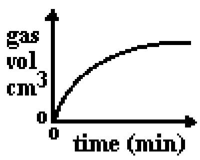
Graph 4.2 shows the increase in the amount of a solid product with time. The graph tends towards a maximum amount possible when all the solid reactant is used up and the graph becomes horizontal. This means the speed has become zero as the reaction has stopped. Here the gradient is a measure of the rate of the reaction.



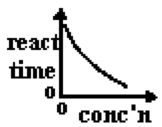
Graph 4.3 shows the decrease in reaction time with increase in temperature as the reaction speeds up. The reaction time can represent how long it takes to form a fixed amount of gas in e.g. in the first few minutes of a metal/carbonate - acid reaction, or the time it takes for so much sulphur to form in the sodium thiosulphate - hydrochloric acid reaction. The time can be in minutes or seconds, as long as you stick to the same unit for a set of results e.g. a set of experiments varying the concentration of one of the reactants. Theory of temperature effect.



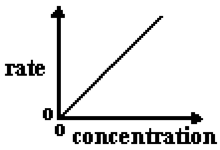
Graph 4.4 shows the increase in speed of a reaction with increase in temperature as the particles have more and more kinetic energy. The rate of reaction is proportional to 1/t where t = the reaction time. See the notes on rate in the Graph 4.7 paragraph below and the theory of temperature effect.



Graph 4.5 shows the increase in the amount of a gas formed in a reaction with time. Here the gradient is a measure of the rate of the reaction. Again, the graph becomes horizontal as the reaction stops when one of the reactants is all used up.

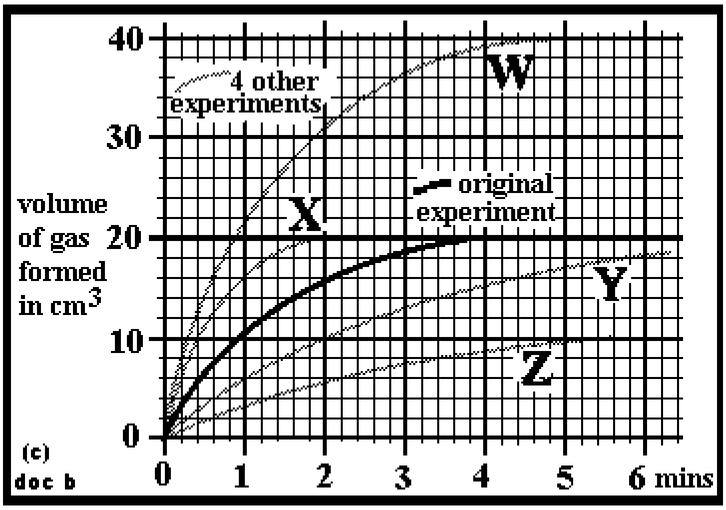


Graph 4.6 shows the effect of increasing concentration, which decreases the reaction time, as the speed increases because the greater the concentration the greater the chance of fruitful collision. See the notes on rate in the Graph 4.3 paragraph above and the theory of concentration effect



Graph 4.7 shows the rate/speed of reaction is often proportional to the concentration of one particular reactant. This is due to the chance of a fruitful collision forming products being proportional to the concentration. The initial gradient of the product-time graph e.g. for gas in cm^3/min (or s for timing the speed/rate) gives an accurate measure of how fast the gaseous product is being formed. The reciprocal of the reaction time, 1/time, can also be used as a measure of the speed of a reaction. The time can e.g. represent

how long it takes to make a fixed amount of gas, or the time it takes for so much sulphur to form in the sodium thiosulphate - hydrochloric acid reaction. The time can be in minutes or seconds, as long as you stick to the same unit for a set of results for a set of experiments varying the concentration or mass of one of the reactants. Theory of concentration effect



Graph 4.8 A set of results for the same reaction

(i) The graph lines W, X, original, Y and Z on the left diagram are typical of when **a gaseous product is being collected**. The middle graph might represent the original experiment 'recipe' and temperature. Then the experiment repeated with variations e.g.

(ii) X could be the same recipe as the original experiment but a catalyst added, forming the same amount of product, but faster.

(iii) Initially, the increasing order of rate of reaction represented on the graph by curves Z to W i.e. W > X > original > Y > Z might represent progressively increasing concentrations of reactant or progressively higher temperature of reaction or progressively smaller lumps-particle/increasing surface area of a solid reactant. All three trends in changing a reactant/reaction condition variable produce a progressively faster reaction shown by the increasing gradient in cm^3/min which represents the rate/speed of the reaction.

(iv) Z could represent taking half the amount of reactants or half a concentration. The reaction is slower and only half as much gas is formed.

(v) W might represent taking double the quantity of reactants, forming twice as much gas e.g. same volume of reactant solution but doubling the concentration, so producing twice as much gas, initially at double the speed (gradient twice as steep).

REDOX Reactions

These are coupled reactions in which one is **Reduction** and the other is **Oxidation**, hence REDOX.

OXIDATION & REDUCTION – REDO X REACTIONS	
OXIDATION - definition and examples	**REDUCTION - definition and examples**
(a) The gain or addition of oxygen to an atom, molecule or ion e.g. ... **(1) S + O_2** ==> **SO_2** [burning sulphur - oxidised]	**(b) The loss or removal of oxygen from a compound etc.** e.g. ... **(1) CuO** ==> **Cu + O_2** [loss of oxygen from copper(II) oxide to form copper]

(2) CH_4 + O_2 ==> **CO_2 + H_2O** [burning methane to water and carbon dioxide, C and H gain O] **(3) NO + O_2**==> **NO_2** [nitrogen monoxide oxidised to nitrogen dioxide] **(4) SO_2 + O_2** ==> **SO_3** [oxidising the sulphur dioxide to sulphur trioxide in the Contact Process for making sulphuric acid]	**(2) Fe_2O_3** ==> **Fe + O_2** [iron(III) oxide reduced to iron in blast furnace] **(3) NO** ==> **N_2 + O_2** [nitrogen monoxide reduced to nitrogen, catalytic converter in car exhaust] **(4) SO_3** ==> **SO_2 + O_2** [sulphur trioxide reduced to sulphur dioxide]
(c) The loss or removal of hydrogen from an atom or substance (i) H_2S +Cl_2 ==> 2HCl +S **H_2S** has been **oxidized** to **S** by the removal of H_2. (ii) $2H_2O$==> $2H_2$ + O_2 H_2O has been oxidized to O_2 by the Removal of O_2. **(d) The loss or removal of electrons from an atom, ion or molecule** e.g. **(1) Fe** ==> **Fe^{2+} + $2e^-$** [iron atom loses 2 electrons to form the iron(II) ion, start of rusting chemistry] **(2) Fe^{2+}** ==> **Fe^{3+} + e^-** [the iron(II) ion loses 1 electron to form the iron(III) ion] **(3) $2Cl^-$** ==> **Cl_2 + $2e^-$** [the loss of electrons by chloride ions to form chlorine molecules]	**(c) The gain or addition of hydrogen to an atom or substance** (i) H_2S +Cl_2 ==> 2HCl +S Cl_2 has been reduced to HCl by the addition or gain of H. (ii) $2H_2$ + O_2==> $2H_2O$ O_2 has been reduced to H_2O by the addition of $H_{2.x}$ **(d) The gain o r addition of electrons by an atom, ion or molecule** e.g. ... **(1) Cu^{2+} + $2e^-$** ==> **Cu** [the copper(II) ion gains 2 electrons to form neutral copper atoms, electroplating or displacement reaction) **(2) Fe^{3+} + e^-** ==> **Fe^{2+}** [the iron(III) ion gains an electron and is reduced to the iron(II) ion] **(3) $2H^+$ + $2e^-$** ==> **H_2** [hydrogen ions gain electrons to form neutral hydrogen molecules, electrolysis of acids or metal-acid reaction]
(e) An oxidising agent is the species that gives the oxygen or removes the electrons	**(f) A reducing agent is the species that removes the oxygen or acts as the electron donor**

REDOX REACTIONS - in a reaction overall, oxidation and reduction must go together

(g) Redox reaction analysis based on the oxygen definitions

- **(1) copper(II) oxide + hydrogen** ==> **copper + water**
  - **CuO(s) + H_2(g)** ==> **Cu(s) + H_2O(g)**
  - copper oxide reduced to copper, hydrogen is oxidised to water
  - hydrogen is the reducing agent (removes O from CuO)

- copper oxide is the oxidising agent (donates O to hydrogen)

- **(2) iron(III) oxide + carbon monoxide** ==> **iron + carbon dioxide**
  - **Fe_2O_3(s) + 3CO(g)** ==> **2Fe(l) + 3CO_2(g)**
  - the iron(III) oxide is reduced to iron, the carbon monoxide is oxidised to carbon dioxide
  - CO is the reducing agent (O remover from Fe_2O_3)
  - the Fe_2O_3 is the oxidising agent (O donator to CO)]
- **(3) nitrogen monoxide + carbon monoxide** ==> **nitrogen + carbon dioxide**
  - **2NO(g) + 2CO(g) ==> N_2(g) + 2CO_2(g)**
  - nitrogen monoxide is reduced to nitrogen
  - carbon monoxide is oxidised to carbon dioxide
  - CO is the reducing agent and NO is the oxidising agent
- **(4) iron(III) oxide + aluminium ==> aluminium oxide + iron** (the Thermit reaction)
  - **Fe_2O_3(s) + 2Al(s)** ==> **Al_2O_3(s) + 2Fe(s)**
  - iron(III) oxide is reduced and is the oxidising agent
  - aluminium is oxidised and is the reducing agent

(h) Redox reaction analysis based on the electron definitions

- **(1) magnesium + iron(II) sulphate** ==> **magnesium sulphate + iron**
  - **Mg(s) + $FeSO_{4(aq)}$** ==> **$MgSO_{4(aq)}$ + Fe(s)**
  - this is the 'ordinary molecular' **equation for a typical metal displacement reaction**, but this does not really show what happens in terms of atoms, ions and electrons, so we use **ionic equations** like the one shown below.
  - The sulphate ion **SO_4^{2-}**(aq) is called a **spectator ion**, because it doesn't change in the reaction and can be omitted from the ionic equation. No electrons show up in the full equations because electrons lost by x = electrons gained by y!!
  - **magnesium + iron(II) ion** ==> **magnesium ion + iron**
  - **Mg(s) + Fe^{2+}(aq)** ==> **Mg^{2+}(aq) + Fe(s)**
  - the magnesium atom loses 2 electrons (oxidation) to form the magnesium ion, the iron(II) ion gains 2 electrons (reduced) to form iron atoms.
  - Mg is the reducing agent (electron donor) and the Fe^{2+} is the oxidising agent (electron remover or acceptor)
  - **Displacement reactions** involving metals and metal ions **are electron transfer reactions**.
- **(2) zinc + hydrochloric acid** ==> **zinc chloride + hydrogen**
  - **Zn(s) + 2HCl(aq)** ==> **$ZnCl_{2(aq)}$ + $H_{2(g)}$**
  - the chloride ion Cl^- is the spectator ion
  - **zinc + hydrogen ion** ==> **zinc ion + hydrogen**
  - **Zn(s) + 2H^+(aq)** ==> **Zn^{2+}(aq) + H_2(g)**
  - Zinc atoms are oxidised to zinc ions by electron loss, so zinc is the reducing agent (electron donor)
  - hydrogen ions are the oxidising agent (gaining the electrons) and are reduced to form hydrogen molecules

- **(3) copper + silver nitrate** ==> **silver + copper(II) nitrate**
  - **$Cu(s) + 2AgNO_3(aq)$** ==> **$2Ag + Cu(NO_3)_2(aq)$**
  - the nitrate ion NO_3^- is the spectator ion
  - **copper + silver ion** ==> **silver + copper(II) ion**
  - **$Cu(s) + 2Ag^+(aq)$** ==> **$2Ag(s) + Cu^{2+}(aq)$**
  - copper atoms are oxidised by the silver ion by electron loss
  - electrons are transferred from the copper atoms to the silver ions, which are reduced
  - the silver ions are the oxidising agent and the copper atoms are the reducing agent
- **(4) iron(II) chloride + chlorine** ==> **iron(III) chloride**
- **(5) halogen** (more reactive) **+ halide salt** (of less reactive halogen) **==> halide salt** (of more reactive halogen) **+ halogen** (less reactive)
  - $\mathbf{X}_{2(aq)} + 2K\mathbf{Y}_{(aq)}$ ==> $2K\mathbf{X}_{(aq)} + \mathbf{Y}_{2(aq)}$
  - $\mathbf{X}_{2(aq)} + 2\mathbf{Y}^-_{(aq)}$ ==> $2\mathbf{X}^-_{(aq)} + \mathbf{Y}_{2(aq)}$
  - where halogen **X** is more reactive than halogen **Y**, F > Cl > Br > I
  - **X** is the **oxidising agent** (electron acceptor)
  - K**Y** is the **reducing agent** (electron donor)

- **The use of Roman Numerals in names**:
  - **This indicates** what is called **the oxidation state of an atom** in a molecule or ion.
  - In simple cases the oxidation number equals the number of electrons to be added (+) or electrons removed (-) to give the neutral element atom. In simple terms, the charge of an atom equals its oxidation number.
  - Neutral atoms (non charged atoms) have the oxidation number of zero (0).
  - It is easy to follow for simple metal ions because it equals the charge on the ion
    - e.g. the **oxidation state** of copper in the copper(II) ion is referred to as **+2**
    - the more electrons removed from the atom or ion by oxidation, the higher its oxidation state
    - e.g. **Fe^{2+} - e^- ==> Fe^{3+}**, iron (II) loses one electron and changes to iron(III) . This gives iron the oxidation state of +3 in the iron(III) ion
      - (via a suitable oxidising agent).
    - **but** for more complex ions things are not so simple and its not appropriate to explain them here.
      - in manganate(VII) ion, the Mn is in the +7 oxidation state
      - in dichromate(VI) ion, the Cr is in the +6 oxidation state

$Cu(s) + 2Ag^+(aq)$ ==> $2Ag(s) + Cu^{2+}(aq)$

✓ The oxidation number of copper on the reactant is zero (0) while on the product it is +2. Since the oxidation number has increased from 0 to +2, **copper has been oxidized**.

- ✓ On the other hand, the oxidation number of silver on the reactant is +2 while on the product is zero (0). Since there is decrease in the oxidation number from +2 to 0, **silver has been reduced**.

Oxidized

$Mg_{(s)} + Fe^{2+}_{(aq)} ==> Mg^{2+}_{(aq)} + Fe_{(s)}$

Reduced

- **but** for more complex ions things are not so simple and its not appropriate to explain them here.
    - in manganate(VII) ion, the Mn is in the +7 oxidation state
    - in dichromate(VI) ion, the Cr is in the +6 oxidation state