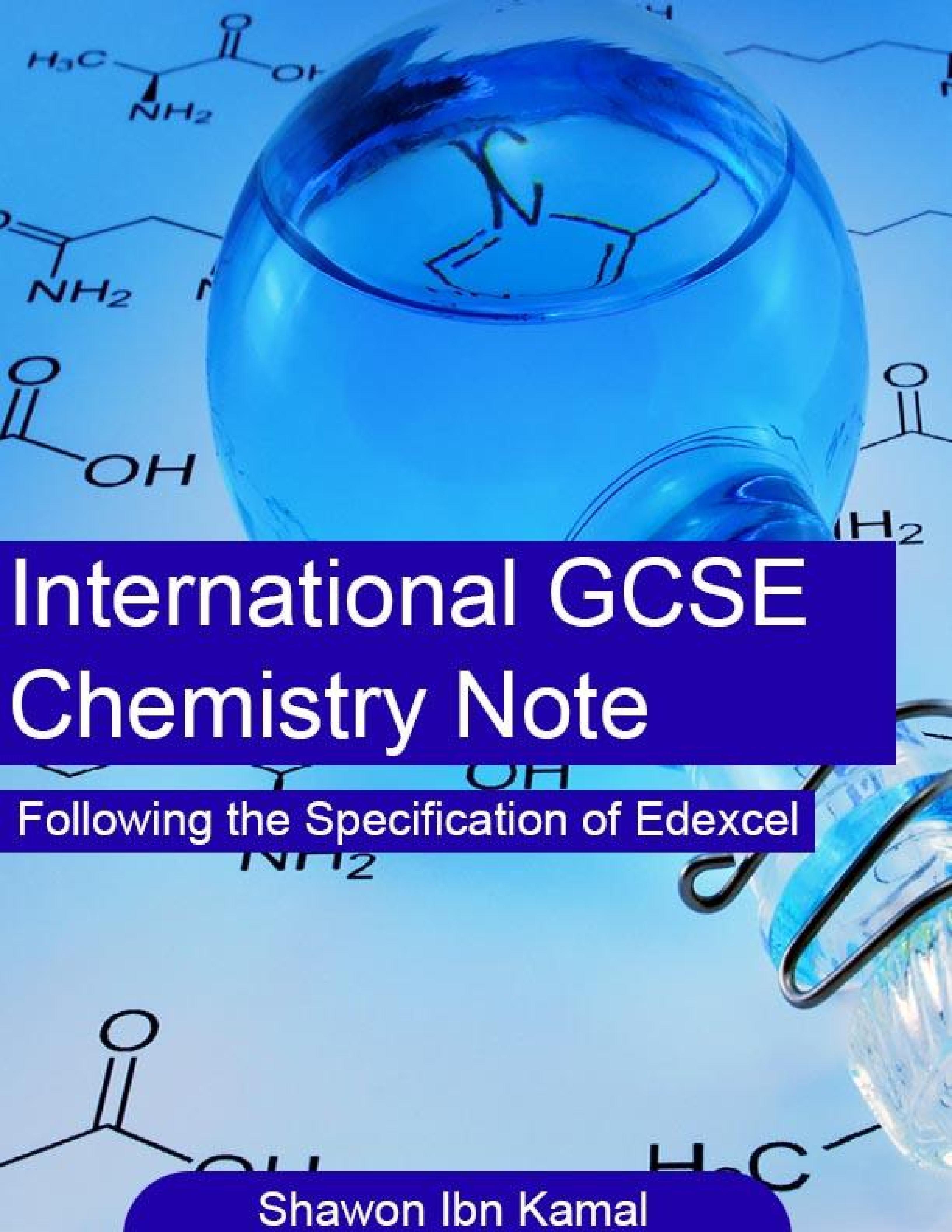


International GCSE Chemistry Note

Following the Specification of Edexcel

Shawon Ibn Kamal

International GCSE Chemistry Note

Following the Specification of Edexcel

By: Shawon Ibn Kamal

Qualification Content:

Section 1: Principles of chemistry

- a) States of matter
- b) Atoms
- c) Atomic structure
- d) Relative formula masses and molar volumes of gases
- e) Chemical formulae and chemical equations
- f) Ionic compounds
- g) Covalent substances
- h) Metallic crystals
- i) Electrolysis

Section 2: Chemistry of the element

- a) The Periodic Table
- b) Group 1 elements – lithium, sodium and potassium
- c) Group 7 elements – chlorine, bromine and iodine
- d) Oxygen and oxides
- e) Hydrogen and water
- f) Reactivity series
- g) Tests for ions and gases

Section 3: Organic chemistry

- a) Introduction
- b) Alkanes
- c) Alkenes
- d) Ethanol

Section 4: Physical chemistry

- a) Acids, alkalis and salts
- b) Energetics
- c) Rates of reaction
- d) Equilibria

Section 5: Chemistry in industry

- a) Extraction and uses of metals
- b) Crude oil
- c) Synthetic polymers
- d) The industrial manufacture of chemicals

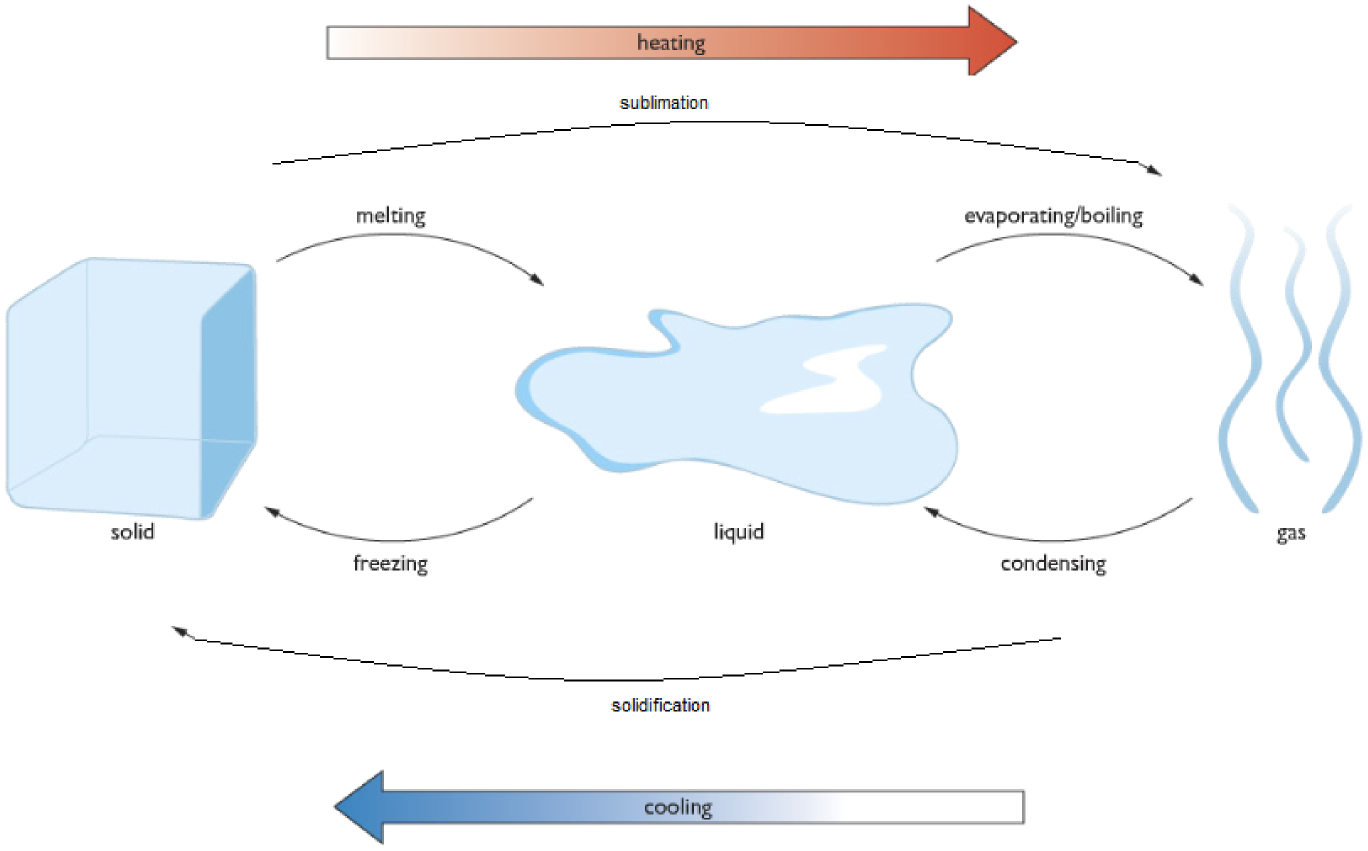
Section 1: Principles of Chemistry

a) States of matter

1.1 understand the arrangement, movement and energy of the particles in each of the three states of matter: solid, liquid and gas

Features	Solid	Liquid	Gas
Arrangement	Regular	Irregular	Random
Movement	Cannot move, vibrate only	Particles can move throughout the liquid slight past each other	Particles can move freely
Energy of Particles	Particles have least kinetic energy	Particles have more kinetic energy than solid	The particles have the most kinetic energy
Distance between particles	Closely packed	Not closely paced	Far apart
Shape	3D structure	Takes the shape of the container	No fixed shape

1.2 understand how the interconversions of solids, liquids and gases are achieved and recall the names used for these interconversions



All types of INTERCONVERSIONS:

- Melting: The process of converting from solid to liquid due to increase of temperature.
- Melting point: The temperature at which a solid starts to melt. Ice melts at 0 degree C.
- Boiling: The process of converting from liquid to gas due to increase of temperature.
- Boiling point: The temperature at which liquid starts to boil. Water boils at 100 degree C.
- Condensation: The process by which a gas turns to a liquid, this process is called condensation.
- Sublimation: The process by which a solid directly turns into a gas without melting.
- Solidification: The process at which gas directly turns to solid.
- Vaporisation: A process by which liquid turns to a gas at its boiling point. (same as boiling)
- Evaporation: A process by which a liquid turns to a gas below its boiling point.
- Volatile: The liquids which evaporates at room temperature that liquids are called volatile.

1.3 explain the changes in arrangement, movement and energy of particles during these interconversions.

- Solid to liquid: When a solid is heated the particles absorb energy and starts to vibrate faster about their fixed position. When the temperature is high enough, the vibration of the particles become sufficient to overcome the attraction between them. The particles begin to break away from their fixed position. The particles can now move slight past each other. The solid is formed liquid.
- Liquid to solid: When a liquid is cooled, the energy is given out by the particles and begins to move slowly. When the temperature is low enough, the particles no longer have the energy to slide over each other. The particles start to settle into a fixed position. When all the particles are settled, the substance's state is solid.
- Liquid to Gas: When a liquid is heated, the particles move fast which break all the forces of attraction in liquid. Bubbles of gaseous particles are formed throughout the whole liquid.
- Gas to liquid: When a gas is cooled the particles eventually move slowly enough that attraction between them hold the particles as liquids. The gas is now condensed to liquid.

b) Atoms

1.4 describe and explain experiments to investigate the small size of particles and their movement including:

- **Dilution of coloured solutions**
- **Diffusion experiments**

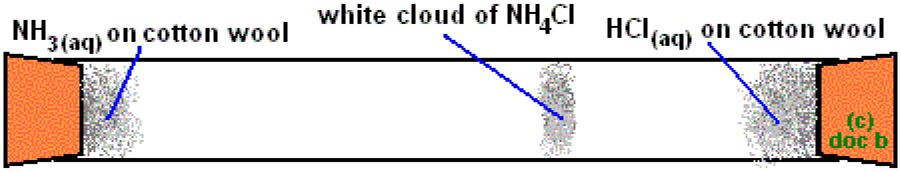
Experiment on dilution

In a beaker of water, a deeply colored substance (Potassium mangnete) is kept. After 10-20 min, the whole solution turned purple. That's because Potassium Mangnete is made of small particles which spread out through the solution. If it was replaced with heavier or bigger particles, it will take long time to diffuse and often it wont.

Diffusion experiment-1

When the glass lid is removed, the brome rises to the top because bromine is made up of small particles which occupies spaces inside both jars by diffusion.

Diffusion experiment-2



From the cotton wool which is soaked in aqueous ammonia, ammonia gas is produced. And from the cotton wool which is soaked in hydrochloric acid, hydrogen chloride gas is produced. Both gases move to each other inside the glass and meet near the cotton wool which is soaked in hydrochloric acid. The gases forms a white smoke of ammonium chloride.

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

This shows that ammonia gas move faster than hydrogen chloride gas in the same length of time. Because ammonia is lighter than hydrogen chloride gas. So ammonia diffuses faster.

The factors that affects the diffusion process are:-
(i) Molecular Mass(Mr)
(ii) Temperature

The molecules with low molecular mass diffuse faster than the higher molecular masses.
Increase in temperature increases the diffusion rate.

1.5 understand the terms atom and molecule

Atoms	Molecules
The smallest particle of an element is called atom.	The smallest particles of an element or compound which can exist independent is called molecules.
Atoms consists of electron, proton and neutron.	Molecule are made up of different types of atom.

Atom can take part in chemical reactions directly.	Molecules cannot take part in chemical reaction without breaking up into atoms.
There are limited types of atom in the earth.	There are unlimited types of molecules.

1.6 understand the differences between elements, compounds and mixtures

Difference between element and compound

Element	**Compound**
Element is a pure substance that cannot be split into simplier substances by chemical process or electricity.	A compound is a pure substance which contain two or more different types of element.
The smallest particle of an element is atom.	The smallest particles of an compound is molecule.
there are limited types of element	There are unlimited types of compound.
The elements are expressed by using chemical symbol.	The compounds are expressed by using chemical formula.

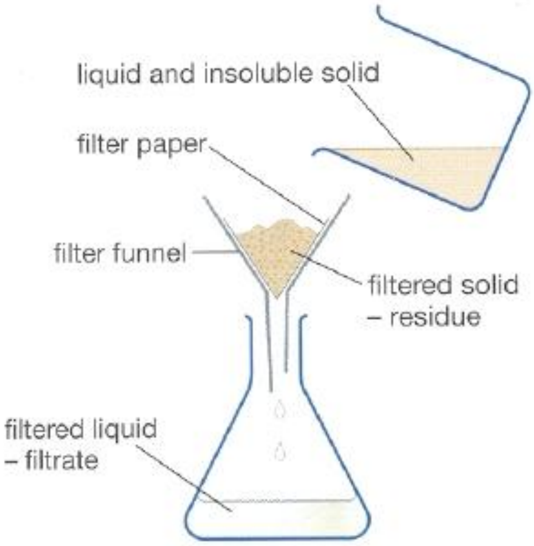
Difference between compound and molecules

Compound	Mixture
Compounds are made up of two or more elements chemically combined.	A gathering of two or more substances without chemical bond.
To make a compound, chemical reactions are needed.	To make a mixture chemical reactions are not needed.
Compounds cannot be separated by physical method of separation.	Mixture can be separated by physical method of separation.
In compounds there must be a fixed proportion of substance.	There is no fixed proportion of substance in mixture.
To make compounds energy changes take place.	To make mixture no energy changes take place.

1.7 describe experimental techniques for the separation of mixtures, including simple distillation, fractional distillation, filtration, crystallisation and paper chromatography

i) Filtration:

A suspension or insoluble solid-liquid mixture can be separated by this process.



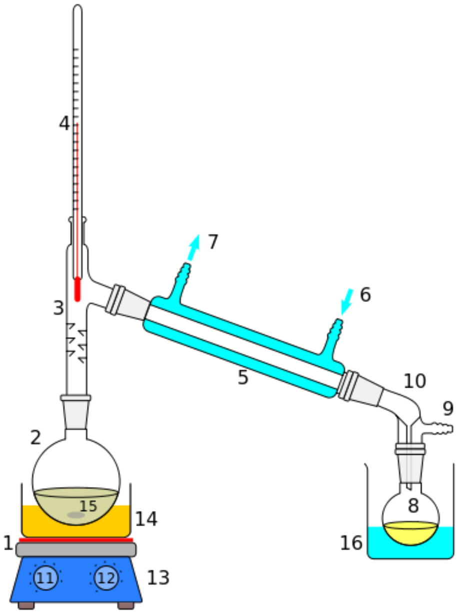
The suspended mixture is poured into the filter funnel. The filter paper has tiny holes, which allows the liquid molecules to pass through. But the solid particles get trapped on the filter paper as residue. The liquid is collected below the funnel as filtrate.

Uses:

- Precipitate from a reaction mixture can be separated by filtration.

ii) Simple Distillation:

Soluble solid can be separated from a solution by distillation process



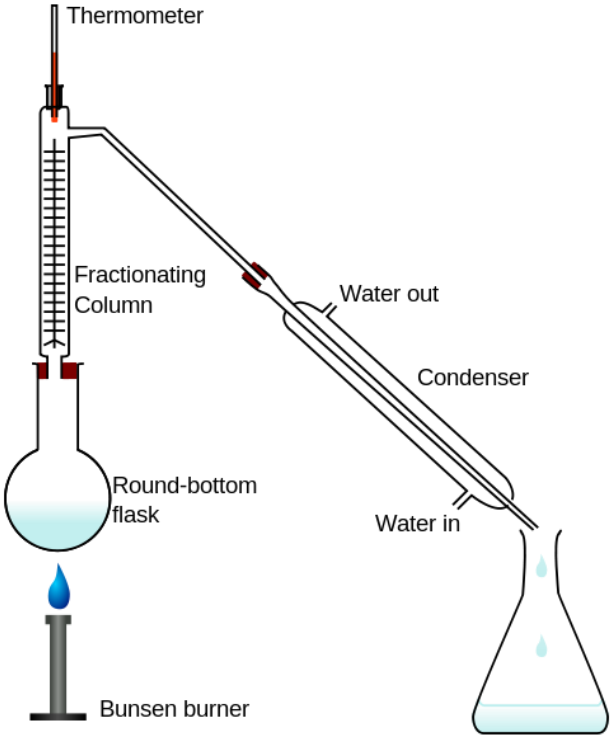
The solution is heated in a round flask. The liquid solvent evaporates and passes through the condenser. While passing, it condenses and drops of liquids are poured in a beaker. The collected liquid is called distillate.

Uses:

- Salt water solution can be separated.

iii) Fractional Distillation:

Two or moremiscible liquid with different boiling points can be separated by fractional distillation.



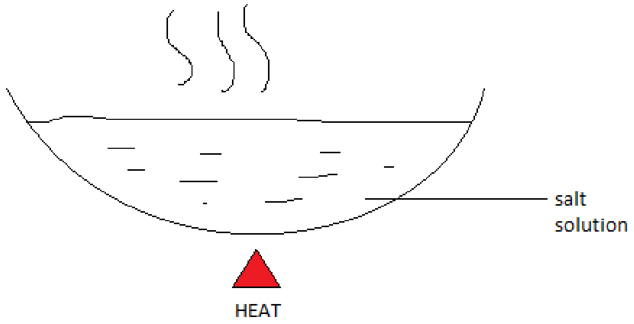
Ethanol water can be separated by this process. The boiling point of ethanol is 78 degree Celsius and boiling point of water is 100 degree Celsius. The temperature of the mixture increases as it is heated. In 78 degree ethanol distills over. The temperature remains constant until all the ethanol has distilled out of the round bottom flask. Then temperature increases until 100 degree Celsius. At this temperature, water starts to boil off. This time heating should be stopped. The ethanol and the water is now separated.

Uses:

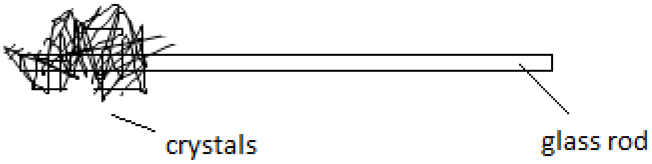
- The fractions can be separated from the crude oil by this process.
- Nitrogen and oxygen can be separated from liquid air.
- Ethanol can be separated from fermented mixture.

v) Crystallization:

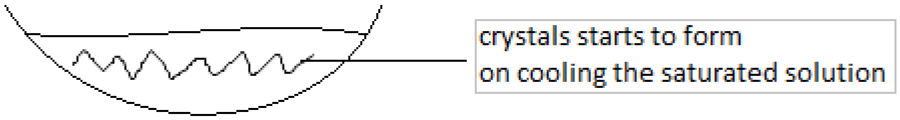
Soluble solid can be separated from a solution by crystallisation. This takes the advantage of different solubility in hot and cold water.



Heat the salt-solution to make it saturated.



Test the solution whether it is saturated or not by dipping a clean dry cold glass rod into the solution. If the solution is saturated, crystals will form in the glass rod. Now cool the solution, so crystals are produced.



Crystals form on cooling because the solubility decreases, and the solid comes out of the liquid as crystals.

Filter the solution to collect the crystals. Then wash the crystals with a liquid cold distilled water to remove impurity. Dry the crystals, keeping it in between few sheets of filter paper.

Paper chromatography:

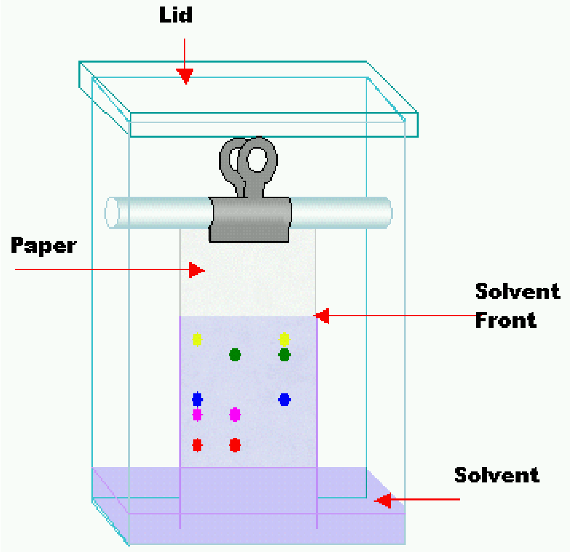
A colored mixture with different solubility can be separated by chromatography.

Follow the steps:

- A small step of ink is placed in the center of a filter paper.
- When the drop has dried up another drop is added in exactly the same spot and this is allowed to dry.
- Ethanol which is a solvent is slowly added drop by drop onto the spot
- The addition of ethanol causes the spot of ink to slowly spread out into different coloured rings.

Ascending chromatography:

To separate the dyes in a mixture using ascending chromatography, we can set up the apparatus as shown below:

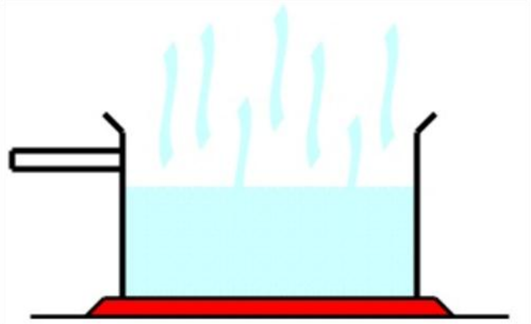


Procedure:

- Draw a pencil line near the end of a chromatography paper.
- Hang the paper inside a glass tank containing a solvent.
- The paper will absorb the solvent and will rise up along with colors.
- Those with higher solubility will spread further than others. Thus the colors will be separated.

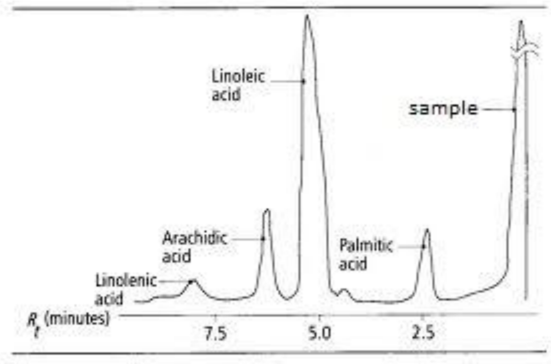
Evaporation

Evaporation is used to separate a soluble solid from a liquid. For example, copper sulfate is soluble in water – its crystals dissolve in water to form copper sulfate solution. During evaporation, the water evaporates away leaving solid copper sulfate crystals behind



1.8 explain how information from chromatograms can be used to identify the composition of a mixture.

Chromatography paper is taken. A sample(unknown) is placed along the pencil line. Other known elements are placed side by the sample in the same line. The sheet of chromatography is coiled into a cylinder and secured with adhesive. It is put in a beaker containing a suitable solvent. The following chromatogram is obtained.

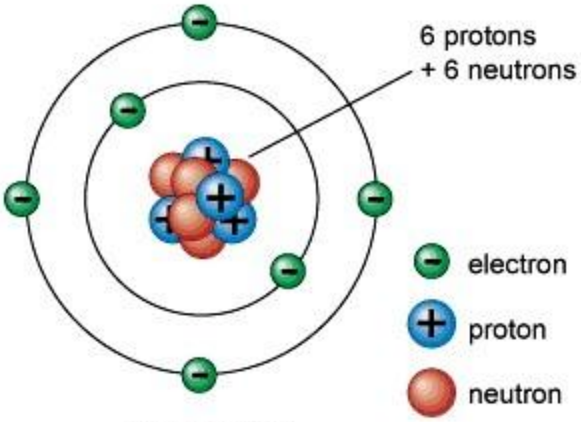


Identical dyes produce spots at the same height in the same color when the same solvent is used. In the diagram above the sample contains linoleic acid.

c) Atomic structure

1.9 understand that atoms consist of a central nucleus, composed of protons and neutrons, surrounded by electrons, orbiting in shells

Atoms are the smallest particle of an element which consists of a nucleus and electron shell. The nucleus is composed of protons and neutrons and the electrons orbit round the nucleus in different shells. A model diagram for an atom is shown below:



Carbon atom

Electron Shells:The electrons move round the nucleus in different energy levels. This energy levels are called electron shells. In first shell two electron can stay. Second shell can hold upto eight electrons. Third shell can hold upto 8 electrons.

1.10 recall the relative mass and relative charge of a proton, neutron and electron

Substance particle	Relative mass	Relative charge	Symbol	Formed in:
Electron	1/1840	-1	e	Electron shells
Protons	1	+1	p	Nucleus
neutrons	1	0	n	nucleus

1.11 understand the terms atomic number, mass number, isotopes and relative atomic mass (Ar)

Atomic number: the number of protons in an atom.

Mass number: The total number of protons and neutrons.

Isotopes: Isotopes are the atoms of the same element with different neutron number.

Relative Atomic Mass(Ar): Relative atomic mass is the weighted average mass of the isotope of the element. It is measured by comparing with the mass of 1/12th of carbon-12.
eg: Chlorine has two isotopes: ${}_{17}{}^{35}Cl$ &${}_{17}{}^{37}Cl$

1.12 calculate the relative atomic mass of an element from the relative abundances of its isotopes

To calculate it use the formula:
Ar = (% x mass number of the isotope) + (% x mass number of the second isotope) and so on….

The natural abundance for chlorine isotopes:
${}^{35}{}_{17}Cl$ is 75% and ${}^{37}{}_{17}Cl$ is 25%

Ar = {(75 x 35)/100} + {(25 x 37)/100}
= 26.5 +9.25
= 35.5

The natural abundance of boron isotopes:
${}^{11}{}_{5}B$ → 80%

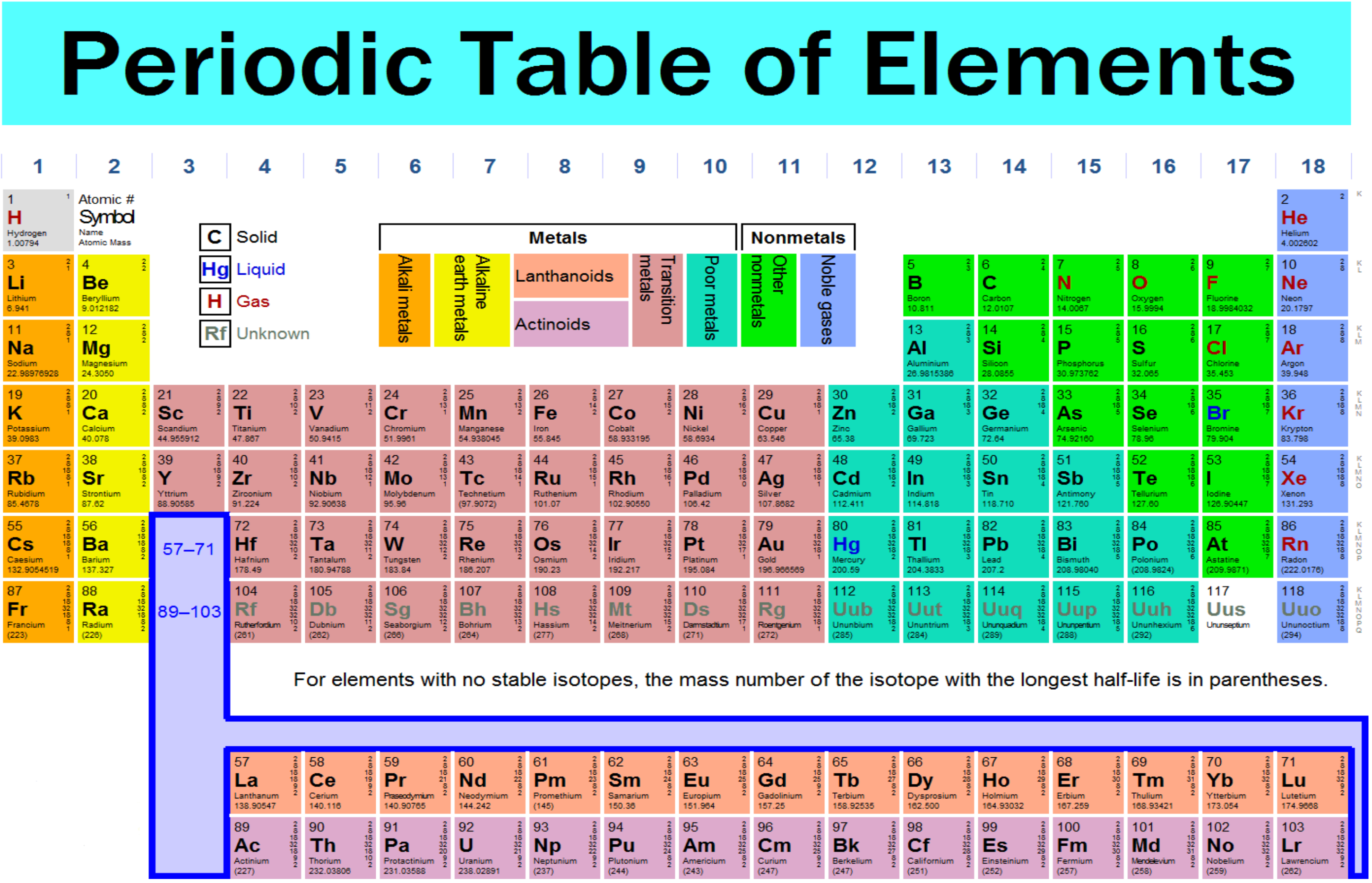
$^{10}_{5}B$ → 20%

Ar = {(80 x 11)/100} + {(20 x 10)/100}

= 8.8+2

= 10.8

1.13 understand that the Periodic Table is an arrangement of elements in order of atomic numb



Periodic Table of Elements

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H Hydrogen 1.00794	Atomic # Symbol Name Atomic Mass																2 He Helium 4.002602
2	3 Li Lithium 6.941	4 Be Beryllium 9.012182											5 B Boron 10.811	6 C Carbon 12.0107	7 N Nitrogen 14.0067	8 O Oxygen 15.9994	9 F Fluorine 18.9984032	10 Ne Neon 20.1797
3	11 Na Sodium 22.98976928	12 Mg Magnesium 24.3050											13 Al Aluminium 26.9815386	14 Si Silicon 28.0855	15 P Phosphorus 30.973762	16 S Sulfur 32.065	17 Cl Chlorine 35.453	18 Ar Argon 39.948
4	19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.955912	22 Ti Titanium 47.867	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938045	26 Fe Iron 55.845	27 Co Cobalt 58.933195	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.64	33 As Arsenic 74.92160	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.798
5	37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90585	40 Zr Zirconium 91.224	41 Nb Niobium 92.90638	42 Mo Molybdenum 95.96	43 Tc Technetium (97.9072)	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.90550	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	52 Te Tellurium 127.60	53 I Iodine 126.90447	54 Xe Xenon 131.293
6	55 Cs Caesium 132.9054519	56 Ba Barium 137.327	57–71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.94788	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.217	78 Pt Platinum 195.084	79 Au Gold 196.966569	80 Hg Mercury 200.59	81 Tl Thallium 204.3833	82 Pb Lead 207.2	83 Bi Bismuth 208.98040	84 Po Polonium (208.9824)	85 At Astatine (209.9871)	86 Rn Radon (222.0176)
7	87 Fr Francium (223)	88 Ra Radium (226)	89–103	104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (277)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (271)	111 Rg Roentgenium (272)	112 Uub Ununbium (285)	113 Uut Ununtrium (284)	114 Uuq Ununquadium (289)	115 Uup Ununpentium (288)	116 Uuh Ununhexium (292)	117 Uus Ununseptium	118 Uuo Ununoctium (294)

C Solid
Hg Liquid
H Gas
Rf Unknown

Metals: Alkali metals, Alkaline earth metals, Lanthanoids, Actinoids, Transition metals, Poor metals
Nonmetals: Other nonmetals, Noble gases

For elements with no stable isotopes, the mass number of the isotope with the longest half-life is in parentheses.

57 La Lanthanum 138.90547	58 Ce Cerium 140.116	59 Pr Praseodymium 140.90765	60 Nd Neodymium 144.242	61 Pm Promethium (145)	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.92535	66 Dy Dysprosium 162.500	67 Ho Holmium 164.93032	68 Er Erbium 167.259	69 Tm Thulium 168.93421	70 Yb Ytterbium 173.054	71 Lu Lutetium 174.9668
89 Ac Actinium (227)	90 Th Thorium 232.03806	91 Pa Protactinium 231.03588	92 U Uranium 238.02891	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)

Periodic Table is an arrangement of elements in order of atomic numbers. It consists of 8 groups and 7 periods.

1.14 deduce the electronic configurations of the first 20 elements from their positions in the Periodic Table

Electronic configuration can be easily figured out by following few rules:

- 1st shell must contain only two electrons
- 2nd shell must contain only 8 electrons
- 3rd shell must contain only 8 electrons
- 4th shell must contain only 18 electrons
- Atomic number is equal to electrons in an “atom”

Example:

Na(Sodium) : 2,8,1
Ca(Calcium): 2,8,8,2

1.15 deduce the number of outer electrons in a main group element from its position in the Periodic Table.

The group number determines outer electron of an element.

Example:

Group I elements have 1 electrons in their outer shell

Group 7 elements have 7
Group 0 have 8(since 8 comes after 7)

d) Relative formula masses and molar volumes of gases

1.16 calculate relative formula masses (Mr) from relative atomic masses (Ar)

The total mass of the molecule calculated from the relative atomic masses of the atoms present in the compound is called relative formula mass.

Example:

H_2O
$=(1 \times 2) + 16$
$=18$

$(NH_4)_2CO_3$
$=(22 + 1 \times 4) \times 2 + 12 + (16 \times 2)$
$=96$

1.17 understand the use of the term mole to represent the amount of substance

A mole is a certain amount of substance. It is a general term to describe an amount of atoms, ions, or molecules.

1.18 understand the term mole as the Avogadro number of particles (atoms, molecules, formulae, ions or electrons) in a substance

A mole is defined as the amount of substance with contains the Avogadro number of particles.

Avogadro Number = 6×10^{23}

1.19 carry out mole calculations using relative atomic mass (Ar) and relative formula mass (Mr)

$$\boldsymbol{mole = \frac{mass\ in\ grams}{MrorAr}}$$

Mole = Number of Particles / 6×10^{23}

1.20 understand the term molar volume of a gas and use its values (24 dm3 and 24,000 cm3) at room temperature and pressure (rtp) in calculations.

Volume of one mole of any gas is molar volume. It is $24dm^3$ or $24000\ cm^3$ at R.T.P.

e) Chemical formulae and chemical equations

1.21 write word equations and balanced chemical equations to represent the reactions studied in this specification

A chemical equation which is expressed by using word is called word equation.

Eg: Carbon + oxygen → carbon dioxide

A chemical equation which is expressed by using symbols is called chemical equation.

Eg: $C(s) + O_2\,(g\,) \rightarrow CO_2(g)$

1.22 use the state symbols (s), (l), (g) and (aq) in chemical equations to represent solids, liquids, gases and aqueous solutions respectively

- (s) for solids
- (l) for liquids
- (g) for gases
- (aq) for aqueous solution

1.23 understand how the formulae of simple compounds can be obtained experimentally, including metal oxides, water and salts containing water of crystallisation

Finding the n in $BaCl_2.nH_2O$

Mass of crucible=30.00g
Mass of crucible +barium chloride crystals, $BaCl_2.nH_2O$ =32.44g
Mass of crucible +anhydrous barium chloride, $BaCl_2$ =32.08g
Mass of $BaCl_2$ =32.08–30.00=208g
Mass of water =32.44 – 32.08 = 0.36g

	$Bacl_2$	**H_2O**
Combiningmasses (g)	20.8	0.36
Number of moles	2.08/208	0.36/18
=	0.01	0.02
Ratio of moles	1:2	
Simplest(empirical)formula	$BaCl_2.2H_2O$	

1.24 calculate empirical and molecular formulae from experimental data

	C	**H**
Combiningmasses(g)	85.7	14.3
Number of moles inatoms	85.7/12	14.3/1
=	7.14	14.3
Ratio of moles	1	
Simplest(empirical)formula	C	

Converting empiricalformulae intomolecularformulae

Relative formula mass of the compound was 56.
56/14 = 4
So you need four lots of CH_2
Themolecular formula is C_4H_8.

1.25 calculate reacting masses using experimental data and chemical equations

Calculate the mass of magnesium oxide that can be made by completely burning 6g of magnesium in oxygen.

Equation for reaction:$2Mg+O_2$---->2MgO

1.Calculate the amount, in moles, of magnesium reacted

Ar of Mg is24

Amount of magnesium=(6÷24)=0.25mol

2.Calculate the amount of magnesium oxide formed

The equation tells us that 2 mol of Mg for 2mol of MgO, hence the amount of MgO formed is the same as the amount of Mg reacted.

Amount of MgO formed is 0.25mol

3.Calculate the mass of MgO formed

Mr of MgO+(24+ 16)=40

Massofmagnesiumoxide=(0.25x40)=10g

1.26 calculate percentage yield

The actual yield is the amount of product that's actually there to be used at the end of the manufacturing process.

The predicted yield is the amount that might have been expected if nothing had got lost along the way.

In practice, some product will be lost during the process when purifying the product by filtration or evaporation or when transferring a liquid or when heating.

The percentage yield is a way of comparing the actual yield with the predicted yield. It's calculated using a formula:

Percentage yield = (actual yield x 100%) / predicted yield

1.27 carry out mole calculations using volumes and molar concentrations.

Mole = Volume(cm^3) /1000 x Concentration(mol/dm^3)

Concentration in g/dm^3 = Concentration in mol/dm^3 x Mr

f) Ionic compounds

1.28 describe the formation of ions by the gain or loss of electrons

Ions are charged particles. Atoms become charged by losing or gaining electrons. Atoms that lose electron become positively charged and atoms that gain electrons become negatively charged.

1.29 understand oxidation as the loss of electrons and reduction as the gain of electrons

Oxidation – Oxidation takes place when electrons are lost. Oxygen is added or oxidation no. is increased.

e.g.:

i. $Na - e \rightarrow Na^+$
ii. $2\,Cl - 2e \rightarrow Cl_2$
iii. $Fe^{2+} - e \rightarrow Fe^{3+}$
iv. $2CO + O_2 \rightarrow 2CO_2$

Oxidation State/Number

The valance condition of an element in a compound is called oxidation state. In other words, the applied valancy of an element inside a compound is called oxidation state.

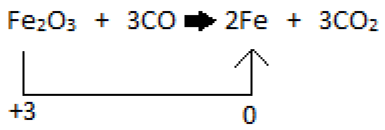
e.g.:

In CO, valancy of C is 2
In CO_2, valancy of C is 4

$KMnO_4 = 0$
$\rightarrow 1 + Mn + 4(-2) = 0$
$\rightarrow Mn - 7 = 0$
$\rightarrow Mn = +7$

Reduction – Reduction takes place when electrons are gained, oxygen is lost. Oxidation number is decreased or hydrogen s added.

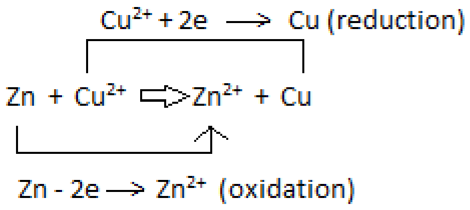
e.g:



Oxidising agent – The substance that oxidise others is called oxidising agent or oxidant. It is reduced after the reaction. The oxidising agent are: H_2SO_4 , HNO_3 , O_3 , Cl_2 , $KMnO_4$, $K_2Cr_2O_7$

Reducing agent – A substance that reduces others is called reducing agent.
e.g.: C, H_2 , CO , metals , H_2S etc.

When in a chemical reaction both oxidation and reduction takes place, then the reaction is called redox.



1.30 recall the charges of common ions in this specification

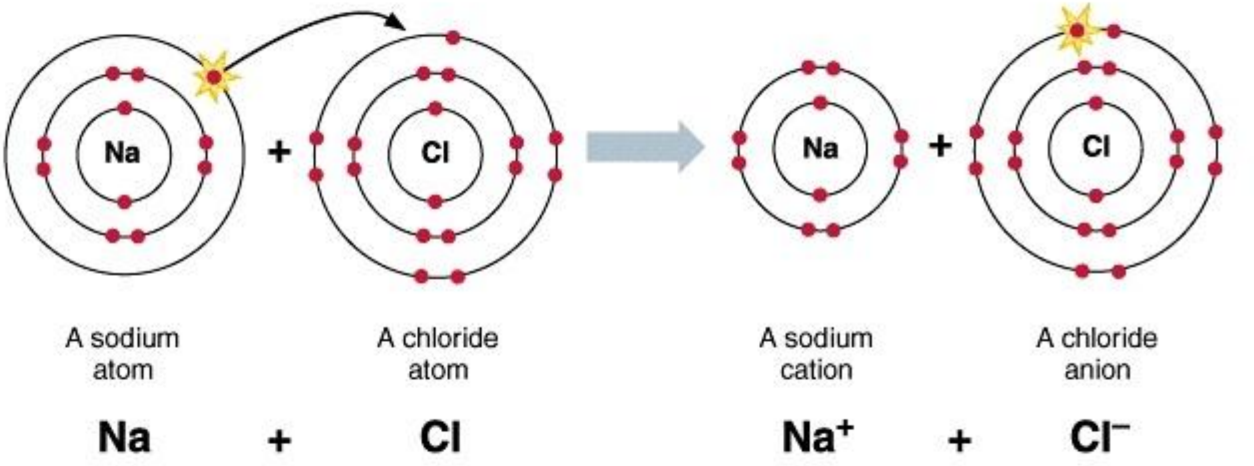
- Group I elements have charge of 1+.
- Group II elements have charge of 2+.
- Group III elements have charge of 3+.
- Group V elements have charge of 3-

1.31 deduce the charge of an ion from the electronic configuration of the atom from which the ion is formed

Here are the steps:

1. Find the number of electrons in the outer electron shell.
2. Find out if it is easy for the atom to gain electron or to donate electron. (in most cases atoms that have below four electrons, donate electrons and atoms that have above 4 electrons, receive electrons)
3. Atoms that gain electron become negative ion and atoms that donate electron forms positive ion.

1.32 explain, using dot and cross diagrams, the formation of ionic compounds by electron transfer, limited to combinations of elements from Groups 1, 2, 3 and 5, 6, 7



1.33 understand ionic bonding as a strong electrostatic attraction between oppositely charged ions

When a metal/metals donates electron, a non-metal/non-metals receives it. The metal becomes positively charged and non-metal becomes negatively charged. As positive and negative charge attract each other, they form a electrostatic force between them. That's how they form ionic compound.

1.34 understand that ionic compounds have high melting and boiling points because of strong electrostatic forces between oppositely charged ions

Melting and boiling point depends on the force that holds the particles. In ionic compounds, the ions are packed tightly for the electrostatic attraction between oppositely charged ions and form a giant structure. So to separate the ions, from the lattice it needs high amount of heat energy to break the millions of bonds, which results high melting and boiling point.

1.35 understand the relationship between ionic charge and the melting point and boiling point of an ionic compound

Ionic charge is directly proportional to the melting and boiling point in an ionic compound.

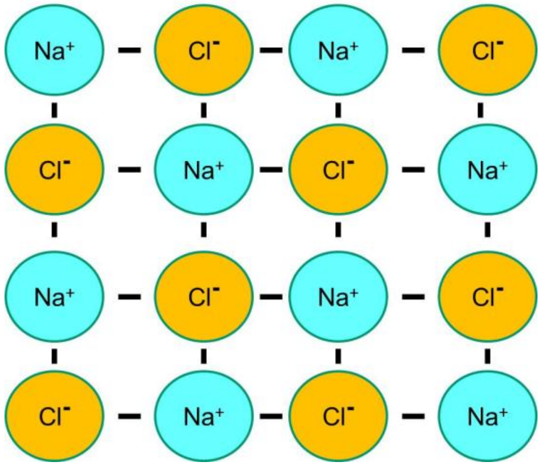
e.g.: Magnesium oxide has higher melting point that NaCl. This is because in MgO, 2+ ions are attracting 2- ions and in sodium chloride, the attraction is weaker because there are only +1 and -1 ions between them.

1.36 describe an ionic crystal as a giant three-dimensional lattice structure held together by the attraction between oppositely charged ions

In ionic compounds the lattice is composed of positive and negative ions in a regular shape. A lattice is a regular array of particles. The lattice is held together by the strong attraction between the positively and negatively charged ions.

1.37draw a diagram to represent the positions of the ions in a crystal of sodium chloride.

Crystal Lattice of NaCl (table salt)



g) Covalent substances

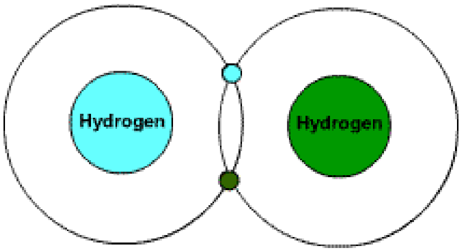
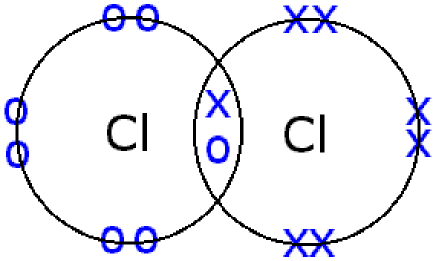
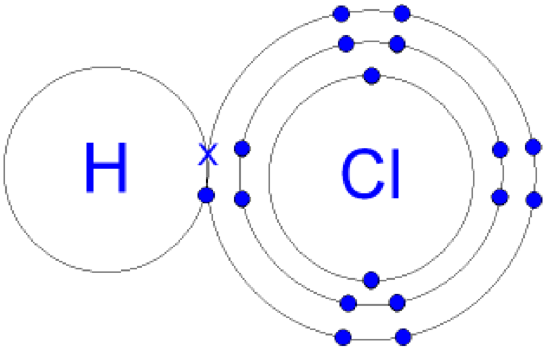
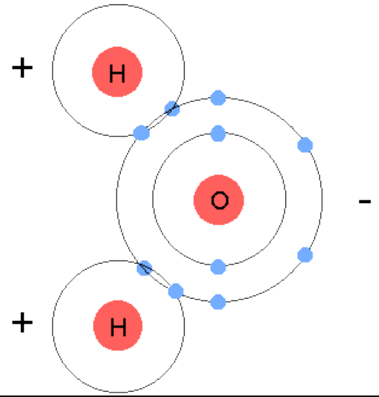
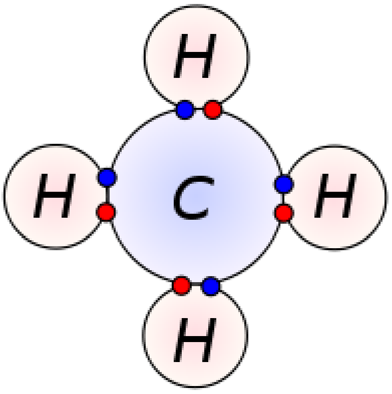
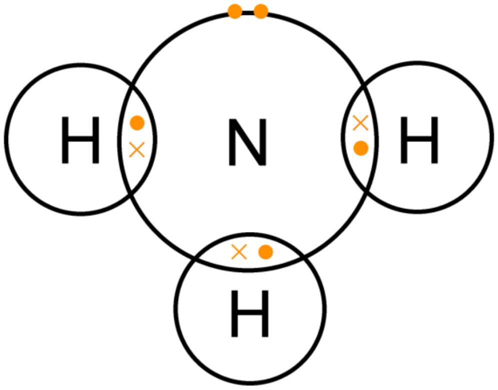
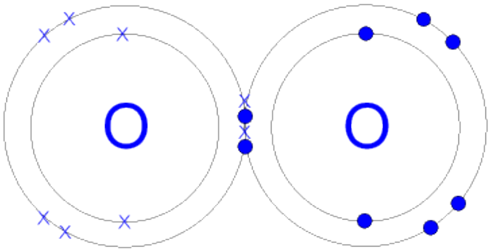
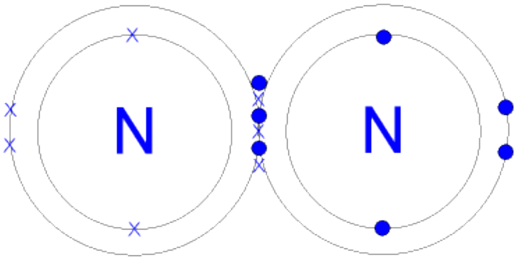
1.38 describe the formation of a covalent bond by the sharing of a pair of electrons between two atoms

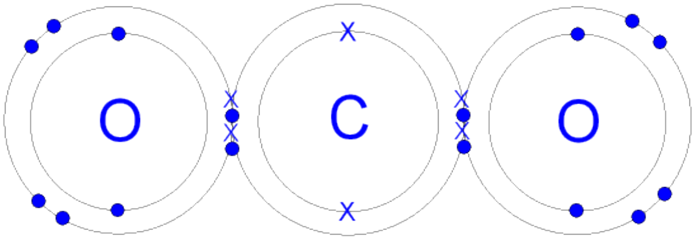
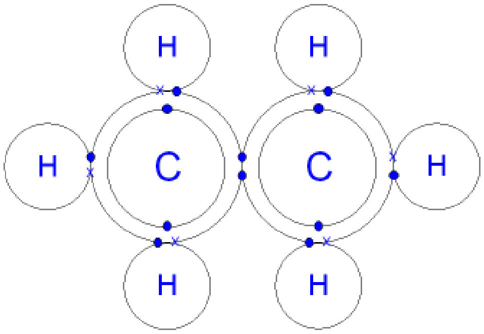
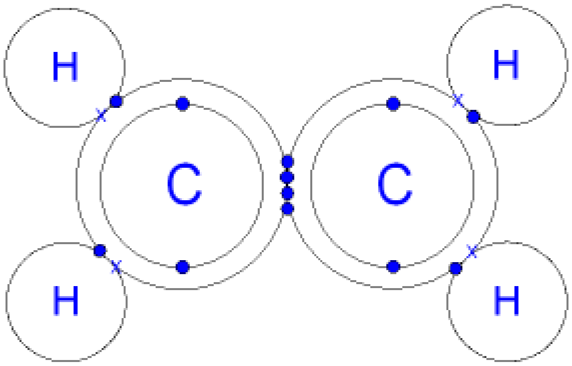
When two or more non-metal share electrons to fulfill their outer electron shell, they form a bond called covalent bond.

1.39 understand covalent bonding as a strong attraction between the bonding pair of electrons and the nuclei of the atoms involved in the bond

When non-metals share electrons, they are attracted by the nucleus. Since nucleus is positively charged it attracts negatively charged electrons.

1.40 explain, using dot and cross diagrams, the formation of covalent compounds by electron sharing for the following substances:

Hydrogen  	**Chlorine** 
Hydrogen chloride  	**Water** 
Methane  ● Electron from hydrogen ● Electron from carbon	**Ammonia** 
Oxygen  	**Nitrogen** 

Carbon dioxide  	**Ethane** 
Ethene  	

1.41 understand that substances with simple molecular structures are gases or liquids, or solids with low melting points

Gases, liquids and gases that have simple molecular structure have low melting points. Reason stated in 1.42

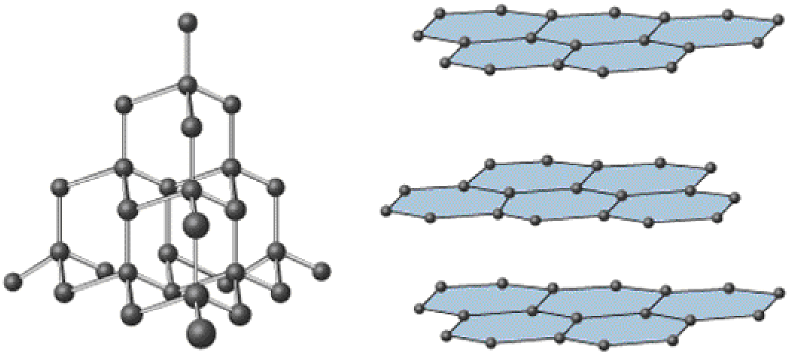
1.42 explain why substances with simple molecular structures have low melting and boiling points in terms of the relatively weak forces between the molecules

The intermolecular attraction between molecules in simple molecular structures is very weak. So less energy is required to break down the forces and hence they have low melting and boiling points.

1.43 explain the high melting and boiling points of substances with giant covalent structures in terms of the breaking of many strong covalent bonds

In giant covalent structures such as diamond, the molecules are held together by billions of strong covalent bond. So a lot of energy is required to break down these bonds and hence they have high melting and boiling point.

1.44 draw diagrams representing the positions of the atoms in diamond and graphite



(a) Diamond (b) Graphite

(a)Diamond: Diamond is one of the allotropes of carbon. Here each carbon atom is joined to four other carbon atoms by strong covalent bonds. It has a giant molecular structure.

(b) Graphite: Graphite is another allotropes of carbon. Within each layer, each carbon atom forms strong covalent bonds with three other carbon atoms. This forms of six carbon atoms that are joined to form two-dimensional flat layers. Therefore, each layer is a giant molecule. The layers of carbon atoms, which are held by weak van der Waal's forces, lie on top of each other.

1.45 explain how the uses of diamond and graphite depend on their structures, limited to graphite as a lubricant and diamond in cutting.

	Diamond	Graphite
Properties	• Hard • Very high melting and boiling points • Non-conductor of electricity	• Soft • Very high melting and boiling points • Conductor of electricity
Uses	• As gemstones • As tips of cutting, grinding and polishing tools	• In pencils • As a dry lubricant • As inert electrodes

Q: Why diamond is hard?/Why it can be used as cutting tools?
Diamonds are used in cutting tools because each carbon is bonded to four other strong covalent bonds and form a giant structure. So to separate, we need to break billions of bond which makes the diamond very hard.

Q: Why diamond do not conduct electricity?
Diamonds do no conduct electricity because each carbon atom has four electrons. In diamonds, these four electrons are engaged in bond formation to other four electrons. So there are no free electrons.

Q: Why graphite can be used as lubricant?
Graphite can be used as lubricants because there are layers in them and the attraction between the layers is weak. So when a force is applied, the layers can slide over each other easily, resulting soft and slippery.

Q: Why graphite is a good conductor of electricity?
In a graphite each carbon atoms are bonded to three other and the fourth electron is free to move through lattice carrying the charge which enables graphite to conduct electricity.

h) Metallic crystals

1.46 understand that a metal can be described as a giant structure of positive ions surrounded by a sea of delocalised electrons

Metal atoms are held strongly to each other by metallic bonding. In the metal lattice, the atoms lose their valence electrons and become positively charged. The valence electrons no longer belong to any metal atom and are said to be delocalised. They move freely between the metal ions like sea of electrons. Hence, this lattice structure is described as a lattice of positive ions surrounded by a "sea of mobile electrons". We can therefore define a metallic bond as the force of attraction between positive metal ions and sea of delocalised electrons.

1.47 explain the electrical conductivity and malleability of a metal in terms of its structure and bonding.

Electrical conductivity: Metals are good at conductors because of the free electron diffusion. When a metal is used in an electrical circuit, electrons entering one end of the metal cause another similar electron to displace from the other end. That's why they are very good conductors.

Malleability and ductility: In metallic bonding, the valence electrons do not belong to any particular metal atom. If sufficient force is applied to the metal, one layer of atom can slide over another without disrupting the metallic bonding. As a result metallic bonds are strong but flexible. Therefore, they can be hammered into different shapes without breaking.

i) Electrolysis

1.48 understand that an electric current is a flow of electrons or ions

Electric current is the flow of electrons.

1.49 understand why covalent compounds do not conduct electricity

In covalent compounds there are no free electrons or ions to move. That's why they don't conduct electricity.

1.50 understand why ionic compounds conduct electricity only when molten or in solution

Ionic compounds normally don't have free moving electrons because they are tightly packed in three dimensional structures. But when they are molten or dissolved in solution they form ions, which allow them to conduct electricity.

1.51 describe experiments to distinguish between electrolytes and non- electrolytes

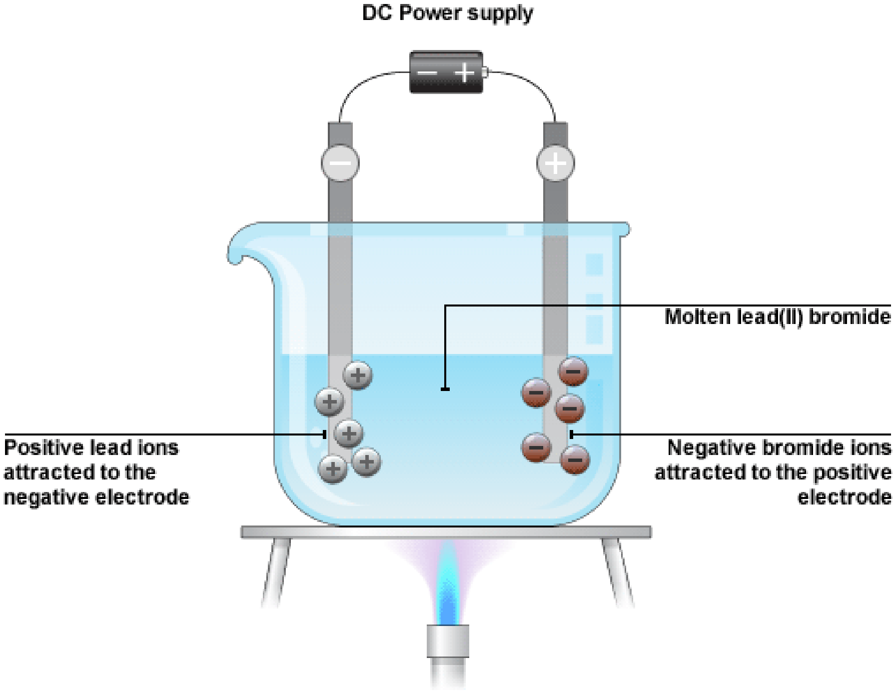
An electrolyte is a substance that conducts electricity.

Experiment:
Take a led bulb and attach the wire at the end of the substance. Now supply electricity from the other end. If the led bulb glows, that substance is an electrolyte or else its non-electrolyte.

1.52 understand that electrolysis involves the formation of new substances when ionic compounds conduct electricity

Electrolysis is the process of using electricity to break down or decompose a compound. The compound is usually dissolved in water or molten.

1.53 describe experiments to investigate electrolysis, using inert electrodes, of molten salts such as lead(II) bromide and predict the products



Lead bromide is heated until it becomes molten. Then it forms two ions: Li^+& Br^-.

The bromide ions moved to the anode and gained electrons and formed bromine molecules.

$2Br^- \rightarrow Br_2 + 2e^-$

The lead ion is positive. So it moves to the negative electrode and by gaining two electron, it forms lead metal which is deposited on the surface of the set.

$Pb^+ + 2e^- \rightarrow Pb(l)$

During the electrolysis, the following observations are seemed:

i. Nothing at all happens until the lead (II) bromide melts. Then the bulb light's up showing electrons are flowing through the wire.
ii. There is bubble around the electrode (anode) as brown bromine gas is given up.
iii. After sometimes metallic red is formed which is found underneath the negative electrode.
iv. When we stop heating the lead (II) bromide solidifies again. Everything stops.

1.54 describe experiments to investigate electrolysis, using inert electrodes, of aqueous solutions such as sodium chloride, copper(II) sulfate and dilute sulfuric acid and predict the products

Electrolysis ofSodium Chloride solution:

In an aqueous solution of sodium chloride contains four different types of ions. They are:

- Ions from sodium chloride – Na^+(aq) and Cl^- (aq)
- Ions from water – H^+(aq) and OH^-(aq)

When dilute sodium chloride solution is electrolysed using inert electrodes, the Na^+ and H^+ ions are attracted to the cathode. The Cl^- and OH^- ions are attracted to the anode.

At the cathode: The Na^+ and H^+ are attracted to the platinium cathode. H^+ ions gain electrons from the cathode to form hydrogen gas.

$2H^+(aq) + 2e^- \rightarrow H_2(g)$

Na^+ remains in the solution.

At the anode:Cl^- and OH^- ions are attracted to the platinum anode. Cl^- ions give up electrons to the anode to form chlorine gas.

$Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

OH^- remains in the solution.

Electrolysis of Aqueous Copper (II) Sulphate Using Inert Electrodes

During electrolysis, the cathode is coated with a layer of reddish-brown solid copper. The blue colour of the solution fades gradually as more copper is deposited. The resulting electrolyte also becomes increasingly acidic.

An aqueous solution of copper(II) sulphate contains four types of ions:
Ions from copper(II) sulphate – Cu^{2+} and SO_4^{2-}
Ions from water – H^+ and OH^-

At the anode: OH^- ions and SO_4^{2-} ions are attracted to the anode. OH^- ions give up electrons more readily than SO_4^{2-} ions.

$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$

The SO_4^{2-} ions remain in solution.

At the cathode: H^+ ions and Cu^{2+} ions are attracted to the cathode. Copper is lower than hydrogen in the reactivity series. Cu^{2+} ions accept electrons more readily than H^+ ions. As a result, Cu^{2+} ions are preferentially discharged as copper metal (atoms).

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

The H^+ ions remain in the solution.

Electrolysis of dilute sulphuric acid

During electrolysis, the cathode is coated with a layer of reddish-brown solid copper. The blue colour of the solution fades gradually as more copper is deposited. The resulting electrolyte also becomes increasingly acidic.

An aqueous solution of copper (II) sulphate contains four types of ions:

Ions from copper (II) sulphate: Cu^{2+} and SO_4^{2-}

Ions from water: H^+ and OH^-

At the anode:

OH^- ions and SO_4^{2-} ions are attracted to the anode. OH^- ions give up electrons more readily than SO_4^{2-} ions. Consequently, OH^- ions are preferentially discharged to give oxygen gas.

$4OH^- (aq) \rightarrow 2H_2O (l) + O_2 (g) + 4e^-$

The SO_4^{2} ions remain in solution.

At the cathode:

H^+ ions and Cu^{2+} ions are attracted to the cathode. Copper is lower than hydrogen in the reactivity series. Cu^{2+} ions accept electrons more readily than H^+ ions. As a result, Cu^{2+} ions are preferentially discharged as copper metal (atoms).

$Cu^{2+} (aq) + 2e^- \rightarrow Cu (s)$

The H^+ ions remain in solution.

1.55 write ionic half-equations representing the reactions at the electrodes during electrolysis

See 1.54

1.56 recall that one faraday represents one mole of electrons

One Faraday is 96500 coulombs. That is the amount of coulombs in one mole of electrons.

1.57calculate the amounts of the products of the electrolysis of molten salts and aqueous solutions.

$$mole = \frac{I \times t}{96000 \times electron\ transfer}$$

$$mole = \frac{Faraday}{electron\ transfer}$$

Section 2: Chemistry of the elements

a) The Periodic Table

2.1 understand the terms group and period

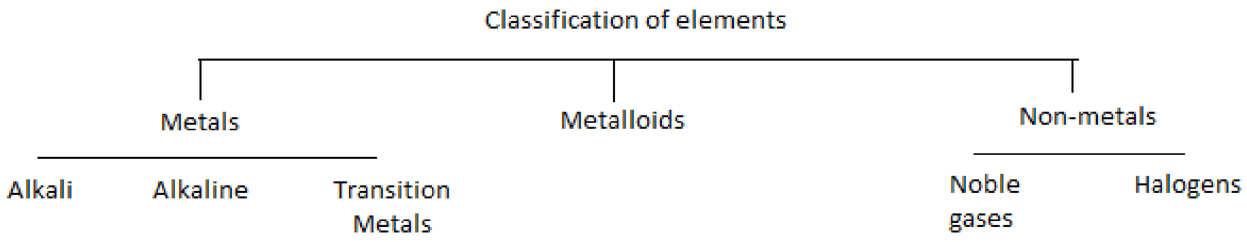
Groups are the vertical columns in the periodic table.

Periods are the horizontal rows in the periodic table.

2.2 recall the positions of metals and non-metals in the Periodic Table

Metals are on the left side of the zig-zag line in the periodic table. Non-metal is on the right.

2.3 explain the classification of elements as metals or non-metals on the basis of their electrical conductivity and the acid-base character of their oxides



Metalloids are the elements of both properties of metals and non-metals, Gallium, Germanium. They exist in the periodic table neat the zig-zag line.

#Difference between metals and non-metals.

Metals:

- Tend to be solids with high melting and boiling points, and with relatively high densities(but as with several of the properties in this list, there are exceptions- like mercury is a liquid.
- Are shiny when they are polished and tend to be easily workable.
- Are good conductors of electricity and heat
- Form positive ions in their compounds
- Have oxides which tend to be basic. This is because when the metallic oxides are dissolved in water, they produce alkalis. Some metallic oxides can show both acidic and alkali property.
  eg: PbO, Al_20_3
- Metal reacts with acids to give a salt and water.

Non-metals:

- Tend to have low melting and boiling points(with the exceptional of carbon and silicon)
- Tend to be brittle as solids and, even if they are crystalline, they don’t have the same sort of shine as metals.
- Don’t usually conduct electricity(with the exceptional of carbon and graphite)
- Are poor conductors of heat
- Tend to form negative ions and covalent compounds

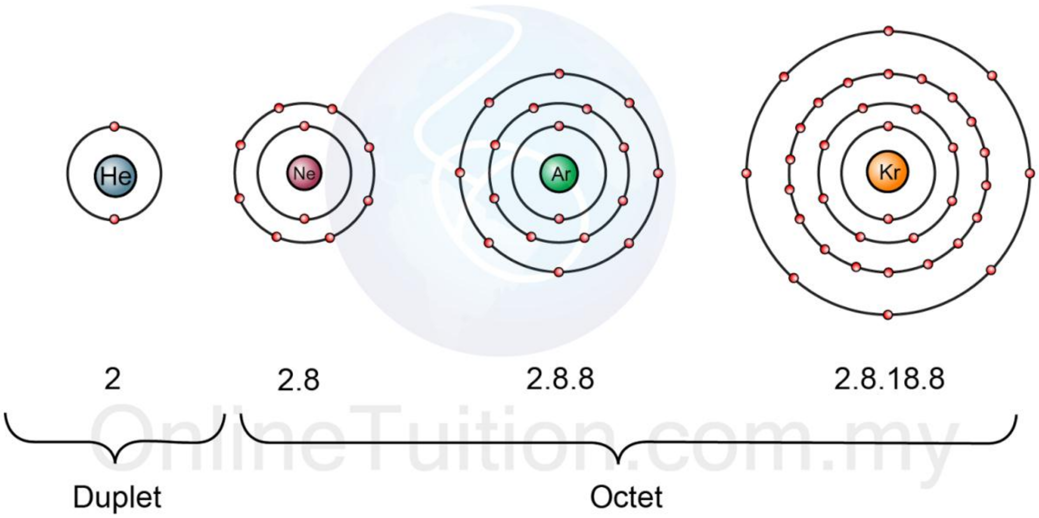
- Non-metallic oxides are acidic.
- Non-metal doesn’t react with acid.

2.4 understand why elements in the same group of the Periodic Table have similar chemical properties

Elements in the same group have similar properties because they have the same number of electrons in their outer electron shell. So when they react, they lose one electron to form single positively charged ions. That’s why they have similar chemical properties though the vigorousness can vary upward and downward position of the group.

2.5 understand that the noble gases (Group 0) are a family of inert gases and explain their lack of reactivity in terms of their electronic configurations.

The noble gases are neon, argon, krypton, xenon and radon. Helium have two electrons in their outer shell which is full and is called duplet electron configuration.



Other noble gases have eight electrons in their outer shell.

The duplet and octet electronic configuration make noble gases inert or unreactive. Because they don’t want to lose or gain electron to become stable, they are already stable. That’s why they are a family of inert gases.

Physical properties of noble gases:

- Colorless gases
- Monatomic: ie. They consist of only one atom
- Densities and boiling points are in a trend. They increases if you go down.

Chemical properties of noble gases:

- They do not react to form ions and so don’t produce ionic compounds. Because they already have their outer shell full and is stable.
- They are unlikely to form covalent compounds. Because it costs too much energy to rearrange the full energy levels to produce the single electron that an atom needs.

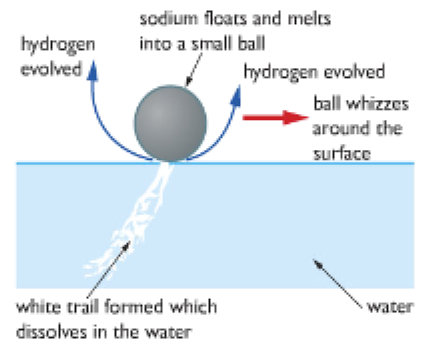
b) Group 1 elements — lithium, sodium and potassium

2.6 describe the reactions of these elements with water and understand that the reactions provide a basis for their recognition as a family of elements

All the group 1 elements react with water to produce a metal hydroxide and hydrogen. The difference is how fast they happen.

Sodium:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$



The sodium floats on water because it is less dense. It melts because its melting point is low and a lot of heat is produced in this reaction. It forms a ball. Because the hydrogen isn't given off symmetrically around the ball, the sodium is pushed around the surface of the water.

A white trail is formed inside the water which is sodium hydroxide, which dissolves to make a strongly alkaline solution. All these metals react with water to produce alkaline metal hydroxides. That's why the Group is called 'alkali metals'.

Lithium:

$$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$$

The reaction is very similar to sodium's reaction, except that it is slower. Lithium's melting point is higher and the heat isn't produced so quickly, so the lithium doesn't melt.

Potassium:

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

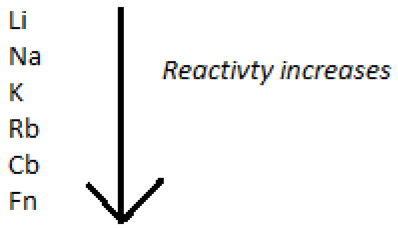
Potassium's reaction is faster than sodium's. Enough heat is produced to ignite the hydrogen, which burns with a lilac flame. The reaction often ends with the potassium spitting around.

Rubidium and caesium:

These react even more violently than potassium, and the reaction can be explosive. Rubidium hydroxide and caesium hydroxide are formed.

2.7 describe the relative reactivities of the elements in Group 1

The reactivity of group one elements increases down the group.



2.8 explain the relative reactivities of the elements in Group 1 in terms of distance between the outer electrons and the nucleus.

Group 1 elements lose electrons to non-metals to react. As you go down the group, the reactivity of the elements increases. Because the number of electrons increase if we go down. Higher electrons can form more shells. Higher electron shell reduces the attraction between proton and electron. So elements can easily lose electrons.

c) Group 7 elements — chlorine, bromine and iodine

2.9 recall the colours and physical states of the elements at room temperature

Element	State in room temperature	Colour At RTP	Colour At Vapour	Colour in Solution
F_2	Gas	Yellow	Yellow	-
Cl_2	Gas	Green	Green	Green – blue
Br_2	Liquid	Red	Red – brown	Orange
I_2	solid	Black	Purple	Dark - brown

2.10 make predictions about the properties of other halogens in this group

- Halogens are oxidising agent.
- They are poisonous.
- They are diatomic.
- They combine with metals to form ionic compounds.
- They get higher melting and boiling points down the group
- They get darker colour down the group.

2.11 understand the difference between hydrogen chloride gas and hydrochloric acid

Hydrogen chloride gas	Hydrochloric acid
It is a gaseous compound of Hydrogen and chlorine gas.	It is an aqueous solution of hydrogen and chlorine.
Have symbol: HCl (g)	Have symbol: HCl(aq)
It exist as a molecule	Here, hydrogen and chloride ions are present.
Hydrogen chloride gas doesn't change the colour of litmus paper when it is dry.	It changes the colour of dry litmus paper.
Hydrogen chloride is acidic in water as it produce hydroxonium ions and chloride ions: $H_2O(l) + HCl(g) \rightarrow H_3O^+(aq) + Cl^-(aq)$	

2.12 explain, in terms of dissociation, why hydrogen chloride is acidic in water but not in methylbenzene

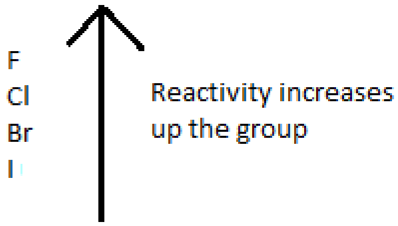
In water, hydrogen chloride dissociates to form hydrogen ion. That's why it shows acidic property. This solution can react with metals to give H_2 gas.

This solution changes the colour of blue litmus paper to red. That means hydrogen chloride is an acid.

Hydrogen chloride in methyl benzene, does not dissociates. It exists as molecule. So it cannot show acid behavior in methyl benzene. That's why this solution does not react with metal carbonates, metal sulphides, and cannot change the colour of litmus.

2.13 describe the relative reactivities of the elements in Group 7

The reactivity falls quickly as you go down the group.



During the reactions, halogens gain one electron to form negatively charged ions. This is because the nucleus attracts the outer shell electron towards it to achieve to fill up its octet. If we go down the group, the distance between the nucleus and the outer shell electron increases. So the attraction for the extra electron decreases. So larger halogens feel less attraction for an extra electron.

2.14 describe experiments to demonstrate that a more reactive halogen will displace a less reactive halogen from a solution of one of its salts

Reacting chlorine with potassium bromide

If you add chlorine solution to colourless potassium solution, the solution becomes orange as bromine is formed.

$2KBr\ (aq) + Cl_2\ (aq) \rightarrow 2KCl\ (aq) + Br_2(aq)$

This is because chlorine is more reactive than bromine. This is because chlorine displaces bromine.

Reacting chlorine with potassium iodide

Adding chlorine solution to potassium iodide solution gives a dark reddish-brown solution of iodine.

$2KI\ (aq) + Cl_2(aq) \rightarrow 2KCl(aq) + I_2(aq\ or\ s)$

2.15 understand these displacement reactions as redox reactions.

In the reaction between chlorine and potassium bromide, bromine loosed electrons. That means it is oxidized. Chlorine gained electron which means it is reduced. When both oxidation and reduction takes place in the same reaction, it is said to be redox reaction.

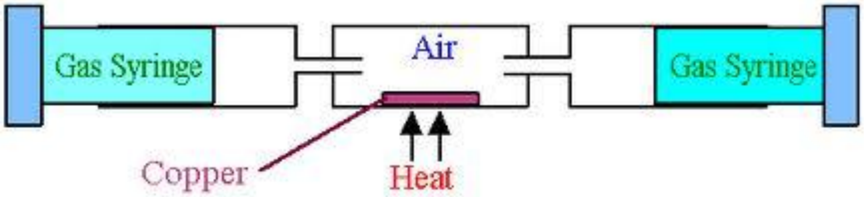
d) Oxygen and oxides

2.16 recall the gases present in air and their approximate percentage by volume

Gas	**Amount in air (%)**	**Amount in air (fraction)**
Nitrogen	78.1	About 4/5
Oxygen	21.0	About 1/5
Argon	0.9	
Carbon dioxide	0.004	

2.17 explain how experiments involving the reactions of elements such as copper, iron and phosphorus with air can be used to investigate the percentage by volume of oxygen in air

Using copper:

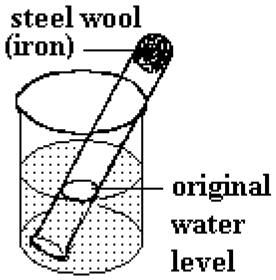


The apparatus originally contains 100cm^3 of air. This is pushed forward and backward over the heated copper, which turns black as copper(II) oxide is formed. The volume of gas reduce as the oxygen is used up.

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

As the copper reacts, the Bunsen is moved along the tube so that it is always heating fresh copper. Eventually all the oxygen in the air is used up. The volume stops reducing. Now there's only 79% of air remaining which proves that air has the percentage composition of 21% in air.

Using rusting or iron:



A test tube is taken and damp iron wool is placed. This is inverted and placed in a beaker containing water. The tube is now left for weeks. Iron uses oxygen and water to form rust. As long as the wool is damp, the rusting will continue.

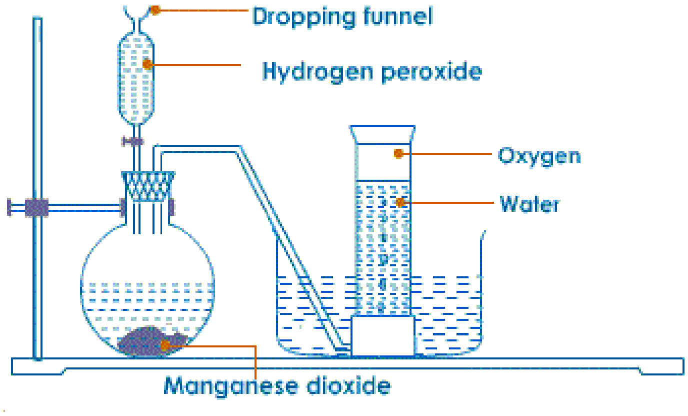
The water level rises as oxygen is used. This is marked with a rubber band. At the end of the experiment, we will see all the oxygen is used up. The original water level was 15cm^3 and now it is 12cm^3. That means 3cm^3 is used up.

$3/15 \times 100 = 20\%$.

2.18 describe the laboratory preparation of oxygen from hydrogen peroxide, using manganese(IV) oxide as a catalyst

Oxygen is made in the lab from hydrogen peroxide solution using manganese(IV) oxide as a catalyst. The reaction is known as catalytic decomposition of hydrogen peroxide.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$



Testing oxygen: Oxygen relights a glowing splint.

2.19 describe the reactions of magnesium, carbon and sulfur with oxygen in air, and the acid-base character of the oxides produced

Burning magnesium

Magnesium burns in air with a bright white flame to give a white, powdery ash of magnesium oxide. The flame is extremely bright in pure oxygen.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Burning carbon

Carbon burns if it is heated very strongly in air or oxygen to give colourless carbon dioxide gas. The carbon may produce a small yellow-orange flame and perhaps some sparks. It depends on the purity of the carbon.

$C(s) + O_2(g) \rightarrow CO_2(g)$

Burning sulfur

Sulfur burns in air with a tiny, almost invisible blue flame. In oxygen it burns much more strongly giving a bright blue flame. Poisonous, colourless sulfur dioxide gas is produced.

$S(s) + O_2(g) \rightarrow SO_2(g)$

Acid base character of oxides:

Few metal oxides react or dissolve in water – to form alkaline solutions. Most metal oxides do not. Shaking a solid magnesium oxide with water doesn't seem to dissolve. However, checking the pH of mixture it proves it is mildly an alkaline.

$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(\text{s and aq})$

Many non-metals react with water to give acidic solutions except water and carbon dioxide. For example sulfur dioxide reacts with water to give sulfurous acid. Sulfurous acid is fairly acidic.

$H_2O(l) + SO_2(g) \rightarrow H_2SO_3(aq)$

2.20 describe the laboratory preparation of carbon dioxide from calcium carbonate and dilute hydrochloric acid

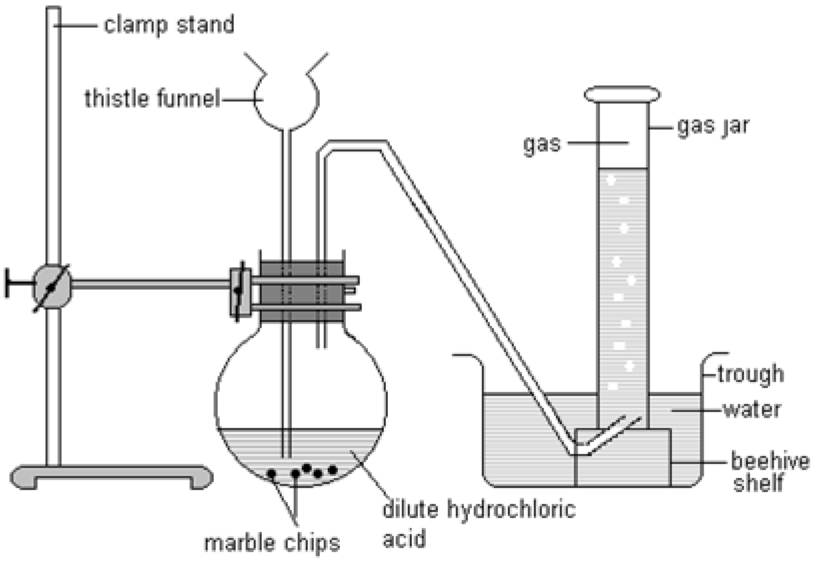


Diagram of Apparatus

Carbon dioxide is most easily made by the reaction between dilute hydrochloric acid and calcium carbonate in the form of marble chips. Carbon dioxide is collected in an inverted jar.

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

Testing carbon dioxide:

Carbon dioxide turns lime water(calcium hydroxide solution) milky. It reacts to give a white precipitate of calcium carbonate.

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

With an excess of carbon dioxide, the precipitate dissolves again to give a colorless solution of calcium hydrogen carbonate.

$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2\ (aq)$

2.21 describe the formation of carbon dioxide from the thermal decomposition of metal carbonates such as copper(II) carbonate

The formation of carbon dioxide by the thermal decomposition of metal carbonate.

K
Na ⇨ These carbonates of the metals do note decompose on heating

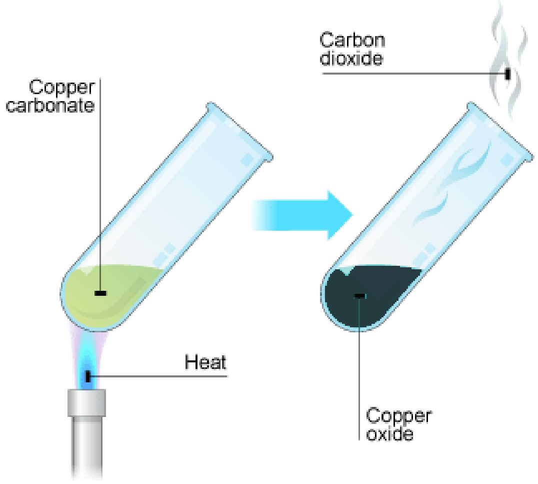
Ca
Mg
Al
Zn
Fe
Pb
Cu

These carbonates can be decomposed by heating to produce metal oxide and carbon dioxide

For example:

Copper carbonate →Copper oxide + Carbon dioxide

$$CuCO_3(green) \rightarrow CuO(black) + CO_2$$



2.22 describe the properties of carbon dioxide, limited to its solubility and density

Properties of Carbon dioxide

- Denser than air
- Colourless
- Odourless
- Does not supper combustion.
- Produces acidic gas reacting with water
- Sparingly soluble in water.

2.23 explain the use of carbon dioxide in carbonating drinks and in fire extinguishers, in terms of its solubility and density

It is used in carbonated drinks because it dissolves in water under pressure. When you open bottle, the pressure falls and the gas bubbles out of solution.

It is also used in fire extinguishers to put out electrical fires, or those caused by burning liquids, where using water could cause problems. The dense gas sinks onto the flames and prevents any more oxygen from reaching them.

2.24 understand that carbon dioxide is a greenhouse gas and may contribute to climate change.

Carbon dioxide is a green house gas.

Carbon dioxide absorbs IR radiation and heats up. As carbon dioxide is present in the air, the air gets warm by absorption of radiation from sunlight. This acts as a greenhouse.

In greenhouse, the heat energy gets trapped inside the house and cannot get out from it. So inside the house, the air is warmer than the outside. Carbon dioxide act as this greenhouse, that's why this gas is called green house gas.

Contribution to climate change:

- Global warming.
- Ice cap melting of N and S pole.
- Flooding in low lying countries.

e) Hydrogen and water

2.25 describe the reactions of dilute hydrochloric and dilute sulfuric acids with magnesium, aluminium, zinc and iron

Magnesium

With dilute hydrochloric acid

$Mg(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

When magnesium react with dilute hydrochloric acid, it forms magnesium chloride and hydrogen gas.

With dilute sulphuric acid

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

When magnesium react with dilute sulphuric acid, it forms magnesium sulphate and hydrogen gas.

Aluminium

With dilute hydrochloric acid

$Al(s) + HCl(aq) \rightarrow AlCl_3(aq) + H_2(g)$

When aluminium react with dilute hydrochloric acid, it forms aluminiumtrichloride and hydrogen gas.

With dilute sulphuric acid

$Al(s) + H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + H_2(g)$

When aluminium react with dilute sulphuric acid, it forms aluminium sulphate and hydrogen gas.

Iron

With dilute hydrochloric acid

$Fe(s) + HCl(aq) \rightarrow FeCl_2(aq) + H_2$

When Iron react with hydrochloric acid, it forms Iron(III) chloride and hydrogen gas.

With dilute sulphuric acid

$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2$

When iron react with sulphuric acid, it forms iron sulphate and hydrogen gas.

2.26 describe the combustion of hydrogen

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

Hydrogen gas burns in air or oxygen to form water(steam).

When hydrogen is pure, hydrogen burns quietly. However, when it is mixed with air or oxygen first, an explosion will occur if a spark or flame is applied.

Testing hydrogen

When a lighted splint is held at the mouth of the test tube that contains hydrogen gas, it will make a popping sound.

2.27 describe the use of anhydrous copper(II) sulfate in the chemical test for water

Water turns white anhydrous copper(II) sulphate blue. Anhydrous copper(II) sulphate lacks water of crystallisation. Dropping water in it replaces the water of crystallisation and therefore colour changes.

$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4x5H_2O(s)$

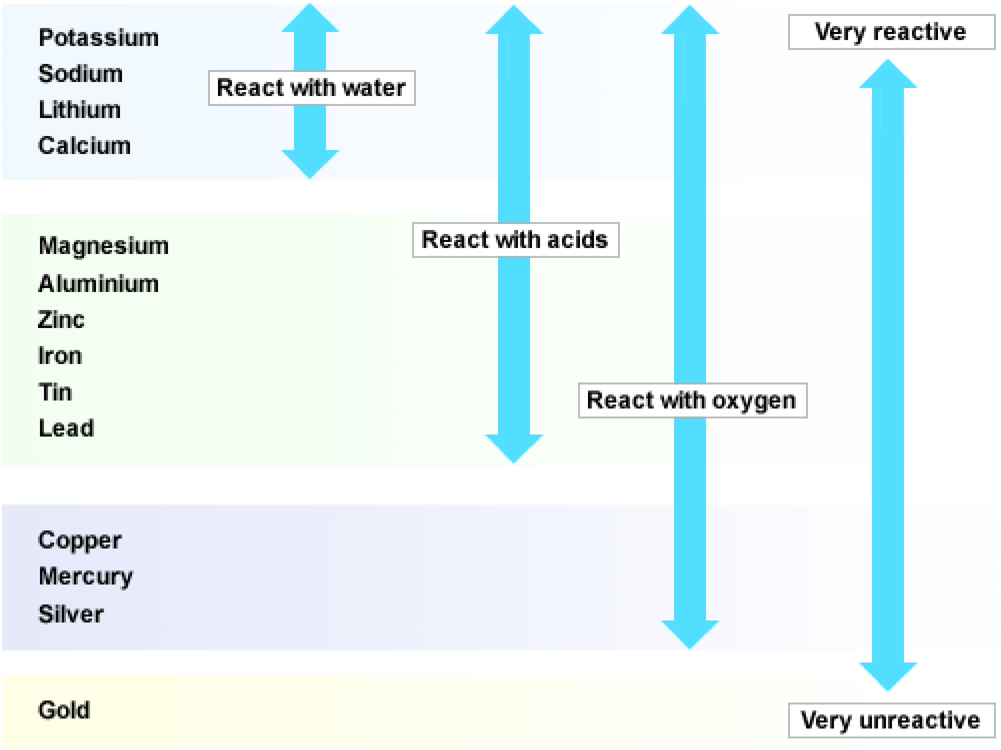
2.28 describe a physical test to show whether water is pure.

If water exactly freezes at 0^{o}C and boils at exactly 100^{o}C at 1 atmospheric pressure, then the water is pure.

f) Reactivity series

2.29 understand that metals can be arranged in a reactivity series based on the reactions of the metals and their compounds: potassium, sodium, lithium, calcium, magnesium, aluminium, zinc, iron, copper, silver and gold

2.30 describe how reactions with water and dilute acids can be used to deduce the following order of reactivity: potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper



Potassium, sodium, lithium, calcium react with water and acids.
Magnesium, aluminium, zinc ,iron ,tin, lead can react with acids too but more slowly. They can't react with water.
Copper can't react with any of them.

The more the vigorous the element is in the reaction, the more reactive the element is and the more things they can react with.

2.31 deduce the position of a metal within the reactivity series using displacement reactions between metals and their oxides, and between metals and their salts in aqueous solutions

If metal is reactive than the metal in metal oxide, the metal will be displaced and reaction will occur in water.
If metal is less reactive than the metal in metal oxide, no reaction will occur.

In this way we can work out their position in reactivity series.

2.32 understand oxidation and reduction as the addition and removal of oxygen respectively

Oxidation: Oxidation takes place during:

i. Addition of oxygen
   e.g: $Cu + O_2 \rightarrow CuO$

ii. Loss of electron.
   e.g: $Ca - 2e \rightarrow Ca^{2+}$

iii. Increase in oxidation state

Oxidation state: The applied valancy of an element in a compound is called an oxidation state.

e.g: Oxidation state of carbon (C) in carbon monoxide (CO) is +2.
Oxidation of carbon (C) in carbon dioxide (CO_2) is +4.

P.S: The total oxidation of a compound is 0

$$\begin{aligned} CO &= 0 \\ C + (-2) &= 0 \\ C - 2 &= 0 \\ C &= +2 \end{aligned}$$

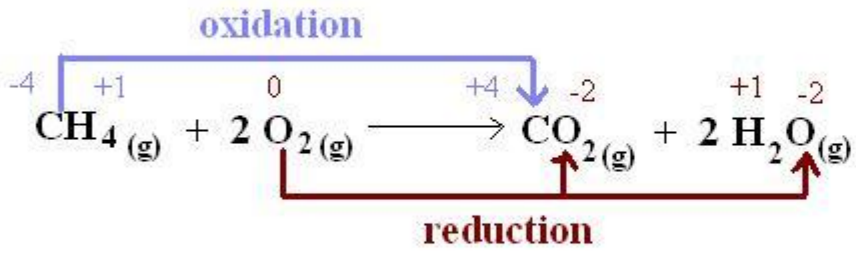
Reduction: Reduction takes place during:

i. Removal of oxygen
ii. Addition of electron
iii. Decrease in oxidation state

2.33 understand the terms redox, oxidising agent, reducing agent

Redox reaction: If in a reaction, both oxidation and reduction take place, it is called redox reaction.

e.g:



$CH_{4(g)} + 2\,O_{2(g)} \longrightarrow CO_{2(g)} + 2\,H_2O_{(g)}$

Oxidising agent/ oxidant: Oxidising agent is a substance that oxidizes others but reduce itself from the reaction.
Reducing agent: Reducing agent is a substance that reduces others but oxides itself from the reaction.

Oxidising agents	Reducing agents
Bromine	carbon
Chlorine	Carbon monoxide
Concentrated sulphuric acid	Hydrogen
Nitric acid	Hydrogen sulphide
Oxygen	Metals
Potassium manganate(VII)	Potassium iodide
Potassium dichromate	Sulphur dioxide
Hydrogen peroxide	Ammonia

2.34 describe the conditions under which iron rusts

What is rusting?

When an object made of iron is exposed to moist air for some time, a reddish-brown substance slowly forms on the surface of the metal. The substance is called rust and have the chemical name hydrated iron (II) oxide.

The process is known as rusting or corrosion of iron.

$4Fe(s) + 3O_2(g) + 2xH_2O(l) \rightarrow 2Fe_2O_3xH_2O(s)$

Conditions under which iron rusts:

- Both air and water are needed for rusting to occur
- The presence of sodium chloride increases the speed of rusting.

2.35 describe how the rusting of iron may be prevented by grease, oil, paint, plastic and galvanising

Grease, oil, paint and plastic prevent air and/or water from coming into contact with iron. This means the reaction that rusts iron can't occur.

Galvanizing is coating in zinc. This Zinc react in the air to form $ZnCO_3$ which prevents air and/or water from coming into contact with the iron.

2.36 understand the sacrificial protection of iron in terms of the reactivity series.

Magnesium or zinc is used as sacrificial protection of iron. If you keep them beside or attach them to iron, magnesium or zinc will corrode instead of iron. Because they are more reactive than iron. As long as magnesium or zinc is present, iron will not rust.

Various methods of rust prevention:

Method	Where it is used	Comment
Painting	Large objects like cars, ships and bridges	If the paint on the metal surface is scratched, rusting will take place under the painted surface.
Oiling or greasing	Tools and machine parts	The protective film of oil or grease gathers dust and must be renewed.
Plastic coating	Kitchenware such as draining racks	If the plastic layer is torn, the iron starts to rust.
Galvanizing (zinc - plating)	Water buckets, dustbins, 'zinc' roofs, kitchen sinks	The metal does not rust even if the zinc layer is damaged. (This is because zinc is more reactive than iron. So zinc corrodes instead of iron).
Tin-plating	Food cans	If the tin layer is scratched, the iron beneath it rusts.
Chrome-plating	Taps, kettles, bicycle handle bars	This gives a bright shiny finish as well as rust protection.
Metal block of zinc or magnesium	Underground pipes, ships, legs of steel piers	Magnesium and zinc corrode in place of iron because they are more reactive metals.
Stainless steel	Cutlery, surgical instruments, pipes in chemical plants	Stainless steel contains chromium and nickel, which do not rust.

g) Tests for ions and gases

2.37 describe tests for the cations:

Li+, Na+, K+, Ca2+ using flame tests

The technique is first of all to clean the end of a piece of platinum or nichrome wire by dipping it onto clean hydrochloric acid and then placing it in a roaring Bunsen flame. This procedure should be repeated until the wire no longer produces a colour in the flame. Then the end of the wire should be dipped again into fresh hydrochloric acid and then into the solid sample under test. The end of the wire should be placed in to a non-roaring, non-luminious Bunsen flame.

This will give this following results:

- Lithum ions : red
- Sodium ions : orange (so strong can mask other colours)
- Potassium ions : lilac
- Calcium ions : brick red
- Sulphur ions : blue
- Carbon ions : orange

NH4+, using sodium hydroxide solution and identifying the ammonia evolved

$$(NH_4)_2SO_4 + NaOH \rightarrow NH_3 + Na_2SO_4 + H_2O$$

Add aqueous sodium hydroxide to the solid or solution, under test and warm the mixture. If ammonium ions are present, then a pungent-smelling gas is produced. The gas produced turns damp red litmus paper blue. It is ammonia, NH_3.

Cu2+, Fe2+ and Fe3+, using sodium hydroxide solution

Of the common hydroxides, only sodium, potassium and ammonium, hydroxides dissolve in water to any extent. Most metal hydroxides are insoluble. That means that if you add sodium hydroxide solution to a solution containing the metal ions, you will get a precipitate of the metal hydroxide.

- Copper(ii) ions + sodium hydroxide → blue precipitate
  When sodium hydroxide is added to compound containing copper ions, it will form a **blue** precipitate.

- Iron(ii) ions + sodium hydroxide →green precipitate
  When sodium hydroxide is added to a compound containing Iron (II) ions, a **green** precipitate of Iron (II) oxide will form.

- Iron(iii)ions + sodium hydroxide → brown precipitate
  When sodium hydroxide is added to a compound containing Iron (III) ions, a**orange-brown** precipitate of Iron (III) oxide is formed.

2.38 describe tests for the anions:

Cl^-, Br^- and I^-, using dilute nitric acid and silver nitrate solution

Make a solution of your suspected chloride, bromide or iodide and add enough nitric acid to make it acidic. Then add few drops of silver nitrate solution.

P.S: The acid is added to react with and remove other substances which might also produce precipitates with silver nitrate solution.

- *Chloride ions + nitric acid + silver nitrate → white precipitate (silver chloride)*
  Chloride ions react with silver nitrate to form a **white** precipitate of silver chloride.

- *Bromide ions + nitric acid + silver nitrate → cream precipitate (silver bromide)*
  Bromide ions react with silver nitrate to from a **cream** precipitate of silver bromide

- *Iodide ions + nitric acid + silver nitrate → yellow precipitate (silver iodide)*
  Iodide ions react with silver nitrate to form a **yellow** precipitate of silver iodide.

SO_4^{2-}, using dilute hydrochloric acid and barium chloride solution

Test: To an aqueous solution of the solid under test, add dilute hydrochloric acid followed by a few drops of barium chloride solution.

Result: This will form a **white** precipitate (of barium sulphate).

$$SO_4^{2-} + HCl + Ba^{2+} \rightarrow \text{white precipitate (barium sulphate)}$$

CO_3^{2-}, using dilute hydrochloric acid and identifying the carbon dioxide evolved

Test: To either the solid, or an aqueous solution of the solid, under test add dilute hydrochloric acid.

P.S: It is best to use dilute nitric acid. Some acid-carbonate combinations can produce an insoluble salt that coats the solid carbonate and stops the reaction, but this doesn't happen if you use nitric acid because all nitrates are soluble.

Result: This will for bubbles of gas (carbon dioxide). This gas will turn limewater milky.

Carbonate + acid → salt + water + carbon dioxide

2.39 describe tests for the gases:

hydrogen

Test: Hold a lighted splint into the gas

Result: It burns with a squeaky pop.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

oxygen

Test: Hold a glowing splint into the gas.

Result: It will relight.

carbon dioxide

Test: Pass the gas through lime water.

Result: It turns white due to the formation of calcium carbonate.

$$CO_2\,(g) + Ca(OH)_2\,(aq) \rightarrow CaCO_3\,(s) + H_2O\,(l)$$

If carbon dioxide is passed in excess, the solution turns clear again due to the formation of soluble calcium hydrogen carbonate.

$$CaCO_3\ (s) + H_2O\ (l) + CO_2\ (g) \rightarrow Ca(HCO_3)_2$$

ammonia

Test: Hold a damp red litmus paper into the gas.
Result: The colour of the paper turns blue.

PS: Ammonia is the only alkaline gas.

Test$_2$: Hold a glass rod soaked in concentrated hydrochloric acid.
Result$_2$: A white fume will form near the glass rod

$$NH_3\ (g) + HCl(g) \rightarrow NH_4Cl$$

chlorine.

Test: Hold a damp blue litmus paper into the gas.
Result: It turns red before it bleaches(turns white).

Colour Chemistry

- Compounds of group I and II are white
- Iron (II) compounds are green
- Iron (III) compounds are red-brown
- Copper (II) compounds are blue.
- Exception: Copper(II) oxide is black,
- Copper (I) oxide is red,
- Anhydrous copper (II) sulphate is white,
- & Copper (II) carbonate is green
- Sulfur is yellow
- Fluorine is yellow
- Chlorine is green – yellow
- Bromine is red
- Iodine is black
- Iodine vapour is purple
- Bromine vapour is red-brown
- Bromine solution is orange
- Iodine solution is dark brown
- Silver chloride is white
- Silver bromide is pale – yellow
- Silver iodide is yellow
- Metal is shiny silver
- Except: Gold is shiny yellow
- & Copper is red brown/pink
- Lead (II) iodide is yellow
- Lead (II) Chloride is white

- Carbon dioxide, Sulfur dioxide, Oxygen, Nitrogen, Ammonia, Noble gases and Carbon monoxide are colourless.
- Nitrogen dioxide is brown.

Section 3: Organic Chemistry

a) Introduction

3.1 explain the terms homologous series, hydrocarbon, saturated, unsaturated, general formula and isomerism.

Organic compound:

The compounds of carbon are called organic compounds. But for conventional reason, the metal carbonate, carbon dioxide and carbon monoxide are not included in organic compounds.

e.g: methane (CH_4), ethanol (C_2H_5OH), carbon tetrachloride (CCl_4), benzene (C_2H_6) etc.

The main branches of organic compounds are:-

i. Aliphatic hydrocarbon
ii. Aromatic hydrocarbon

Homologous series:

It is a family of compounds with same functional group and general formula with similar physical and chemical properties.

Characteristics of homologous series:

- They have a general formula.
- They differ in molecular formula.
- A homologous series have similar chemical properties.
- Their physical properties are in a trend.
- Each homologous series have a functional of group.
- Each member in a homologous series differs in molecular formula from next by CH_2.

Hydrocarbon:

Compounds that are made of only hydrogen and carbon. Hydrocarbons are of two types: Saturated and Unsaturated

Saturated	Unsaturated
Contain C-C single bond(eg: Alkanes)	Contain C=C double bond. (eg: alkenes)
Give substitution reaction	Give addition reaction
A part of saturated compound comparatively is less reactive	A part of unsaturated is more reactive than saturated
Alkanes do not polimarise	Alkene can be polimarised
Saturated hydrocarbons cannot change the colour of bromine water	Unsaturated hydrocarbons can change the colour of bromine water

General formula:

General formula confirms which compound is in which homologous series.

•Alkanes → C_nH_{2n+2}

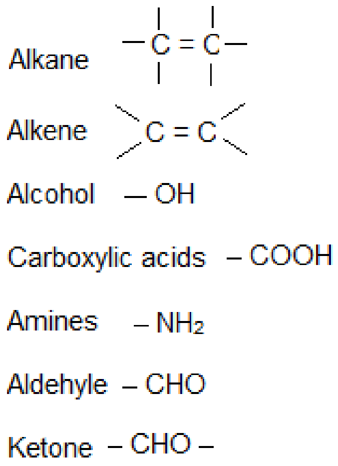
•Alkenes → C_nH_{2n}

•Alcohols → C_nH_{2n+1}+ OH

- Carboxylic acid → C_nH_{2n+1} COOH
- Amines → C_nH_{2n+1} NH_2

Functional Group

Functional group is an atom or group of atom which controls the property of homologus series.



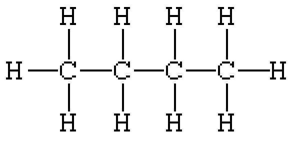
Alkane	$-\mathrm{C}=\mathrm{C}-$
Alkene	$>\mathrm{C}=\mathrm{C}<$
Alcohol	– OH
Carboxylic acids	– COOH
Amines	– NH_2
Aldehyle	– CHO
Ketone	– CHO –

Isomerism:

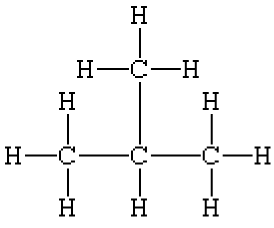
Molecules with identical molecular formulae but with different structural formulae are called isomer.

Example:

Isomers of Butane:

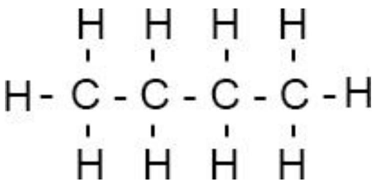


n-butane, $CH_3CH_2CH_2CH_3$



isobutane, $(CH_3)_3CH$

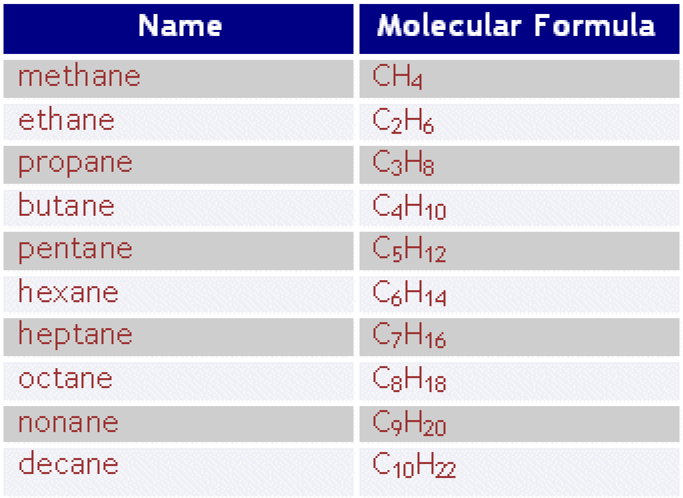
Types of formula

Name	Molecular Formula	Structural Formula	Displayed Formula
Butane	C_4H_{10}	$CH_3CH_2\ CH_2\ CH_3$	

b) Alkanes

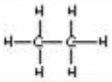
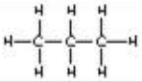
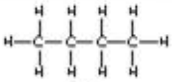
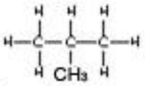
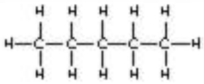
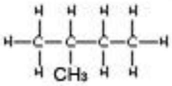
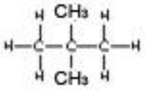
3.2 recall that alkanes have the general formula C_nH_{2n+2}

Alkanes have the general formula C_nH_{2n+2}.n is the number of carbons.



Name	Molecular Formula
methane	CH_4
ethane	C_2H_6
propane	C_3H_8
butane	C_4H_{10}
pentane	C_5H_{12}
hexane	C_6H_{14}
heptane	C_7H_{16}
octane	C_8H_{18}
nonane	C_9H_{20}
decane	$C_{10}H_{22}$

3.3 draw displayed formulae for alkanes with up to five carbon atoms in a molecule, and name the straight-chain isomers

Carbons	Name	Structure
1	methane	
2	ethane	
3	propane	
4	butane	 n - Butane  2 - methyl propane
5	pentane	 n - pentane  2 - methyl butane  2,2 - dimethyl propane

3.4 recall the products of the complete and incomplete combustion of alkanes

If air is present enough, then alkanes burn completely to produce carbon dioxide and water.

eg:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

If the air is not enough then alkanes burn incompletely and produce carbon monoxide or soot (carbon) and water.

$$CH_4 + O_2 \rightarrow CO + H_2O$$

or

$$CH_4 + O_2 \rightarrow C\ (s) + H_2O$$

Complete combustion is more efficient than incomplete combustion and during complete combustion more heat energy is produced.

3.5 describe the substitution reaction of methane with bromine to form bromomethane in the presence of UV light.

If hydrogen atoms of alkanes are replaced by halogen atoms then this type of reaction is called **substitution** reaction.

If chlorine is replaced, it is **chlorination.** If bromine is replaced, it is **bromination.** If halogen is replaced, it is **halogenations**.

Alkanes give substitution reaction by reacting with bromine in the presence of ultra-violet light. A hydrogen atom in the alkane is replaced by a bromine atom.

A mixture of methane and bromine gas is brown because of the presence of bromine. If it is exposed to sunlight, it loses its colour, and a mixture of bromomethane and hydrogen bromide gases is formed.

Methane + chlorine →monochloromethane + hydrogen chloride

CH_4 (g) + Cl_2 (g) → CH_3Cl (g) + HCl (g)

Monochloromethane + chlorine → dichloromethane + hydrogen chloride

CH_3Cl (g) + Cl_2 (g) → CH_2Cl_2 (g) + HCl (g)

Dichloromethane + chlorine →trichloromethane + hydrogen chloride

CH_2Cl_2 (g) + Cl_2 (g) → $CHCl_3$ (l) + HCl (g)

Trichloromethane + chlorine →tetrachlomethane + hydrogen chloride

$CHCl_3$ (l) + Cl_2 (g) → CCl_4 (l) + HCl (g)

Overall reaction:

CH_4 (g) + Cl_2 (g) → CCl_4 (l) + HCl (g)

c) Alkenes

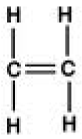
3.6 recall that alkenes have the general formula C_nH_{2n}

Alkenes have the general formula C_nH_{2n}

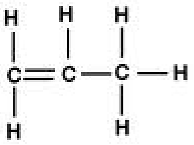
No. of C atoms	Name of alkene	Molecular formula
2	Ethene	C_2H_4
3	Propene	C_3H_6
4	Butene	C_4H_8
5	Pentene	C_5H_{10}
6	Hexene	C_6H_{12}
7	Heptene	C_7H_{14}
8	Octene	C_8H_{16}
9	Nonene	C_9H_{18}
10	Decene	$C_{10}H_{20}$

3.7 draw displayed formulae for alkenes with up to four carbon atoms in a molecule, and name the straight-chain isomers (knowledge of cis- and trans- isomers is not required)

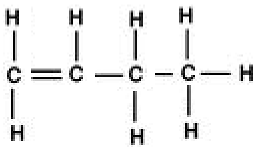
Ethene:



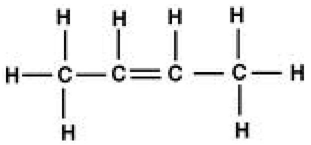
Propene:



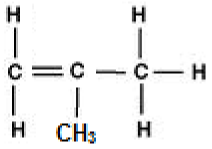
Butene:



But - 1 - ene



But - 2 - ene



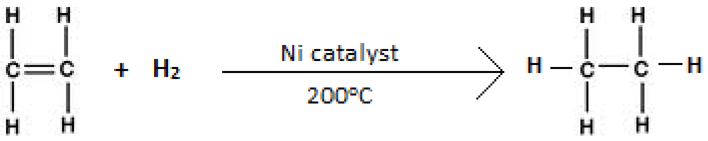
2 - methyl propene

3.8 describe the addition reaction of alkenes with bromine, including the decolourising of bromine water as a test for alkenes.

Addition reaction:

(i) With hydrogen

In the presence of nickel catalyst, alkenes react with hydrogen to produce alkanes. This process is adds hydrogen, therefore it is called hydrogenation.



(ii) With halogen

An alkene will make its double bond into a single bond, to bond to two bromines. Bromine is added to the molecule. This is called halogenations because halogen is added. When alkenes are put in bromine water it turns from brown to colourless.

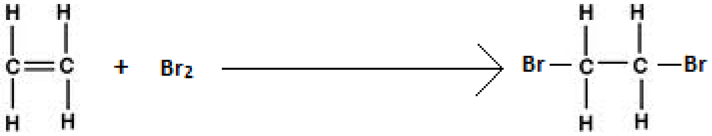
This reaction can be used to check saturation because only alkenes can add bromine but not alkane.

Test: Shake the compound with bromine water solution.

Result: If the compound is unsaturated, then the colour will change from orange to colourless.

For example:

$$C_2H_4(g) + Br_2\ (aq) \rightarrow C_2H_4Br_2\ (aq)$$



d) Ethanol

Ethanol is the second member of a homologus series, alcohol. All alcohol has a functional group –OH. The general formula of alcohol is C_nH_{2n+1} OH.

3.9 describe the manufacture of ethanol by passing ethene and steam over a phosphoric acid catalyst at a temperature of about 300°C and a pressure of about 60–70 atm

Ethanol is manufactured by the catalytic addition of steam to ethane. The mixture of ethane and steam is passed through phosphoric (V) acid at 300^0C and 60 atm. Phosphoric(V) acid (H_3PO_4) acts as a catalyst in this reaction.

$$\text{Ethane + steam} \xrightarrow[300\ C,60\ atm]{phosphoric\ acid} \text{ethanol}$$

$$CH_2 = CH_2(g) + H_2O(g) \xrightarrow[300\ C,60\ atm]{phosphoric\ acid} CH_3CH_2OH$$

3.10 describe the manufacture of ethanol by the fermentation of sugars, for example glucose, at a temperature of about 30°C

Fermentation is a chemical process in which microorganisms such as yeast act on carbohydrates to produce ethanol and carbon dioxide. Sugars and starch are examples of carbohydrates.

Yeast contains zymase enzymes (biological catalysts) which cause starch or sugar to break down to glucose. The glucose is then broken down to ethanol and carbon dioxide.

A glucose solution is mixed with yeast and the mixture is produced after a few days.

The enzymes in yeast work best at around 37^oC. If the temperature is raised beyond 37^o, the enzymes will die and the fermentation stops.

The fermentation of sugars produces only a dilute solution of ethanol which is about 12%.This is because when the alcohol content exceeds this value, the yeast dies and fermentation stops. Then the fermented mixture is filtered to remove the solid impurities and the filtrate is fractionally distilled for further purification.

$$\text{Glucose solution} \xrightarrow{yeast} \text{ethanol + carbon dioxide}$$
$$C_6H_{12}O_6\ (aq) \xrightarrow{yeast} 2C_2H_5OH\ (aq) + 2\ CO_2\ (g)$$

Purification of fermented mixture:

Follow: 1.7(iii)

3.11 evaluate the factors relevant to the choice of method used in the manufacture of ethanol, for example the relative availability of sugar cane and crude oil

	Fermentation	Hydration of ethane
Uses of resources	Uses renewable resources- sugar beet or sugar cane, corn and other starchy materials.	Uses finite resources – once all the oil has been used up there won't be any more
Type of process	A batch process – everything is mixed together in a reaction vessel and then left for several days. That batch is then removed and a new reaction is set up – this is inefficient	A continuous flow process – a stream of reactant is constantly passed over the catalyst. This is more efficient than a batch process
Rate of reaction	Slow, taking several days for each batch	Rapid
Quality of product	Produces very impure ethanol that needs further processing	Produce much purer ethanol
Reaction conditions	Uses gentle temperature and ordinary pressure	Uses high temperature and pressures, needing a high input of energy

3.12 describe the dehydration of ethanol to ethene, using aluminium oxide.

Dehydration refers to the removal of water from a compound. The dehydration of ethanol produces ethane. Ethanol vapour is passed over hot aluminium oxide, acting as a catalyst.

$$C_2H_5OH(g) \rightarrow C_2H_4(g) + H_2O(l)$$

If propanol is used instead of ethanol, then by the loss of water molecules, propene will be produced.

Other reactions of alcohol:

(i) With sodium

Ethanol reacts with sodium to produce sodium ethoxide and hydrogen. This time, sodium sinks in ethanol, (whereas sodium floats in water).

$$C_2H_5OH + Na \rightarrow C_2H_5ONa + H_2$$

(ii) With phosphorus pentachloride

Ethanol reacts with phosphorus (V) chloride to produce chloroethane, phosphorus oxychloride and steamy fumes of hydrogen chloride gas.

$$C_2H_5OH + PCl_5 \rightarrow C_2H_5Cl + POCl_2 + HCl$$

(iii) With acidified potassium dichloride VI (oxidation reaction).

If ethanol is warmed with acidify potassium dichloro VI, then, ethanol oxidises to ethanoic acid.

$$CH_3CH_2OH + |O| \rightarrow CH_3COOH + H_2O$$

Here, the colour changes from orange to green of the reaction mixture.

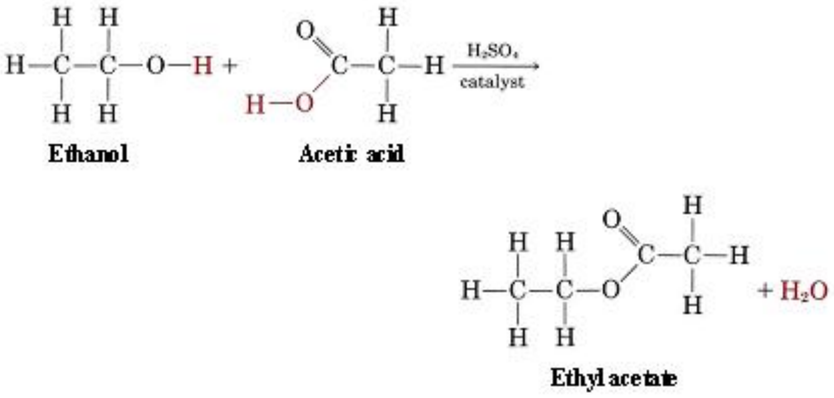
(iv) Esterification

Alcohol + carboxylic acid → ester + water

If ethanol is heated with ethanoic acid in the presence of concentrated acid & catalyst then a ester is formed.

$$CH_3CH_2OH + CH_3COOH \rightarrow CH_3COOCH_2CH_3$$

Ethanol + ethanoic acid →ethylethanoite



(8.4)

Section 4: Physical chemistry

a) Acids, alkalis and salts

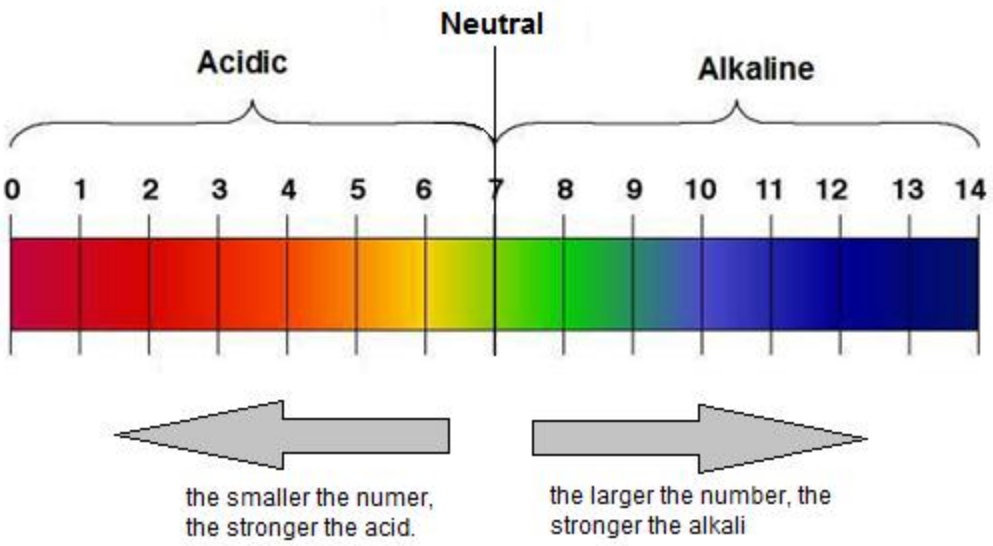
4.1 describe the use of the indicators litmus, phenolphthalein and methyl orange to distinguish between acidic and alkaline solutions

Indicators are mixture of dyes which changes their colour to show whether a solution is acidic or alkaline. Some indicators and their colour in different solutions are given below.

Indicators	Colour in acid	Colour in alkali	Colour in neutral
1. Litmus	Red	Blue	Purple
2. Methyl orange	Red	Yellow	Orange
3. Phenolphthalein	*Colourless*	Pink	*Colourless*
4. Universal indicator	Red	Violet	Green

4.2 understand how the pH scale, from 0–14, can be used to classify solutions as strongly acidic, weakly acidic, neutral, weakly alkaline or strongly alkaline

pH scale: Some acids are weaker than others. It is the same with alkalis. The strength of an acid or alkali is shown using the pH scale which goes from 0 to 14.



On this scale,

- i. An acidic solution has a pH number less than 7.
- ii. An alkaline solution has a pH number greater than 7.
- iii. A neutral solution has a pH number of exactly 7.

Strongly acidic sol^n: pH 0 – 3
Weakly acidic sol^n: pH 3 – 7
Neutral sol^n: ph 7

Weakly alkaline sol^{n}: pH 7 - 11
Strongly acidic sol^{n}: pH 11 - 14

4.3 describe the use of universal indicator to measure the approximate pH value of a solution

Universal indicator is a mixture of several different indicators. Unlike litmus, universal indicator can show us exactly how strongly acidic or alkaline a solution is by the gradual colour change over a range of pH. This is measured using the pH scale. The pH scale runs from pH 0 to pH 14.

pH: 1 – 3: red colour; strong acid
pH: 3 – 7: orange; weak acid
pH: 7: green; neutral
pH: 7 – 11: blue – green; weak alkali
pH: 11 – 14: Violet; strong alkali

4.4 define acids as sources of hydrogen ions, H+, and alkalis as sources of hydroxide ions, OH^-

Acids: Acids are those which provides H^+ ions when they are dissolved in water. E.g:

$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$
$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$
$H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2-}$

$CH_3COOH\ (aq) \rightleftharpoons CH3COO^- + H^+$
$HCOOH\ (aq) \rightleftharpoons HCOO-\ (aq) + H+\ (aq)$

Acids can be of two types –

i. **Strong acid.** Those which ionizes fully in water.
ii. **Weak acid.** Those which ionizes partially/incompletely in water. Eg: Organic acid

If an acid ionizes to give just one hydrogen ion, it is monobasic aid. Similarly, dibasic acid are those which provides hydrogen ions. Tribasic acid provides three hydrogen ions.

Alkali: Alkalis are those which provide hydroxide ions when they are dissolved in water. E.g:

$NaOH \rightarrow Na^+ + OH^-$
$KOH \rightarrow K^+ + OH^-$
$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$

Alkalis can also be strong or weak. Those which ionize completely are strong alkali and those which doesn't ionize completely are weak alkalis.

$NH_3(aq) = NH_4OH \rightleftharpoons NH_4^+ + OH^-$

4.5 predict the products of reactions between dilute hydrochloric, nitric and sulfuric acids; and metals, metal oxides and metal carbonates (excluding the reactions between nitric acid and metals)

Metals:
Metals above hydrogen, react with acid to produce metal salt and hydrogen.

- Hydrochloric acid + metal → metal chloride salt + hydrogen
- Nitric acid + metal → Not needed
- Sulphuric acid + metal → metal sulphate + hydrogen

Only nitric acid can react with copper

Metal oxides:

- Hydrochloric acid + metal oxide → metal chloride salt + water
- Nitric acid + metal oxide → metal nitrate salt + water
- Sulphuric acid + metal oxide → metal sulphate + water

e.g.:

$ZnO (s) + HCl (aq) \rightarrow ZnCl_2 (s) + H_2O (l)$
$FeO + H_2SO_4 (aq) \rightarrow FeSO_4 (s) + H_2O (l)$
$MgO + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O (l)$

Metal carbonates:

Acids react with metal carbonate to produce metal salt, carbon dioxide & water.

- Hydrochloric acid + metal carbonate → metal chloride salt + water + carbon dioxide
- Nitric acid + metal carbonate → metal nitrate salt + water + carbon dioxide
- Sulphuric acid + metal carbonate → metal sulphate + water + carbon dioxide

e.g.:

$CaCO_3 + HCl \rightarrow CaCl_2 + H_2O + CO_2$
$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$
$FeCO_3 + HNO_3 \rightarrow Fe(NO_3)_2 + H_2O + CO_2$

4.6 understand the general rules for predicting the solubility of salts in water:

- all common sodium, potassium and ammonium salts are soluble
- all nitrates are soluble
- common chlorides are soluble, except silver chloride , lead (II) chloride
- common sulfates are soluble, except those of barium, lead and calcium
- common carbonates are insoluble, except those of sodium, potassium and ammonium

4.7 describe experiments to prepare soluble salts from acids

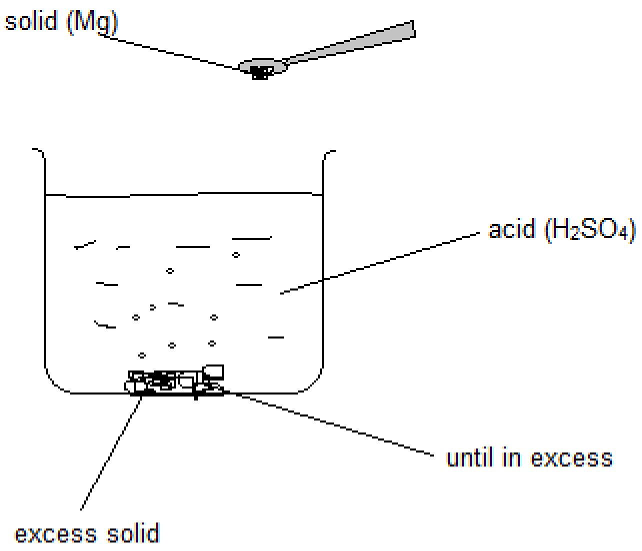
These all involve reacting a solid with an acid:

- acid + metal (but only for moderately reactive metals from magnesium to iron in reactivity series)
    - acid + metal oxide or hydroxide

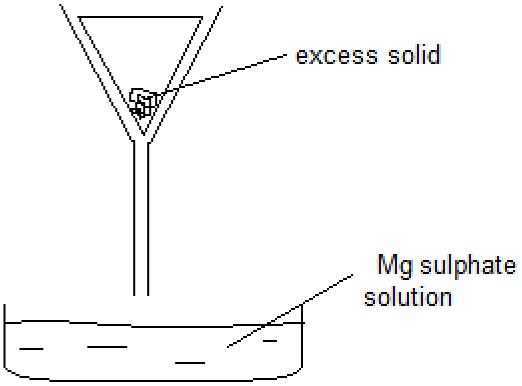
- acid + carbonate

Method 1: Making soluble crystals of magnesium sulphate:

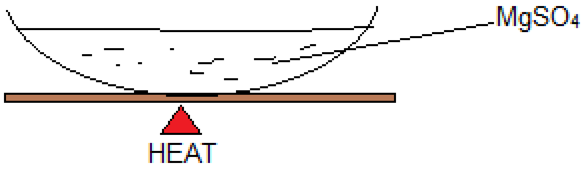
1. Add enough magnesium to dilute sulphuric acid, so that there is no acid left and the solution stops bubbling
   $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4 + H_2O$



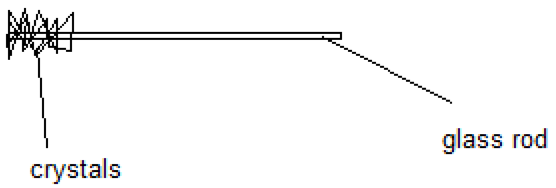
2. Filter the unused magnesium



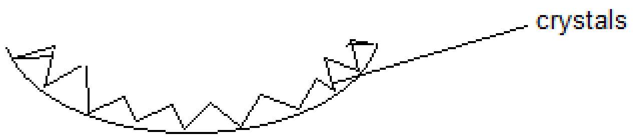
3. Heat the solution to evaporate water and to make the solution saturated.



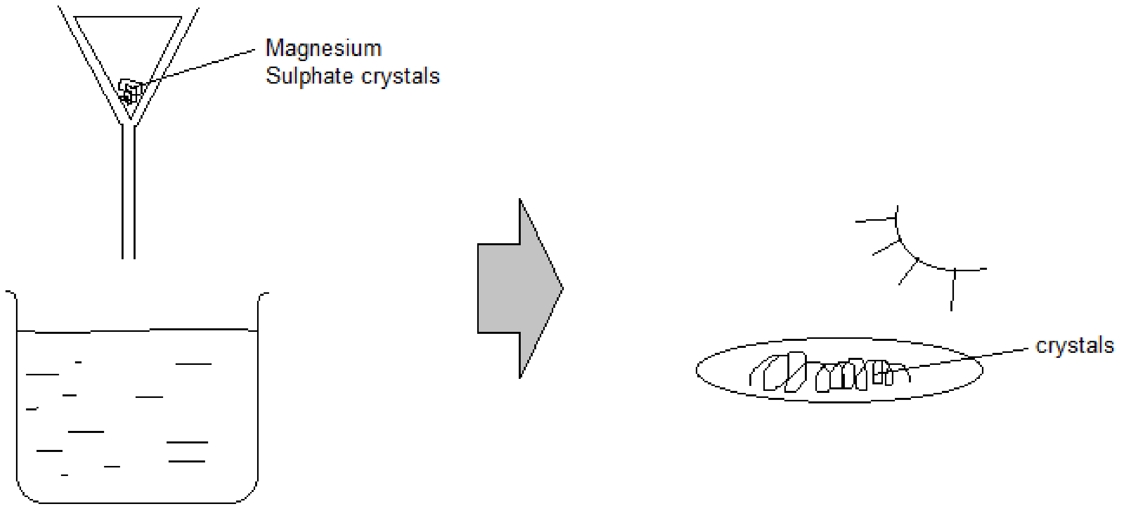
4. Test the solution whether it is saturated or not, using a glass rod.



5. Leave the solution to cool and crystallize.



6. Filter the crystals and dry them by keeping under the sunlight.

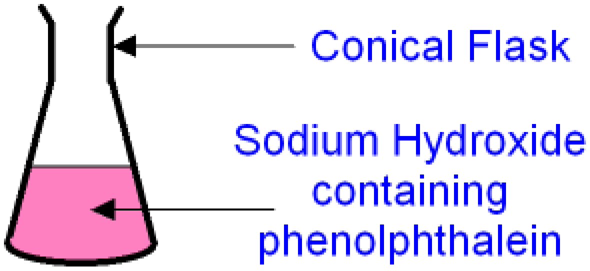


Method 2: Titration

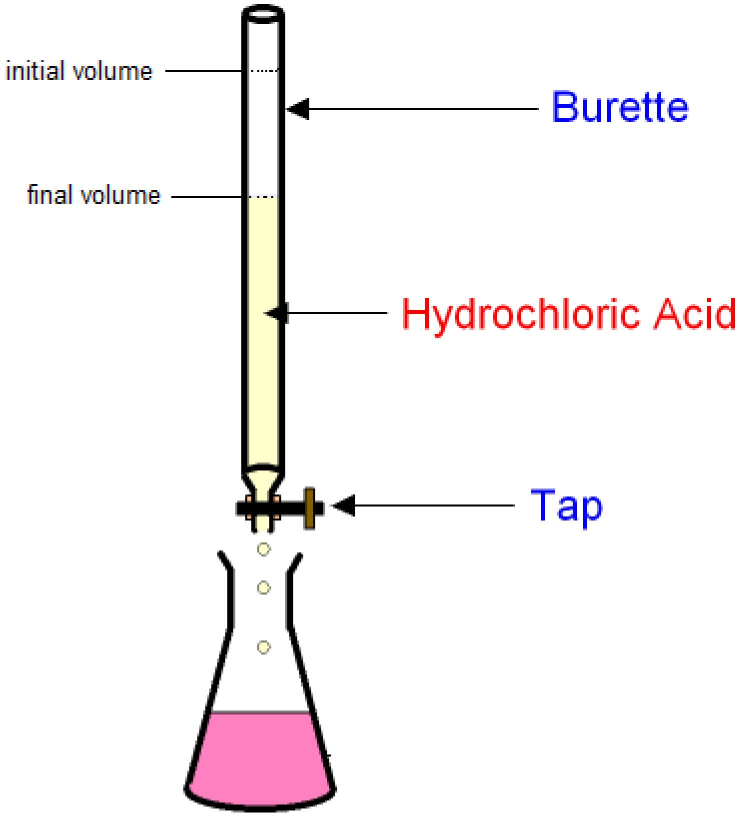
Acid + alkali → salt + water

$$HCl + NaOH \rightarrow NaCl + H_2O$$

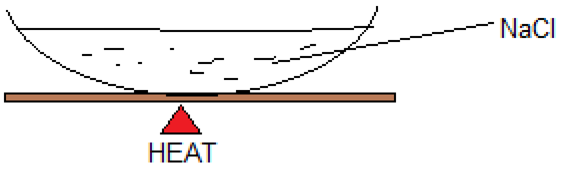
1. Take 25 cm^3 of alkali in a conical flask and add a few drops of phenalpthalein in it.



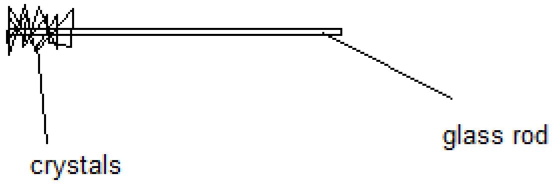
2. Run acid through burette until the colour change. Measure the amount of acid added by subtracting final volume from initial volume.



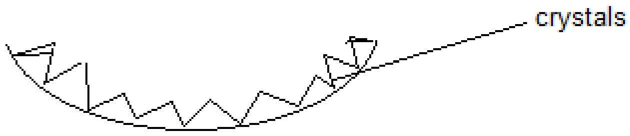
3. Repeat the experiment without using indicator, with same volume of acid and alkali.
4. Heat the solution to evaporate water and to make the solution saturated.



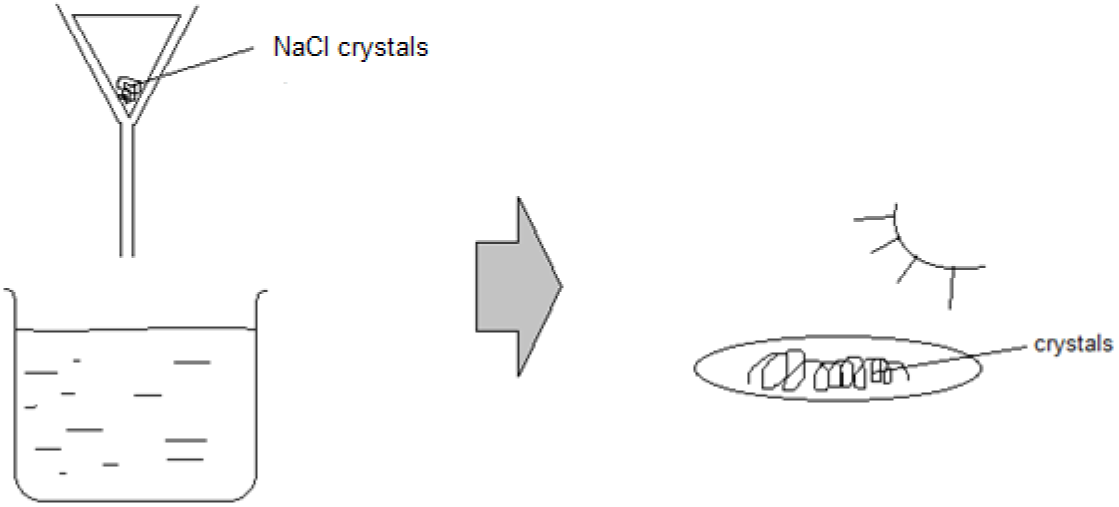
5. Test the solution whether it is saturated or not, using glass rod.



6. Leave the solution to cool and crystalise.



7. Filter the crystals and dry them by keeping them under the Sun.



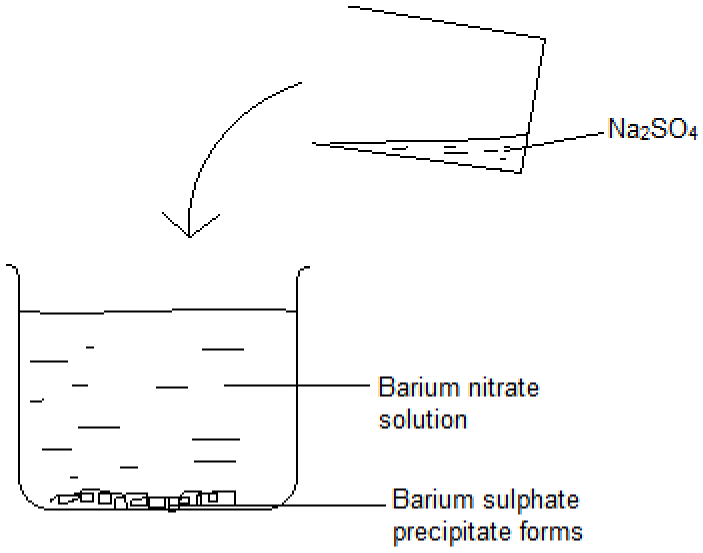
4.8 describe experiments to prepare insoluble salts using precipitation reactions

The insoluble salts are:

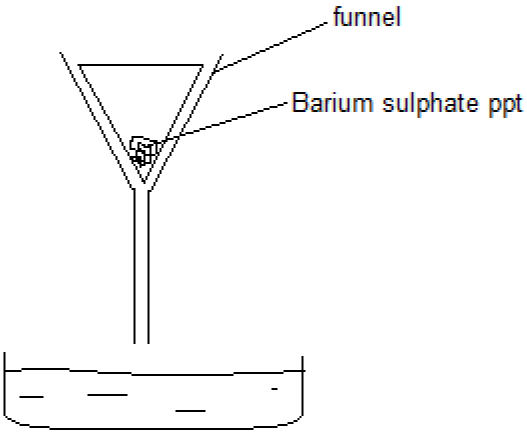
i. Barium sulphate
ii. Calcium sulphate
iii. Lead (II) sulphate
iv. Silver chloride
v. Silver bromide
vi. Silver iodide
vii. Lead (II) iodide
viii. Lead (II) chloride
ix. Carbonates except potassium, sodium and ammonia

Preparation of an insoluble salt:

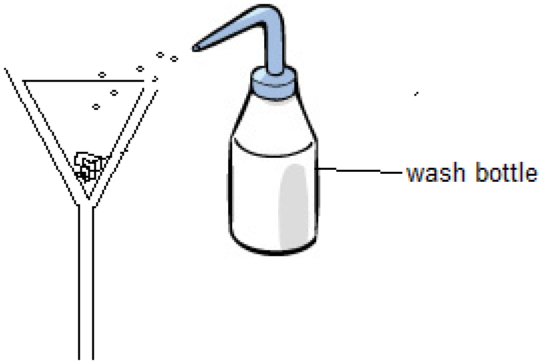
1. Two soluble salt solutions are mixed.
   Take barium nitrate in a beaker and add sodium sulphate solution in it until no precipitate forms.



2. Filter to collect the residue



3. Wash the residue with cold distilled water.



4. Leave residue to dry on filter paper/dry in a warm oven

Example: Describe how to prepare a dry solid sample of silver chloride, AgCl, a salt which is insoluble in water.

Sodium chloride + silver nitrate → silver chloride + sodium nitrate

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

Silver nitrate solution contains silver ions and nitrate ions in solution. The positive and negative ions are attracted to each other, but the attraction aren't strong enough to make them stick together. Similarly, sodium chloride solution contains sodium ions and chloride ions – again, the attractions, aren't strong enough for them to stick together.

When you mix the two solutions, the various ions meet each other. When silver ions meet chloride ions, the attractions are so strong that the ions clump together and form a solid. The sodium and nitrate ions remain in the solution.

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

Example: Making pure barium sulfate

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Barium chloride and dilute sulphuric acid are mixed together. Hydrogen ions and chloride ions are spectator ions and aren't involved in the reaction at all. Barium and sulphate ions attract together to form white precipitate of barium sulphate.

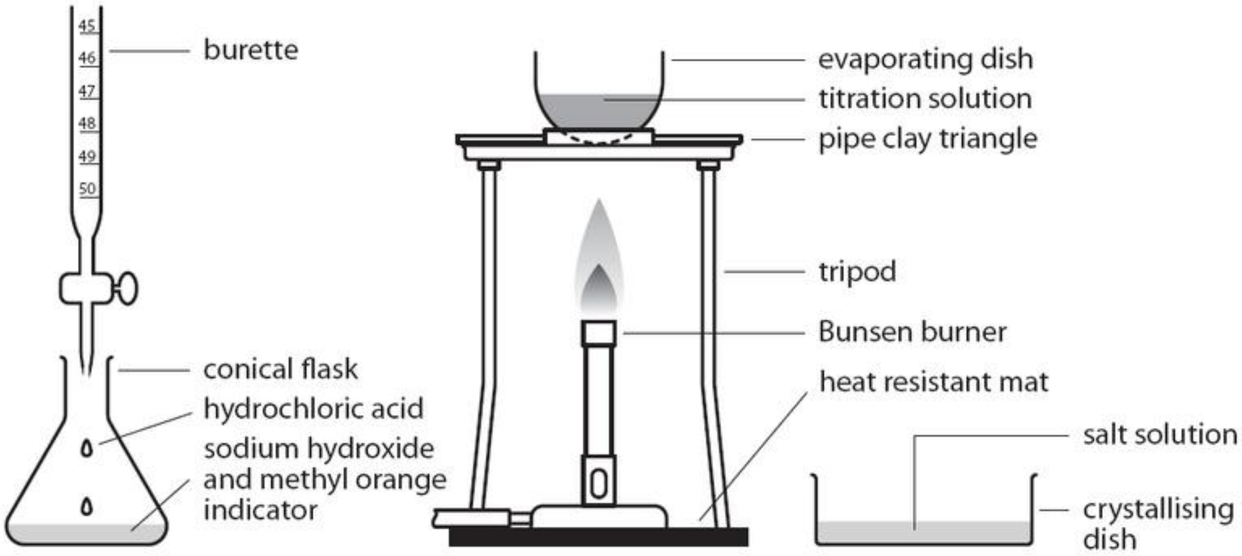
The mixture is filtered to get the precipitate. The solid barium sulphate is impure because of the presence of the spectator ions and any excess barium chloride solution or sulphuric acid. It is washed with pure water while it is still on the filter paper and then left to dry.

4.9 describe experiments to carry out acid-alkali titrations.

Titration is a quantitative experimental technique used in neutralisation reactions between acids and alkalis. It involves adding an acid from a burette into a conical flask containing the alkali. A pipette is used to measure out the exact volume of alkali and an indicator is added so the end point of the titration can be seen.

Procedure:

i. fill the acid up to the mark in the burette
ii. pipette 25.0cm^3 sodium hydroxide into a conical flask
iii. add a few drops of methyl orange indicator
iv. add acid from the burette drop wise with swirling of flask
v. stop when colour change is permanent (turns pink/red)
vi. note burette readings
vii. repeat until concordant results are obtained (results are within 0.1 of each other)
viii. take average of results



Uses:

If they ask you how to prepare a soluble salt using an acid and an alkali, titration must be used. You first carry out a normal titration, and find out the exact amount of acid needed to neutralise the alkali. Then you repeat it without an indicator so that the salt is not contaminated with its colour. You remove the salt from the neutralised solution by evaporation, then you dry it.

Calculations:

$$\left(\frac{MV}{n}\right) acid = \left(\frac{MV}{n}\right) alkali$$

Where,

M = molarity or concentration in mol/dm^3

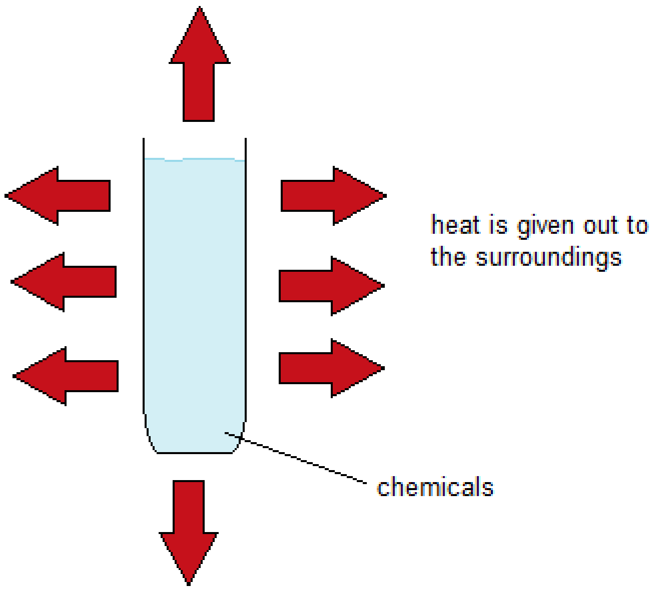
V = volume in cm^3

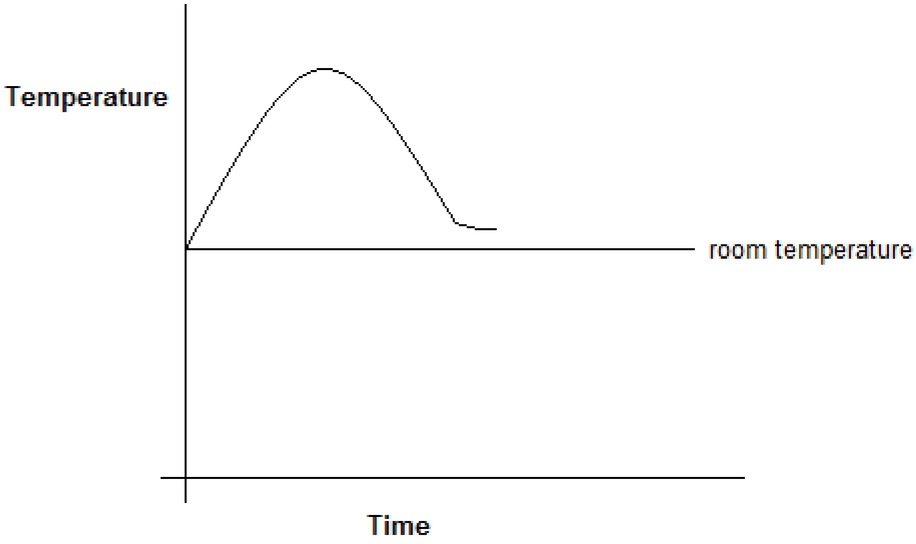
n = number of moles shown in the chemical equation

b) Energetics

4.10 understand that chemical reactions in which heat energy is given out are described as exothermic and those in which heat energy is taken in are endothermic

<u>Exothermic:</u> The chemical reactions which give out heat energy to the surrounding is called exothermic reaction.





Examples:

i. Combustion

$CH_3 + O_2 \rightarrow CO_2 + H_2O$

ii. Neutralization reaction

$HCl + NaOH \rightarrow NaCl + H_2O$

iii. Respiration

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + energy$

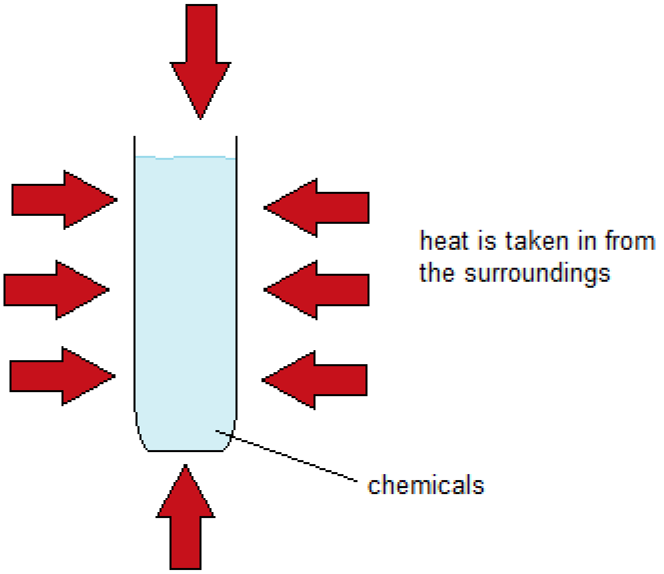
iv. Condensing

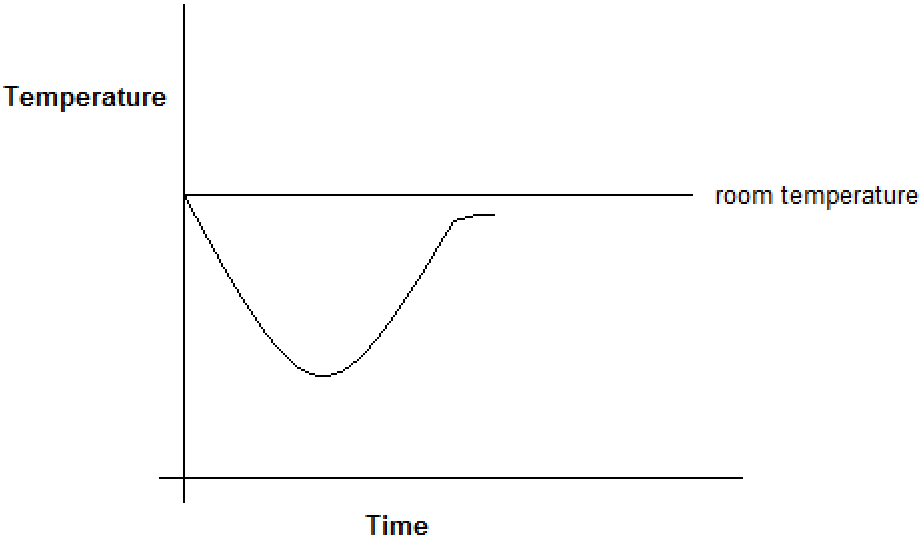
$H_2O\ (g) \rightarrow H_2O\ (l)$

v. Oxidation

$Mg + O_2 \rightarrow MgO$

<u>Endothermic:</u> The chemical reaction in which heat is taken in from the surroundings is called endothermic reaction.





Examples:

i. Thermal decomposition
   $CaCO_3 \xrightarrow[heat]{} CaO + CO_2$
ii. Photosynthesis
   $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$
iii. Vaporisation
   $H_2O\ (l) \rightarrow H_2O\ (g)$
iv. Dissolving ammonium nitrate

NH_4NO_3 (s) + (aq) → NH_4NO_3 (aq)

4.11 describe simple calorimetry experiments for reactions such as combustion, displacement, dissolving and neutralisation in which heat energy changes can be calculated from measured temperature changes

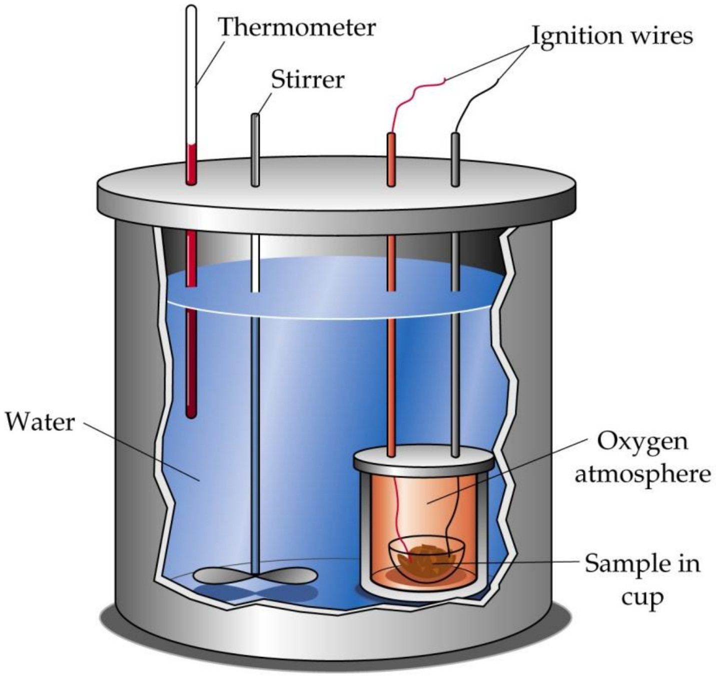


Fig: Calorimeter

Measuring energy changes involving solution (Dissolving, Displacement & Neutralization)

Measuring the heat evolved when magnesium reacts with a acid

When magnesium reacts with dilute sulfuric acid, the mixture gets very warm. The reaction is:

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

50 cm^3 of dilute sulfuric acid is run into a polystyrene cup using a pipette or burette and the temperature of the acid is measured. A small amount of magnesium powder is placed in a weighing bottle, and the mass of the bottle plus magnesium is recorded.

The magnesium is then tipped into the acid, and the maximum temperature reached is measured on the thermometer.

The mass of the empty weighing bottle is found and then the experiment is repeated to check the reliability of the result.

Volume of acid used = 50.0 cm^3

	Experiment 1	Experiment 2
Mass of weighing bottle + Mg(g)	10.810	10.800
Mass of weighing bottle afterwards (g)	10.687	10.685
Mass of Mg used (g)	0.123	0.115
Initial temperature (°C)	17.4	17.3
Maximum temperature (°C)	27.5	26.7

Temperature rise (°C)	10.1	9.4

If you do a reaction using a known mass of solution and measure the temperature rise, the amount of heat given out during the reaction is given by:

Heat given out = mass x specific heat x temperature rise

The specific heat is the amount of heat needed to raise the temperature of 1 gram of a substance by 1°C.

For water, the value is 4.18 J $g^{-1}C^{-1}$ (joules per gram per degree Celsius).

You can normally assume that dilute solutions have the same specific heat and density as water. You can also assume that negligibly small amounts of heat are used to warm up the cup and the thermometer.

In this case, we will take the mass of the solution as 50g. The mass of the magnesium is so small that it can be ignored. There are other major sources of error in the experiment which will make much more difference to the results.

Calculation:

(i) Heat evolved when 0.123 g Mg reacts = M x C x $\Delta\Theta$
= 50 x 4.18 x 10.1 J
= 2111J
= 2.111 kJ

Energy changes in reactions are always quoted in kJ. Diving by 1000 gives 2.111 kJ evolved when 0.123g of Mg react.

(ii) Now we need to calculate how much heat is evolved when 24.3 g of Mg react (Molar enthalpy change). 24.3 is an accurate value for the relative atomic mass of the magnesium.

If: 0.123 g Mg produce= 2.111 kJ

Then: 24.3g Mg produce $=\frac{24.3}{0.123}\ x\ 2.111\ kJ = 417kJ$

The amount of heat given out by the reaction is therefore:

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$ $\triangle H = -417\ kJ\ mol^{-1}$

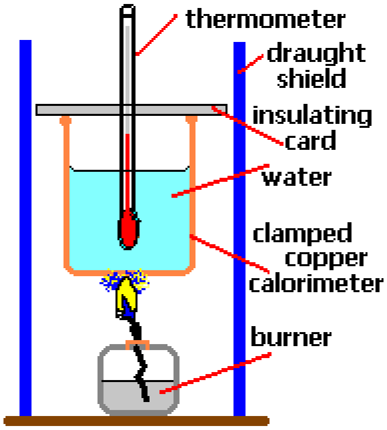
Or

Divide Heat change by the moles of Magnesium

$$mole\ of\ Mg = \frac{0.123g}{24.3g} = 0.005\ mol$$

$$molar\ enthalpy\ change = \frac{2.111}{0.005} = 422\ kJ/mol$$

Measuring the heat evolved in burning liquids (Combustion)



$$C_2H_5OH \rightarrow 2CO_2 + 3H_2O; \Delta H = ?$$

Volume of water = 100 cm^3
Mass of burner + ethanol before experiment = 37.355 g
Mass of burner + ethanol after experiment = 36.575 g
Original temperature of water = 21.5 ^oC
Final temperature of water = 62.8 ^oC
Mass of ethanol burnt = 0.780 g
Water temperature increase = 41.3 ^oC
Mass of water being heated = 100 g
Heat gained by water = 100 x 4.18 x 41.3 J
= 17260 J
= 17.26kJ

Burning 0.780g of ethanol produces 17.26 kJ

Amount of heat produced when 1g of ethanol burns = $\frac{17.26}{0.780} = 22.1\ kJ$

The amount of heat produced from 1 mole of ethanol = $\frac{17.26}{0.780}\ x\ 46\ kJ\ = 1020kJ$

The accepted value for ethanol is 1370 kJ of heat evolved when 1 mol ethanol is burnt. But the experimental value is very low. This is because large amount of heat losses for many source of error.

i. There is heat loss from the warming water to the surroundings.
ii. There is heat loss from the flame which goes straight into the air rather than into the water.
iii. Some heat are being used to raise the temperature of the flask and the thermometer.

4.12 calculate molar enthalpy change from heat energy change

Follow 4.11

4.13 understand the use of ΔH to represent enthalpy change for exothermic and endothermic reactions

ΔH is the symbol that represents the amount of energy lost or gained in a reaction.

+ΔH is endothermic (because it gains heat)
-ΔH is exothermic (because it loses heat)

The unit of ΔH is kJ/mol

4.14 represent exothermic and endothermic reactions on a simple energy level diagram

Endothermic reaction

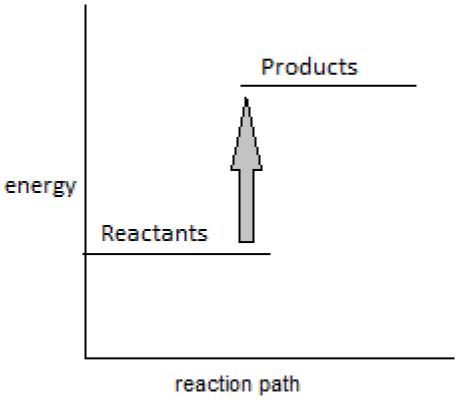


Fig: Energy Level Diagram

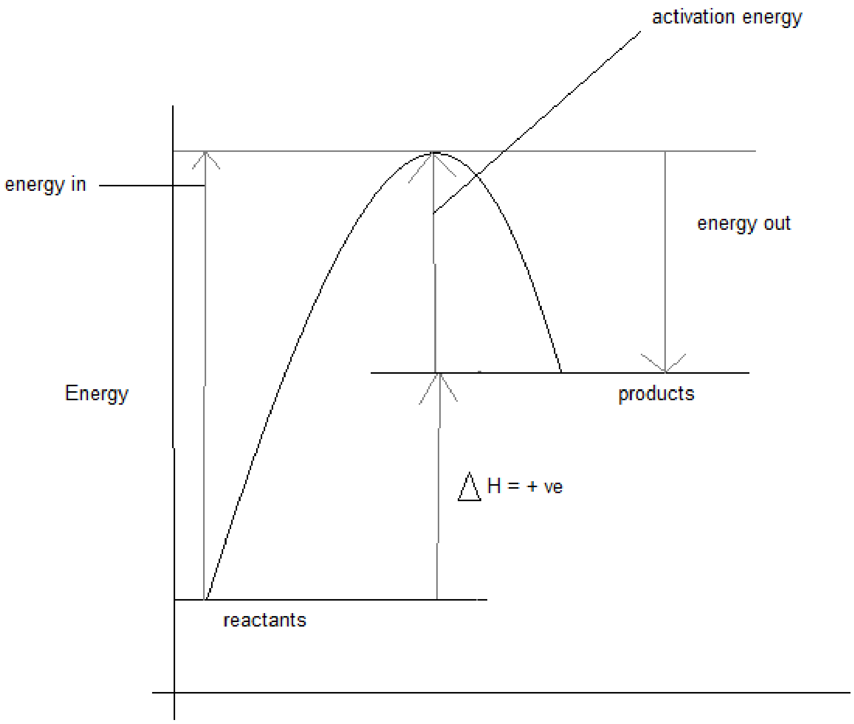


Fig: Energy profile diagram

Exothermic reaction

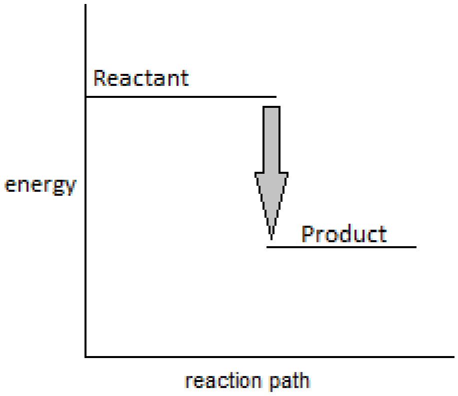


Fig: Energy Level Diagram

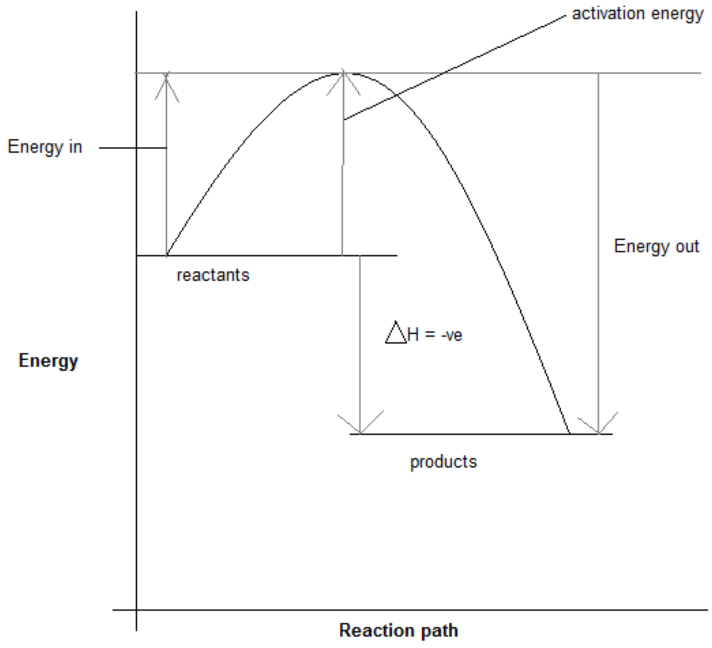


Fig: Energy profile diagram

4.15 understand that the breaking of bonds is endothermic and that the making of bonds is exothermic

During chemical reactions, bonds in the reactants have to be broken, and new ones formed to make the products. Breaking bonds needs energy and energy is released when new bonds are made. So for a chemical reaction, if the energy of breaking bond is greater than the energy of making bond, then the reaction is endothermic.

Endothermic reaction = Bond breaking energy **>** Bond making energy

If the energy required to break bond is less than making bond, then the reaction is exothermic.

Exothermic reaction = Bond making energy **>** Bond breaking energy.

ΔH will be energy taken in to break bond (E_{in}) – Energy given out to make bonds (E_{out})

$$\Delta H = E_{in} - E_{out}$$

4.16 use average bond energies to calculate the enthalpy change during a simple chemical reaction.

Bond Energies

Bonds	H – H	C – C	C – H	C – O	O = O	O = H	H – Cl	N ☰ N	N = H	Cl - Cl
Energies/mol	= 436	= 348	= 413	= 443	=466	= 463	= 431	= 946	= 488	=242

$H_2 + O_2 \rightarrow 2H_2O$

Bond Breaking:

When the bonds in the hydrogen molecule and oxygen molecule are breaking, they have 2 – free hydrogen atom. This requires energy to be absorbed. That means the process is endothermic.

Energy required to break 1 mol of H-H bonds = +436kJ

$\triangle H = 436$ kJ

Energy required to break 2 mol of H-H bonds = 2 x 436 kJ = +872kJ

$\triangle H = +872$ kJ

Energy required to break 1 mol of O=O bonds = +496 kJ

Total energy = 872 + 496 = +1368kJ

Bond making:

Energy released on forming 2 mol of O-H bonds = -468kJ x 2 = -926 kJ

Therefore, energy released on forming 2 mol of H_2O molecules = 2 x -926 = -1852 kJ

The overall enthalpy change for the reaction = (+1368) + (-1852) = -484 kJ

c) Rates of reaction

4.17 describe experiments to investigate the effects of changes in surface area of a solid, concentration of solutions, temperature and the use of a catalyst on the rate of a reaction

Rate of reaction means the amount of reactant used up or the amount of product obtained per unit time.

$$Rate\ of\ reaction = \frac{Amount\ of\ reactant\ used\ up}{time\ taken}$$

Or

$$Rate\ of\ reaction = \frac{Amount\ of\ product\ obtained}{time\ taken}$$

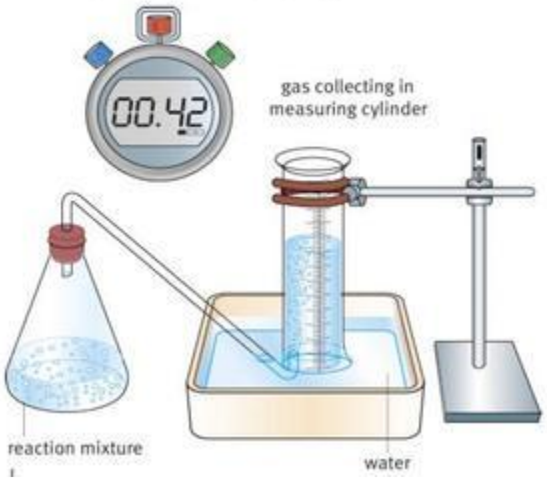
Or

$$Rate\ of\ reaction = \frac{Volume\ of\ gas\ produced}{time\ taken}$$

Reaction rate can be measured in four ways:

i. Measuring a gas product:

Collecting and measuring a gas product



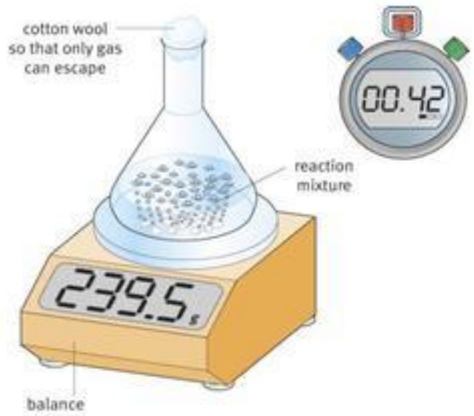
Obviously, this method is only suitable when one of the products formed is a gas. Ideally, this gas needs to be insoluble, otherwise some may dissolve in solution. In practice, this is not always easily achieve.

The gas formed is collected at regular intervals, such as every 30 or 60 seconds. Commonly, the gas is collected in a measuring cylinder, but a gas syringe could also be used.

The volume of gas collected is plotted against time taken; a line of best fit is then drawn, completing the graph. The rate of reaction is calculated by taking the initial rate from the graph.

ii. Measuring loss of mass:

Measuring the loss of mass as a gas forms

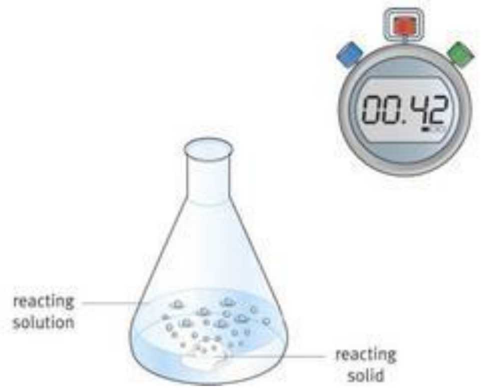


Quite similar to measuring the volume of a gas product, this particular method relies on the accurate measurement of the mass of gas lost.

Again, the mass lost is recorded at regular intervals, allowing a graph to be drawn with a curve of best fit. The initial rate of reaction can be calculated using this curve.

iii. Timing a reactant disappearing:

Timing how long it takes for a small amount of solid reactant to disappear

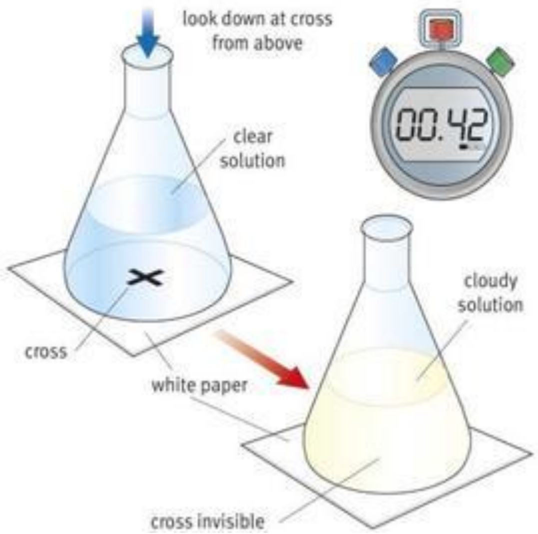


This method relies on human perception; by that, it is meant that it requires the chemist to judge when the reaction is complete. Hence, this method is open to human error.

Mixing the solid and solution reactants together and timing until you can no longer see any solid allows the chemist to calculate the average rate of reaction.

iv. <u>Timing a cloudy solution</u>:

Timing how long it takes for a solution to turn cloudy



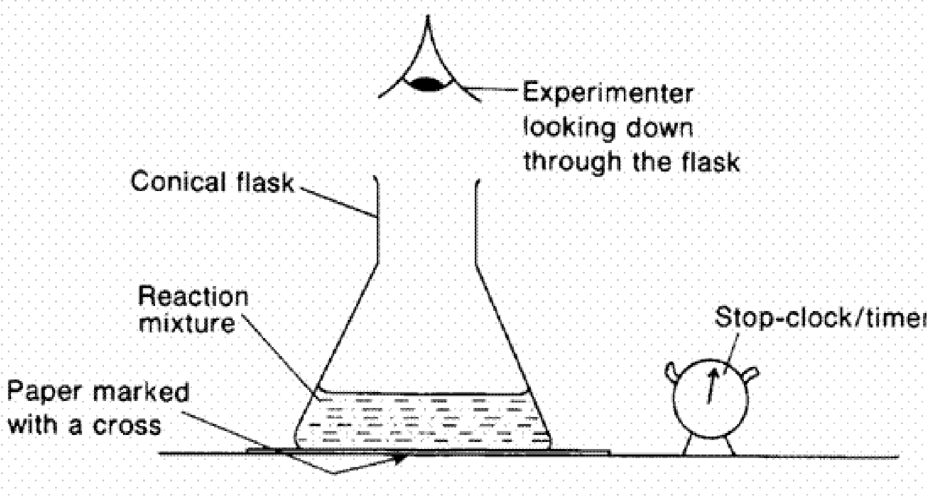
This is for reactions that produce an insoluble solid, such as sulphur.

Initially, you are required to draw a cross on a piece of paper, which is then placed under your glassware. Mix the solutions in the flask and start timing; you stop once you can no longer see the cross on the paper through the solution.

Again, this method relies on the individual. Each person may perceive the reaction completed at different timings.

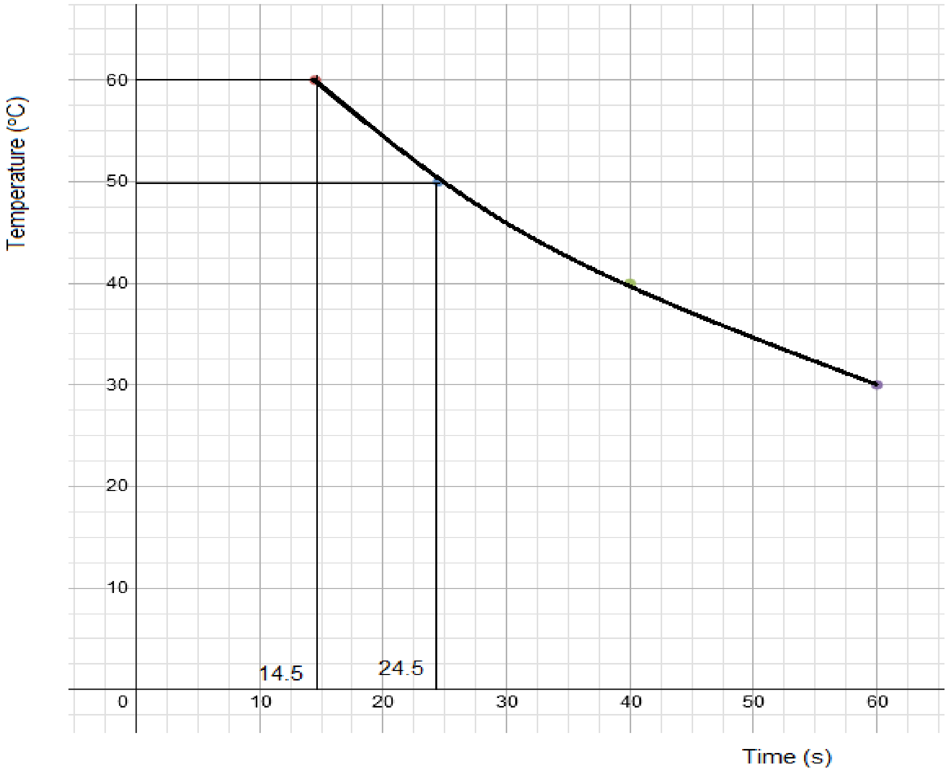
Experiment: How temperature affects the speed of Reaction

Sodium thiosulphate reacts with hydrochloric acid to form yellow deposit of sulfur.



$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + SO_2(g) + H_2O(l) + S(s)$

1. Measure out 10 cm^3 of the sodium thiosulphate solution (40 g dm^{-3}) into a conical flask and add 40 cm^3 of water.
2. If necessary, very gently warm the solution to a temperature of approximately 30 °C.
3. Add 5 cm^3 of 2 mol dm^{-3} hydrochloric acid and start the stop clock.
4. Swirl the mixture and place over the cross marked on the white sheet of paper.
5. Look down vertically onto the cross, and time its 'disappearance'
6. Repeat the experiment at about 40°C, 50°C, 60°C, 70°C by heating the thiosulphate solution before you add the acid. It is advisable to heat the solution to a little above these temperatures, to allow for the temperature to fall before you actually add the acid and start the clock.
    7. Plot graph – Temperature against time taken for the cross to disappear at different temperature.

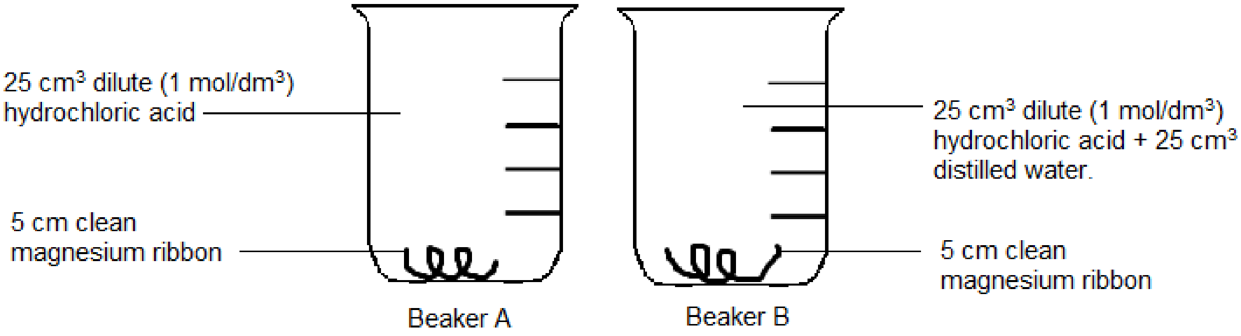


Observation:

i. The higher the temperature, the shorter the time taken for the cross to disappear from view. This means that the higher the temperature, the faster the rate of reaction.
ii. Observe the gradient of the graph at various temperature. The rate of reaction increases rapidly as the temperature increases.

Experiment: To study the effect of concentration on the speed of the reaction between magnesium and hydrochloric acid

1. Two beakers are taken: Beaker A: 50 cm3 dilute (1 mol/dm3) hydrochloric acid and 5cm clean magnesium ribbon. Beaker B: 25 cm3 dilute (1 mol/dm3) hydrochloric acid + 25 cm3 distilled water and 5cm clean magnesium ribbon.



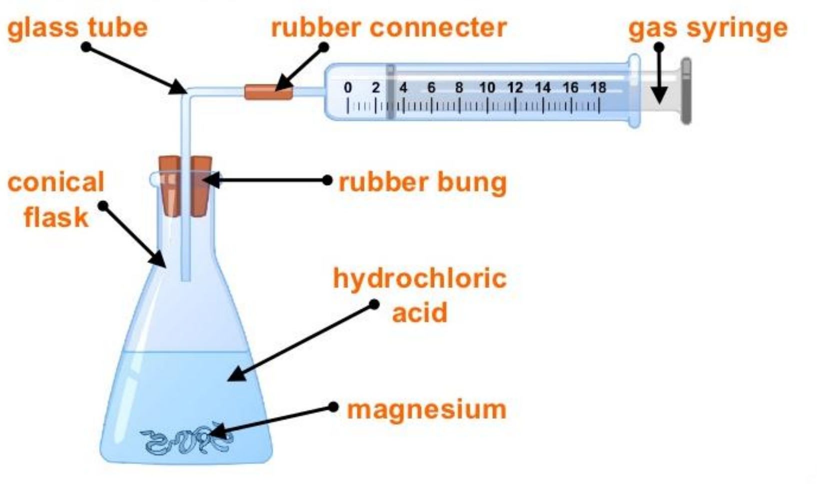
2. The time taken for each piece of magnesium ribbon to dissolve is recorded.

Beaker	A	B

Time taken for magnesium to dissolve (g)	39	78

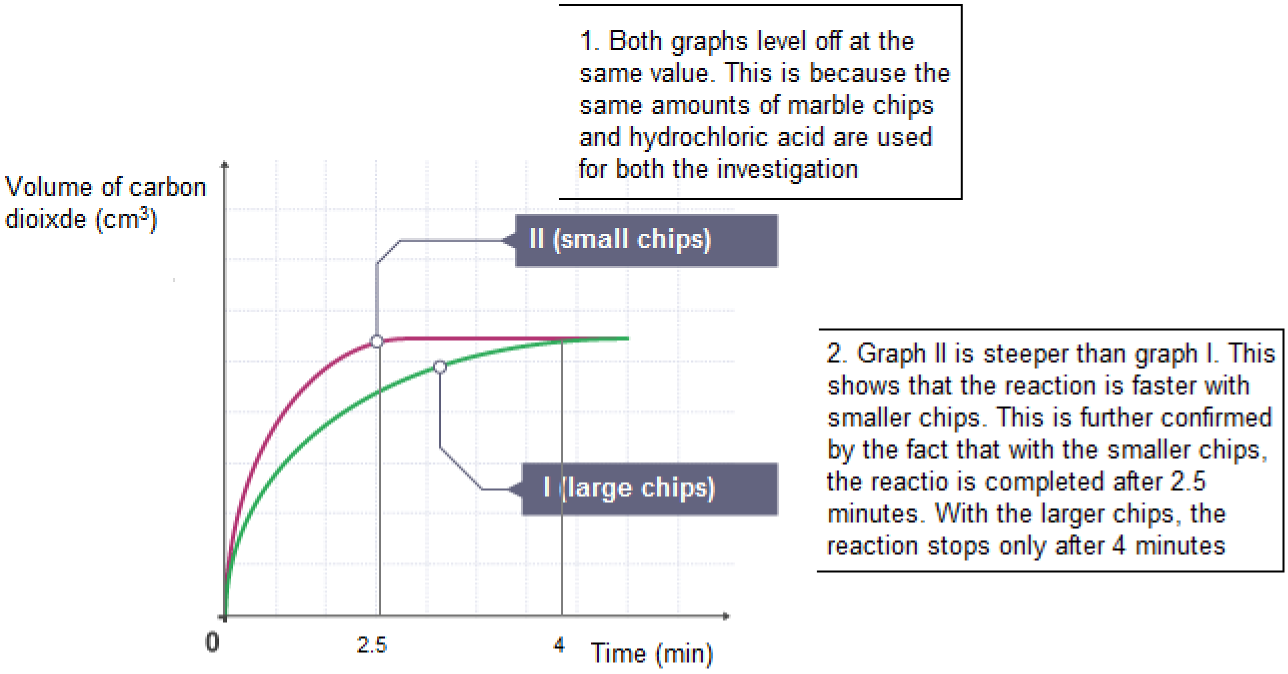
Observation: In the above experiment, the acid in beaker A is twice as concentrated as the acid in beaker B. the time taken for magnesium to react completely in beaker A is shorter. Thus, we can conclude that a reaction proceeds faster when the concentration of a reactant is increased.

Experiment: To study the effect of particle size on the speed of reaction.



1. A conical flask is set up and attached to a gas syringe(to measure the amount of CO2 produced). Marbles chips are placed in hydrochloric acid in the flask. The volume of gas produced is recorded at one-minute intervals for investigation 1.
2. The experiment is repeated for investigation II, with marble chips that have been crushed into much smaller pieces.

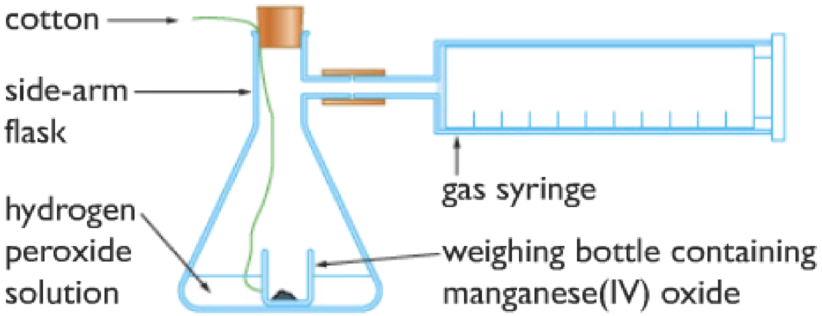
Observation:



1. Both graphs level off at the same value. This is because the same amounts of marble chips and hydrochloric acid are used for both the investigation

2. Graph II is steeper than graph I. This shows that the reaction is faster with smaller chips. This is further confirmed by the fact that with the smaller chips, the reactio is completed after 2.5 minutes. With the larger chips, the reaction stops only after 4 minutes

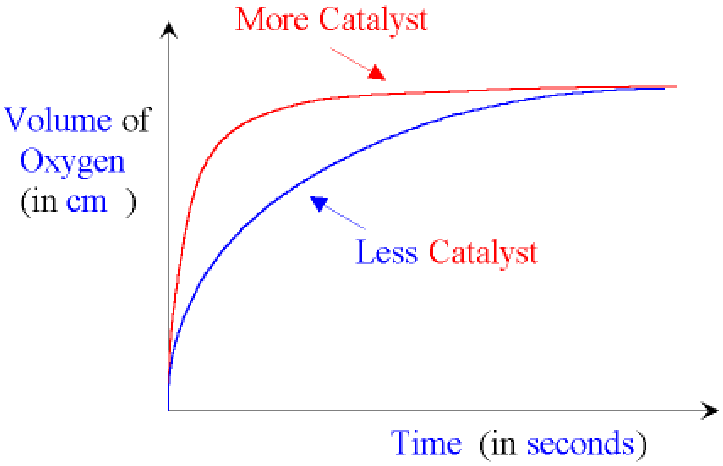
Experiment: Investigating the effect of catalyst in decomposition of hydrogen peroxide



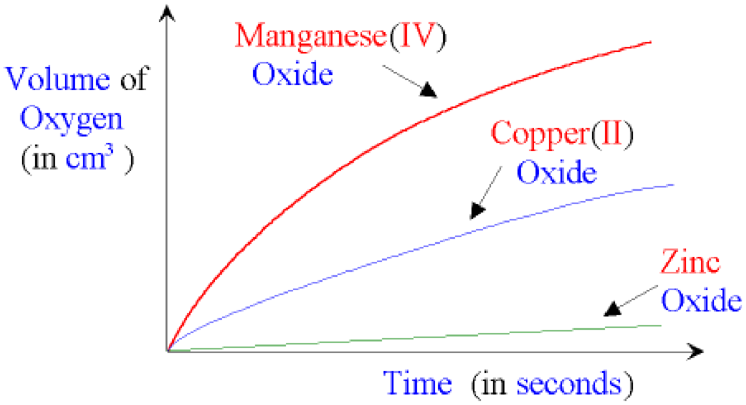
1. Hydrogen peroxide solution is placed in a beaker with a weighing bottle containing catalyst (manganese (IV) oxide).
2. The weighing bottle is dropped, and the reaction starts.
3. Measure the time, need to complete the reaction. (Usually when the volume of oxygen in gas syringe doesn't increase anymore).
4. Repeat the above experiment with no catalyst, with different mass of catalyst and different catalyst itself.

Observation:

1. Experiment using more mass of catalyst produce a steeper graph. This means adding more catalyst increases the rate of reaction.



2. With different catalyst, it is observed that only some catalyst produce higher rate of reaction than others. In this experiment, Manganese IV Oxide is proved to be the best catalyst for this reaction.



4.18 describe the effects of changes in surface area of a solid, concentration of solutions, pressure of gases, temperature and the use of a catalyst on the rate of a reaction

Temperature: The higher the temperature, the faster the rate of reaction.

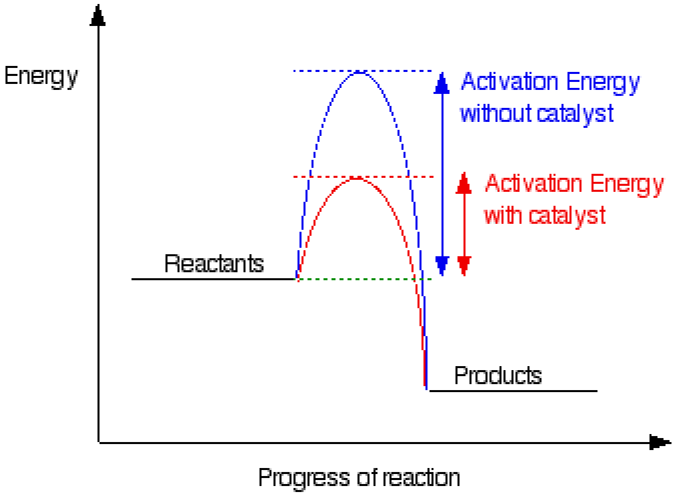
Concentration: The more concentrated the substance used, the faster the speed of a chemical reaction.

Pressure: The speed of reactions which involve gases is faster at higher pressures.

Particle Size: Smaller particles like powders have a much greater surface area than larger lumps or crystals. Within a greater surface area, the other reactant can attack it more easily and thereby increase the speed of reaction.

4.19 understand the term activation energy and represent it on a reaction profile

Activation energy is the minimum energy that molecules must possess during their collisions in order for a chemical reaction to occur. In other words, it is the total energy which is needed to break all the bonds in the reactant.



4.20 explain the effects of changes in surface area of a solid, concentration of solutions, pressure of gases and temperature on the rate of a reaction in terms of particle collision theory

Collision theory: It states that for a reaction to occur between two particles,

a) The reactant particles must collide with each other.
b) They must collide with a certain minimum energy which is called activation energy.

Successful collision: A collision that results in the formation of products is called successful or effective collision.

Factors that affect rate of reaction:

Temperature: When the temperature is raised, the reactant particles have a greater heat energy, causing them to move about more and with a greater kinetic energy. They, therefore, stand an increase of frequency of successful collision which results in higher rate of reaction.

Concentration: When concentrated substance is used, there is a greater likelihood that reacting molecules will collide with one another with sufficient energy to form products.

Pressure: When the pressure is increased, the particles get closer. So more particles are in per unit volume and this results more frequency of successful collision. And hence, the rate of reaction increases.

Surface area: When the substance is powdered then more number of particles are exposed to get collided. So the number of successful collision per unit time increases resulting the higher speed.

4.21 explain that a catalyst speeds up a reaction by providing an alternative pathway with lower activation energy.

A catalyst is a substance which increases the rate of chemical reaction, without itself being chemically changed at the end of the reaction.

Catalysts work by proving a more direct route from reactants to products. If we plot energy against time for a reaction, there is an energy barrier over which the reactants have to pass. The height of this barrier above the energy of the reactants is called the activation energy. What catalysts do, is to lower this activation energy, so as to allow the reactants to change into products more quickly.

d) Equilibria

4.22 understand that some reactions are reversible and are indicated by the symbol ⇌ in equations

Reaction which can be reversed is called reversible reaction. It is indicated by the symbol ⇌.

E.g.:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

When nitrogen and hydrogen react, it produces ammonia. This is forward reaction.

$$N_2 + 3H_2 \rightarrow 2NH_3\text{(forward reaction)}$$

The produced ammonia can be decomposed to give nitrogen and hydrogen. This is backward reaction.

$$2NH_3 \rightarrow N_2 + 3H_2 \text{ (backward reaction)}$$

4.23 describe reversible reactions such as the dehydration of hydrated copper(II) sulfate and the effect of heat on ammonium chloride

Heating copper(II) sulphate crystals

If you heat blue copper(II) sulphate gently, the blue crystals turn to a white powder and water is driven off. Heating causes the crystals to lose their water of crystallisation and white anhydrous copper(II) sulphate is formed.

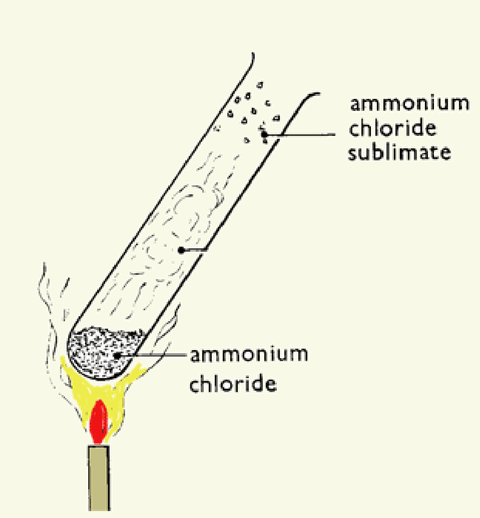
$$CuSO_4 \, . \, 5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$$

Now if you add water to the white solid, it will turn blue and will get warm.

$$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4 \, . \, 5H2O(s)$$

Heating ammonium chloride

If you heat ammonium chloride, the white crystals disappear from the bottom of the tube and reappear further up.



Heating ammonium chloride splits it into the colourless gases ammonia and hydrogen chloride.

$$NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$$

These gases recombine when the condition is changed from hot to cool.

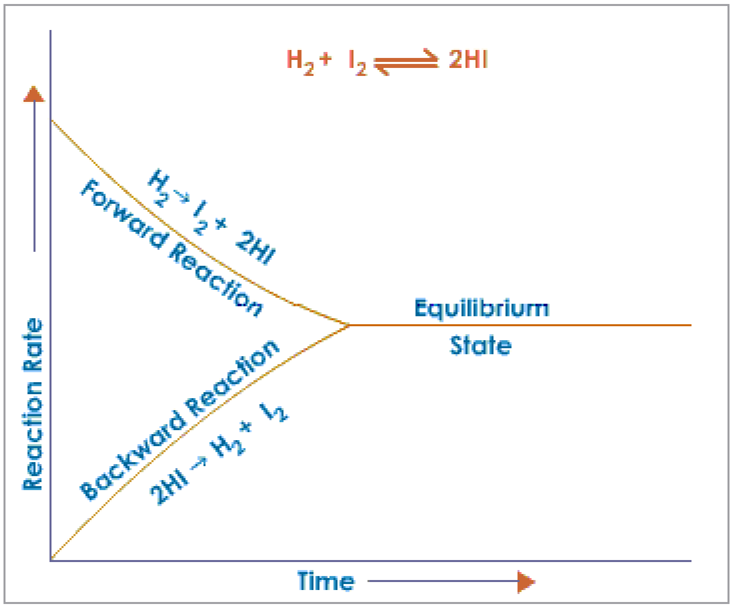
$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

4.24 understand the concept of dynamic equilibrium

When reversible reactions reach equilibrium the forward and reverse reactions are still happening but at the same rate, so the concentrations of reactants and products do not change.

Conditions in dynamic equilibrium:

i. Rate of forward reaction is equal to rate of backward reaction.
ii. Concentration of reactant and product remains same.
iii. Both forward and backward reactions are occurring.



The balance point can be affected by temperature, and also by pressure for gasses in equilibrium.

<u>Position of equilibrium</u>: Position of equilibrium means the concentration of reactant and products at the equilibrium state.

- When the position of equilibrium shifts to the left, it means the concentration of reactant increases.
- When the position of equilibrium shifts to right, this means the concentration of product increases.

4.25 predict the effects of changing the pressure and temperature on the equilibrium position in reversible reactions.

Le Chatelier's Principle

"When a reversible reaction is in equilibrium and you make a change, it will do what it can to oppose that change."

The effect of pressure:

Changing the pressure has little effect on an equilibrium mixture without gases - but can have a big effect on an equilibrium mixture containing gases. If the pressure is increased, the position of equilibrium moves in the direction of the fewest molecules.

Look again at the Haber process, which makes ammonia (NH3):

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

On the left, there are 1 + 3 = 4 molecules of gas. On the right, there are two molecules of gas. If the pressure is increased, the position of equilibrium will move to the right and more ammonia will be made. If the pressure is reduced, the position of equilibrium will move to the left and less ammonia will be made.

The effect of temperature:

- If the forward reaction is exothermic and the temperature is increased, the yield of products is decreased. If the temperature is decreased, the yield of products is increased.
- If the forward reaction is endothermic and the temperature is increased, the yield of products is increased. If the temperature is decreased, the yield of products is decreased.

The effect of concentration

When the concentration of a product is increased, the position of equilibrium will move to the left (as to remove the extra product).

$$BiCl_3\ (aq) + H_2O\ (l) \rightleftharpoons BiOCl\ (s) + 2HCl\ (aq)$$

If the concentration of hydrochloric acid in the equilibrium mixture is increased, the position of equilibrium moves to the left and the amount of white solid in that new equilibrium mixture is reduced.

The effect of catalyst

If we remove the products from an equilibrium mixture, more reactants are converted into products. If a catalyst is used, the reaction reaches equilibrium much sooner, because the catalyst speeds up the forward and reverse reactions by the same amount. The concentration of reactants and products is nevertheless the same at equilibrium as it would be without the catalyst.

Section 5: Chemistry in industry

a) Extraction and uses of metals

5.1 explain how the methods of extraction of the metals in this section are related to their positions in the reactivity series

Metals can be extracted in four different ways: (i) With Carbon, (ii) By electrolysis, (iii)By more reactive material& (iv) Heating

Metals that are below zinc in reactivity series can be extracted using carbon and carbon monoxide. Carbon is cheap and can also be as the source of heat.

Metals above zinc in reactivity series are usually extracted by electrolysis. Unfortunately, the large amounts of electricity involved make this an expensive process.

If a metal is more reactive than a metal in a compound, the reactive material will displace it and our targeted metal will be extracted.

Metals below Copper can be extracted from their compounds only by heating.

Q: Explain why more reactive metal are extracted by electrolysis.
A: Extraction process is a reduction process. The more reactive metal are, the harder they are to reduce. They need a powerful method of extraction which is electrolysis.

$$\overbrace{CuO + H_2 \text{ — } Cu}^{\text{reduction}} + H_2O$$

$$\underbrace{Fe_2O_3 + 3C \text{ — } 2Fe}_{\text{reduction}} + 3CO$$

$CaO + C$ — No reaction(carbon cannot displace oxygen from CaO)

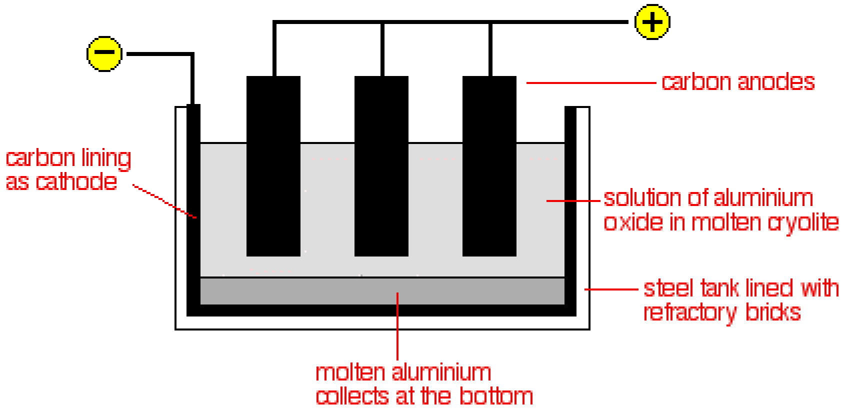
Q: Why iron can be extracted by carbon reduction process?
A: Carbon is more reactive than iron and that's why it can displace iron from their oxides.

5.2 describe and explain the extraction of aluminium from purified aluminium oxide by electrolysis, including:

the use of molten cryolite as a solvent and to decrease the required operating temperature

the need to replace the positive electrodes

the cost of the electricity as a major factor



Aluminium is extracted from aluminium oxide which comes from bauxite.Aluminium oxide has very high melting point, and it isn't practical to electrolyse molten aluminium oxide. So aluminium oxide is dissolved in molten cryolite to decrease working temperature.

The electrolyte is a solution of aluminium oxide in molten cryolite at a temperature of about 1000°C. The molten aluminium is siphoned off from time to time and fresh aluminium oxide is added to the cell. The cell operates at 5-6 volts and with current 100,000amps. The heat generated by the huge current keeps the electrolyte molten. The large amounts of electricity needed are a major expense.

Cathode:

Aluminium ions are attracted to the cathode and are reduced to aluminium by gaining electrons.

$Al^{3+}(l) + 3e^{-} \rightarrow Al(l)$

The molten aluminium produced sinks to the bottom of the cell.

Anode:

The oxide ions are attracted to the anode and lose electrons to from oxygen gas.

$2O^{2-}(l) \rightarrow O_2(g) + 4e^{-}$

This creates a problem. Because of the high temperatures, the carbon anodes burn in the oxygen to from carbon dioxide. The anodes have to be replaced regularly, and this also adds to the expense of the process.

 # The use of molten cryolite

The melting point of alumina is very high. If the alumina is dissolved in cryolite, it becomes 950°C. This reduces the expense for the heating purpose.

 # The need to replace the positive electrode

The oxide ions give up electrons to form oxygen gas.

$2O^{2-}(l) \rightarrow O_2(g) + 4e^{-}$

Oxygen gas attacks the carbon anode at high temperature to form gaseous products mainly carbon dioxide. This causes the carbon anodes to burn away, so they must be replaced regularly.

 # The cost of electricity as a major factor.

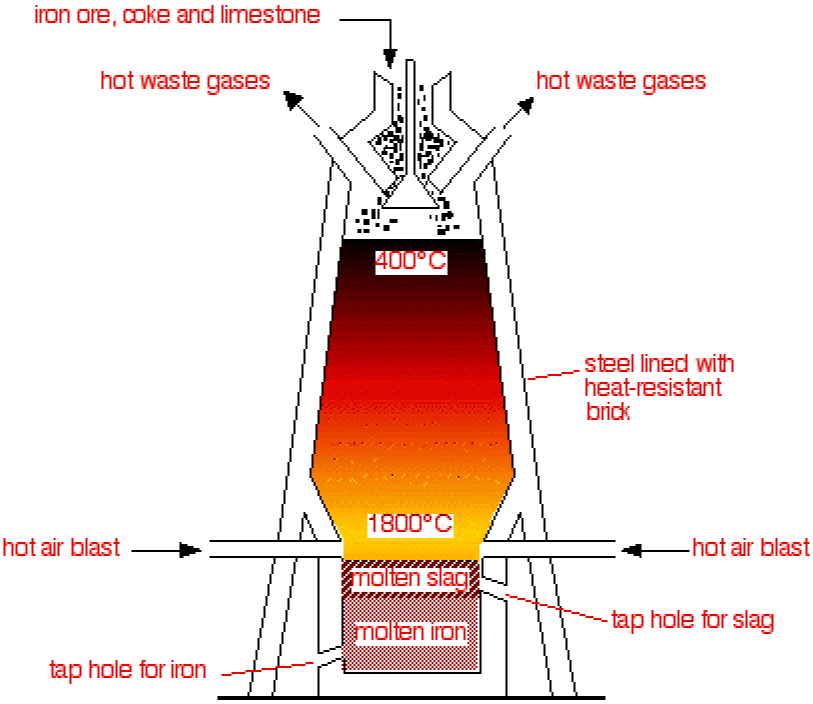
The cell operates about 5-6 volts but with current of up to about 1,00,000 amps. The heat generated by the huge current keeps the electrolyte molten. The large amount of electricity are major expense.

5.3 write ionic half-equations for the reactions at the electrodes in aluminium extraction

At the anode: $2O^{2-}(l) \rightarrow O_2(g) + 4e^-$

At the cathode: $Al^{3+}(l) + 3e^- \rightarrow Al(l)$

5.4 describe and explain the main reactions involved in the extraction of iron from iron ore (haematite), using coke, limestone and air in a blast furnace



Raw materials:

i. Iron ore
ii. Coke
iii. Limestone

Coke is used as the starting material. It is an impure carbon and it burns in how air blast to form carbon dioxide. This is strongly exothermic reaction.

$C(s) + O_2(g) \rightarrow CO_2(g)$

At the high temperatures in the furnace, the carbon dioxide is reduced by more carbon to give carbon monoxide.

$CO_2(g) + C(s) \rightarrow 2CO(g)$

Carbon monoxide is the reducing agent. Iron is gained from hematite (Fe_2O_3)

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$

The iron melts and flows to the bottom of the furnace, where it can be tapped off.

In the hotter parts of the furnace, some of the iron oxide is also reduced by carbon itself.

$Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(l) + 3CO(g)$

Limestone is added to the furnace to remove impurities in the ore. Limestone thermally decomposed to calcium oxide and carbon dioxide. It is an endothermic reaction.

$CaCO3(s) \rightarrow CaO(s) + CO_2(g)$

Silicon dioxide occurs naturally as quartz and it is a form of impurities that needs to be removed. Calcium oxide reacts with it to form calcium silicate. This melts and trickles to the bottom of the furnace as a molten slag, which floats on top of the molten iron and can be tapped off separately.

$CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$

5.5 explain the uses of aluminium and iron, in terms of their properties.

Uses of aluminium:

Pure aluminium isn't very strong, so aluminium alloys are used instead. As aluminium has low density, it resists corrosion and is strong. It has various uses. Like it is used in aero planes, pans etc. For its good conductivity of electricity, it is used as cables.

Uses of iron:

Types of iron	Iron mixed with	Properties	Some uses
Wrought iron	(pure iron)	Soft	Decorative work such as gates and railings
Mild steel	Up to 0.25% carbon	Hard	Nails, car bodies, ship building, girders
High-carbon steel	0.25-1.5% carbon	Very hard, sometimes brittle	Cutting tools, masonry nails
Cast iron	About 4% carbon	Hard but brittle	Manhole covers, guttering, engine blocks
Stainless steel	Chromium and nickel	Resistant to corrosion	Cutlery, cooking utensils, kitchen sinks

b) Crude oil

5.6 understand that crude oil is a mixture of hydrocarbons

Crude oil is a mixture of hydrocarbons – compounds containing carbon and hydrogen only. If any fraction from crude oil is burned, then carbon dioxide and water is produced – this shows that crude oil is a complex mixture of hydrocarbons.

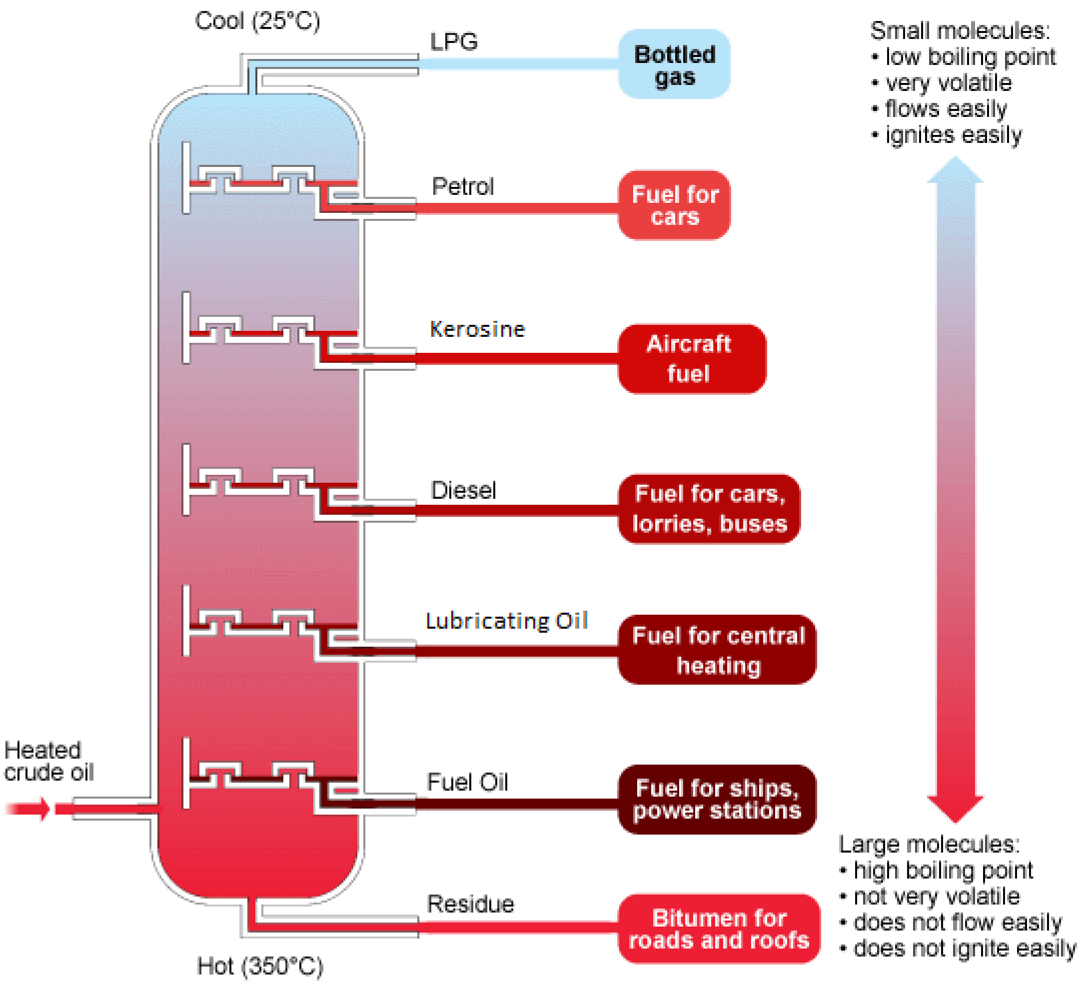
E.g: A fraction petrol, if it is burnt, it produces carbon dioxide and water.

$$C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$$

If crude oil is separated by fractional distillation, various fraction are produced. All the fraction are mixture of different kinds of hydrocarbon.

5.7 describe and explain how the industrial process of fractional distillation separates crude oil into fractions

The process of refining involves separating the hydrocarbons into fractions or batches using a technique called fractional distillation. Each fraction separates as they have different boiling points. The crude oil is heated in a furnace to around 400°C.



1. In the furnace at the bottom of the fractionating column, petroleum is heated into vapour.
2. The vapour is pumped into huge fractionating column. The column is cool at the top but hot at the bottom.
3. Hot vapour rises up the column and begins to cool down. The smaller hydrocarbons are collected at the top of the fractionating column as gases. The bigger hydrocarbons are collected at the lower sections of the fractionating column.

5.8 recall the names and uses of the main fractions obtained from crude oil: refinery gases, gasoline, kerosene, diesel, fuel oil and bitumen

Refinery gases: Refinery gases are a mixture of methane, ethane, propane and butane, which can be separated into individual gases if required. These gases are commonly used as LPG (liquefied petroleum gas) for domestic heating and cooking.

Gasoline (petrol): As with all other fractions, petrol is a mixture of hydrocarbons with similar boiling points. It is used in cars engine mainly.

Naptha: It is used as raw material in petrochemical industries.

Kerosine: Kerosine is used as fuel for jet aircraft, as domestic heating oil and as 'paraffin' for small heaters and lamps.

Diesel oil (gas oil):This is used for buses, lorries, some cars, and railway engines where the line hasn't been electrified. Some is also cracked to make other organic chemicals and produced more petrol.

Lubricating oil: It is used as oil for car engines and other machines.

Fuel oil: This is used for ships' boilers and for industrial heating.

Paraffin wax: Used to make candles, polish, wax paper, water proofing and grease.

Bitumen: Bitumen is a thick black material, which is melted and mixed with rock chippings to make the top surfaces of roads.

5.9 describe the trend in boiling point and viscosity of the main fractions

Viscosity: Viscosity refers to the less ease of flow of a liquid. If the number of carbon increases, the attraction between the hydrocarbon molecules also increases which results the liquid become more viscous with the increasing length of the chain of hydrocarbon. The liquid flow less easily with the increasing molecular mass. That means the hydrocarbons become more viscous.

Melting point/ Boiling point: As the molecules get larger, the intermolecular attraction becomes greater. So more heat is needed to separate the molecules which results increasing the molecular size, increases the boiling point.

Volatility: Volatility means the state of changing physical state. With the increasing molecular size, the hydrocarbon liquids become less volatile. As the attraction between the molecules increases with the increasing molecular size.

Flambility: The small hydrocarbons burn easily. So, the bigger the size, the less flammable the hydrocarbons are.

Summary:

As the molecules get bigger, the following changes occur:

- boiling point increases
- The liquids become less volatile.
- The liquid flow less easily (they become more viscous)
- they burn less easily

5.10 understand that incomplete combustion of fuels may produce carbon monoxide and explain that carbon monoxide is poisonous because it reduces the capacity of the blood to carry oxygen

If there isn't enough air (or oxygen), you get incomplete combustion. This leads to the formation of carbon or carbon monoxide instead of carbon dioxide.

$2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$

Carbon monoxide is colourless, odourless and is very poisonous. Carbon monoxide is poisonous because it combines with hemoglobin, preventing it from carrying oxygen. People can be made ill or even die, because of lack of oxygen in your body.

5.11 understand that, in car engines, the temperature reached is high enough to allow nitrogen and oxygen from air to react, forming nitrogen oxides

In car engines there is a high enough temperature to cause a reaction between oxygen and nitrogen in the air.

$N_2(g) + O_2(g) \rightarrow NO(g)$

5.12 understand that nitrogen oxides and sulfur dioxide are pollutant gases which contribute to acid rain, and describe the problems caused by acid rain

Acid rain is formed when acidic air pollutants such as sulphur dioxide and nitrogen dissolve in rainwater. Sulphur dioxide dissolves in water to form sulphurous acid (H_2SO_3).

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

In the presence of oxygen in the air, the acid is slowly oxidized to sulphuric acid (H_2SO_4).

Oxides of nitrogen also contribute to acid rain. In the presence of oxygen and water, nitrogen dioxide is converted to nitric acid.

$4NO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 4HNO_3(aq)$

Carbon dioxide in the air dissolves in rainwater to form carbonic acid, which is a weak acid.

$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

The pH value of normal rain is slightly below 7. The pH value of acid rain is approximately 3.4.

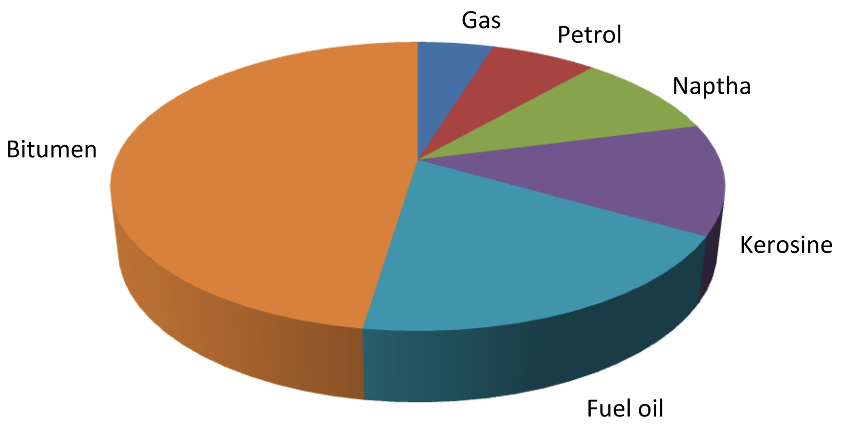
Problems of acid rain:

- Acid rain reacts with metals and with carbonates in marble and limestone. When this happens, metal bridges and stone buildings are damaged.
- Acid rain can reduce the pH value of natural water bodies from 6.5 and 8.5 to below 4. This will kill fish and other aquatic life.
- Acid rain also leaches important nutrients from the soil and destroys plants. Without these nutrients, plant growth is stunted. In some cases, acid rain dissolves aluminium hydroxide in the soil to produce aluminium ions, which are toxic to plants.

5.13 understand that fractional distillation of crude oil produces more long-chain hydrocarbons than can be used directly and fewer short-chain hydrocarbons than required and explain why this makes cracking necessary

The amounts of each fraction you get will depend on the proportions of the various hydrocarbons in the original crude oil, not in the amount in which they are needed.

If 100 ml of crude oil is distilled:

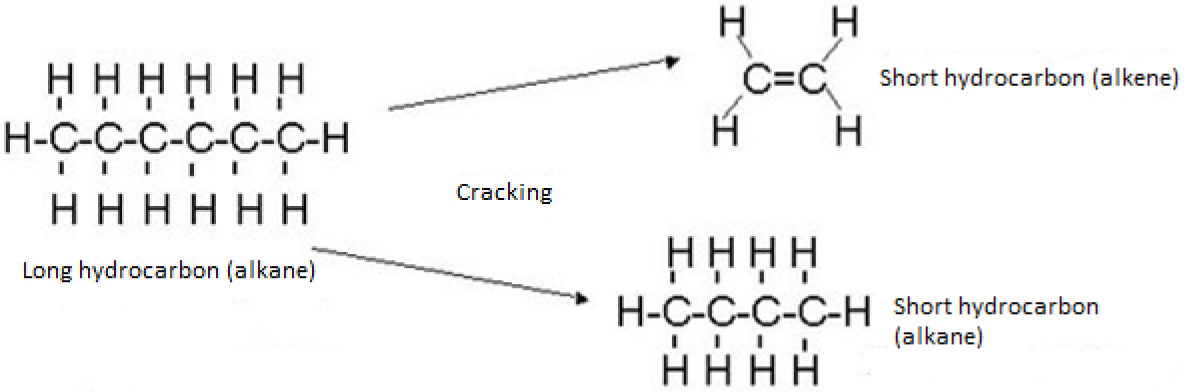


Long chain hydrocarbons which can't be used directly aren't very useful. With increasing chain length, the hydrocarbons become less flammable, more viscous and therefore less useful.

Short chain hydrocarbons burn well and flow well. Therefore, they are useful but these are produced less in the fractional distillation of crude oil. Thus chemists convert these large, less useful, heavy fractions into smaller, more useful ones by means of cracking. There are other necessities of cracking too.

- Cracking is used to produce petrol.
- Cracking is used to produce alkenes which are raw material in plastic industry.
- Cracking is used to produce hydrogen gas which is a raw material in manufacture of ammonia in Haber process.

5.14 describe how long-chain alkanes are converted to alkenes and shorter-chain alkanes by catalytic cracking, using silica or alumina as the catalyst and a temperature in the range of 600–700°C.



Cracking is simply splitting of larger molecules to simpler ones. The molecules are broken up in an random way which produce a mixture of alkanes and alkenes.

The gas oil fraction is heated to give a gas and then passed over a catalyst of mixed silicon dioxide and aluminium oxide at about 600 – 700°C. Cracking can also be carried out at higher temperature without a catalyst.

c) Synthetic polymers

5.15 understand that an addition polymer is formed by joining up many small molecules called monomers

Monomer: Monomers are small units which are joined to form polymer.

Polymer: When repeating monomers are joined together by polymerisation, them form a macromolecule called polymer.

There are two types of polymers:

i. Natural Polymers.
e.g.: starch, carbohydrate, fat, protein etc.
ii. Synthetic polymers (Man-made polymers).
e.g.: polythene, polypropene, polychloroethene (PVC), polystyrene, nylon, terylene.

Polymerisation: The process of joining together a large number of small molecules (monomers) to from a macromolecule/polymer is called polymerisation.

There are two basic types of reactions for forming polymers:

i. Addition polymerisation (polythene, polypropene, polychloroethene (PVC), polystyrene, nylon)
ii. Condensation polymerisation (nylon, terylene)

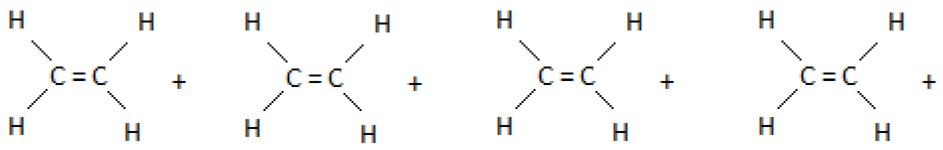
Addition polymerisation occurs when unsaturated monomer units join together without losing any molecules or atoms.

Condensations polymerisation occurs when two monomer join together alternately. And each time two monomers combine, a small molecule is lost. This is known as a condensation reaction.

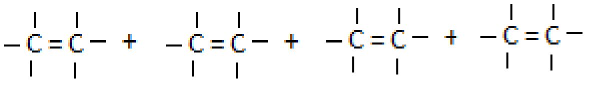
Polymerisation of Ethene to form Poly (ethene)

At high temperature, pressure, and in the presence of a catalyst, the carbon-carbon bonds of the alkene molecules (monomers) break. Each monomer forms single bonds with two other monomers. Eventually, they join to form a giant molecule.

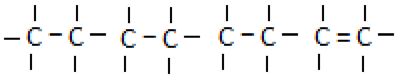
This is what happens during the addition polymerisation of ethene.



1. One bond in each double bond breaks.



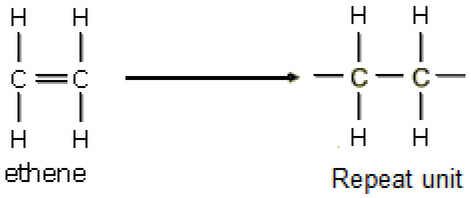
2. Each monomer forms single bonds with two other monomers.

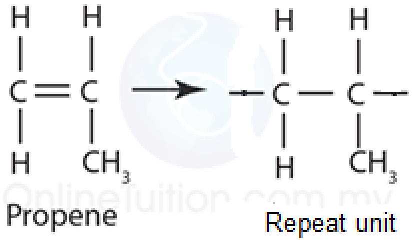


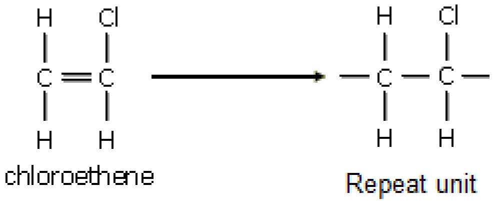
3. Poly (ethene) is formed.

The equation of the polymerisation of ethene is therefore written as

5.16 draw the repeat unit of addition polymers, including poly(ethene), poly(propene) and poly(chloroethene)

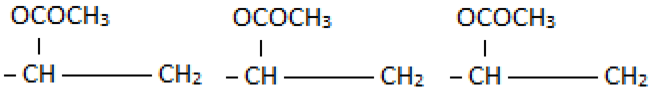






5.17 deduce the structure of a monomer from the repeat unit of an addition polymer

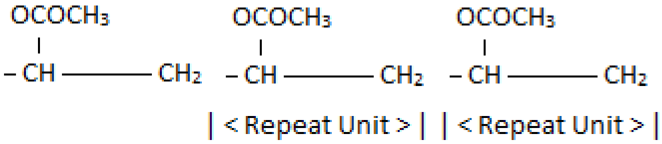
In the previous unit, we learnt how to draw a polymer given its monomer. How can we do the reverse? Lets examine this using PVA (polyvinyl acetate) as an example. PVA is used as the main constituent in emulsion paints. It has the formula.



$$-\overset{OCOCH_3}{\overset{|}{CH}}-CH_2-\overset{OCOCH_3}{\overset{|}{CH}}-CH_2-\overset{OCOCH_3}{\overset{|}{CH}}-CH_2$$

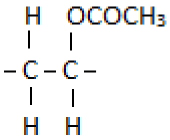
To deduce the structural formula of its monomer, we follow the steps below.

1. Identify the repeat unit in the polymer.



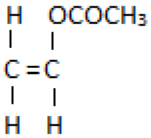
$$-\overset{OCOCH_3}{\overset{|}{CH}}-CH_2-\overset{OCOCH_3}{\overset{|}{CH}}-CH_2-\overset{OCOCH_3}{\overset{|}{CH}}-CH_2$$

2. Write down the formula of the repeat unit.



$$-\overset{H}{\underset{H}{C}}-\overset{OCOCH_3}{\underset{H}{C}}-$$

3. Convert the carbon-carbon single bond into a carbon-carbon double bond. The structural formula of the monomer, vinyl acetate, is



$$\overset{H}{\underset{H}{C}}=\overset{OCOCH_3}{\underset{H}{C}}$$

Vinyl acetate

5.18 describe some uses for polymers, including poly (ethene), poly(propene) and poly(chloroethene)

Poly (ethene):

Poly (ethene) comes in two types – low-density poly (ethene) (LDPE) and high-density poly (ethene) (HDPE). Low-density poly (ethene) is mainly used as a thin film to make polythene bags. It is very flexible and not very strong.

High-density poly (ethene) is used where rather greater strength and rigidity is needed – for example, to make plastic bottles such as milk bottles.

Poly (propene): Poly(propene) is somewhat stronger than poly(ethene). It is used to make ropes and crates (among many other things).

Poly (chloroethene): It is quite strong and rigid and so can be used for drainpipes, or replacement windows. It can also be made flexible by adding 'plasticisers'. That makes it useful for sheet floor coverings, and even clothing. These polymers don't conduct electricity and PVC is used for electrical insulation.

5.19 explain that addition polymers are hard to dispose of as their inertness means that they do not easily biodegrade

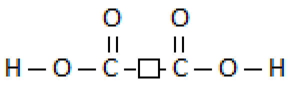
Addition polymers are unreactive. So they don't easily biodegrade.

5.20 understand that some polymers, such as nylon, form by a different process called condensation polymerisation

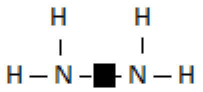
Some polymers are made by reacting two different types of monomers. Each of the monomers involved has a functional group at each end of the molecule. When these monomers react, a polymer is produced, and a small molecule such as water is also produced as a by-product of the reaction. This type of reaction is called condensation polymerisation.

Nylon – A Synthetic Polymide

Nylon was the first synthetic fibre made by condensation polymerisation. It is made from the monomers dicarboxylic acid and diamine. We can represent these two monomers units as follows:

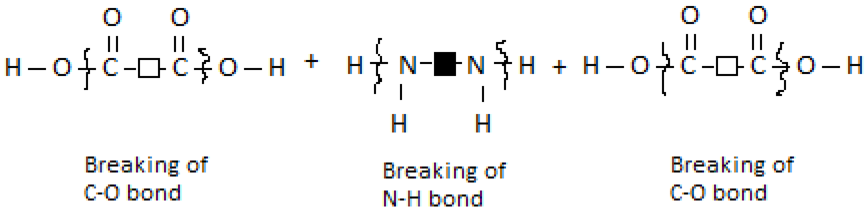


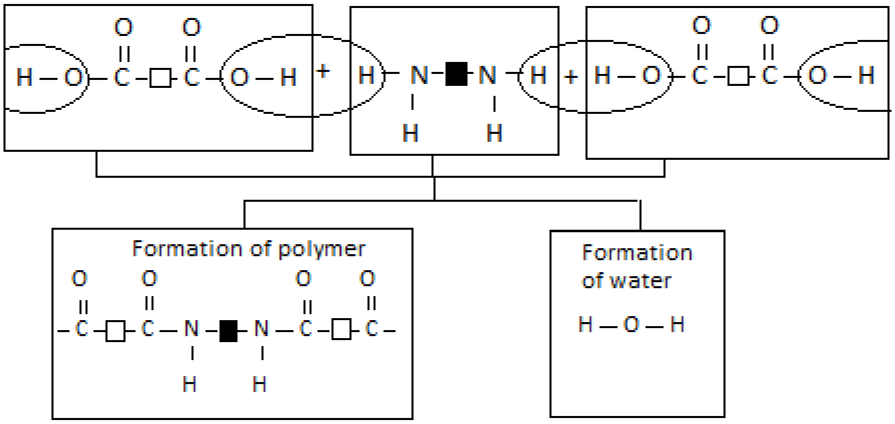
Dicarboxylic acid -
a molecule with two -COOH groups



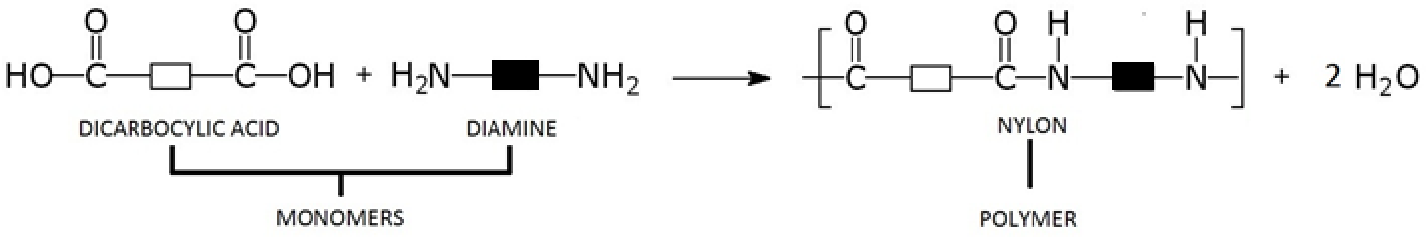
Diamine -
a molecule with two -NH_2 groups

This is how the two monomers react:





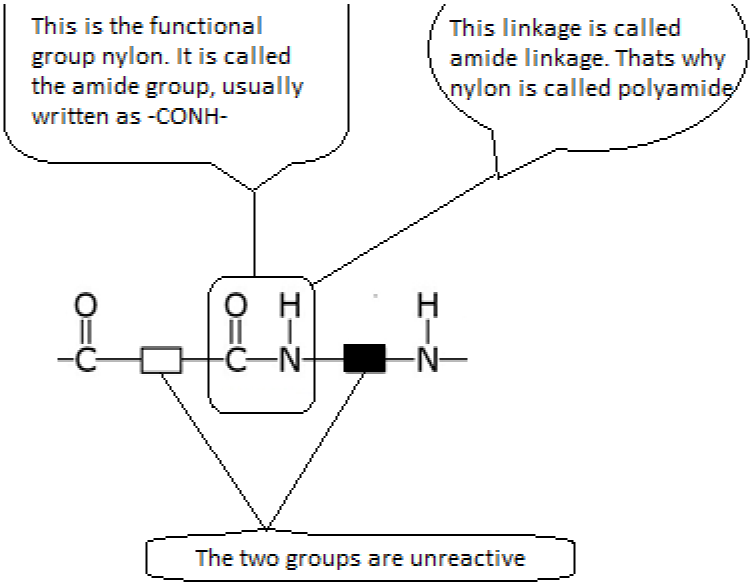
The reaction between a dicarboxylic acid and a diamine can be represented by the equation shown below.



Notice the following:

1. Each monomer for making nylon has two identical functional groups. The chain can be extended in two directions by the removal of water molecules.
2. A whole family of nylons can be formed using different dicarboxylic acid and diamine monomers.

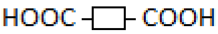
The reactive group in nylon is shown below.



Terylene – A Synthetic Polyester

Terylene is another example of a condensation polymer. It is included in a polyester group.

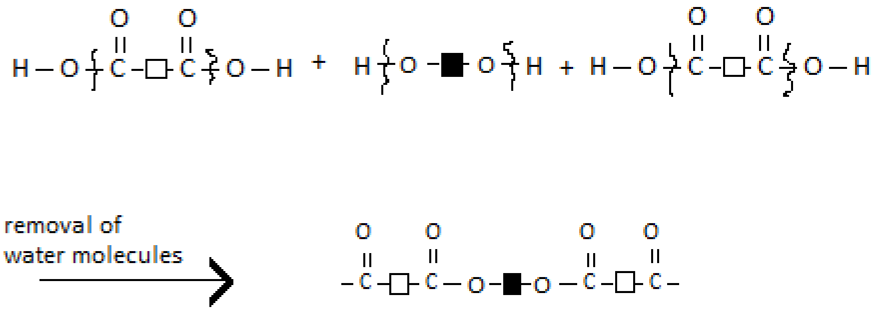
The monomers are:



Dicarboxylic acid

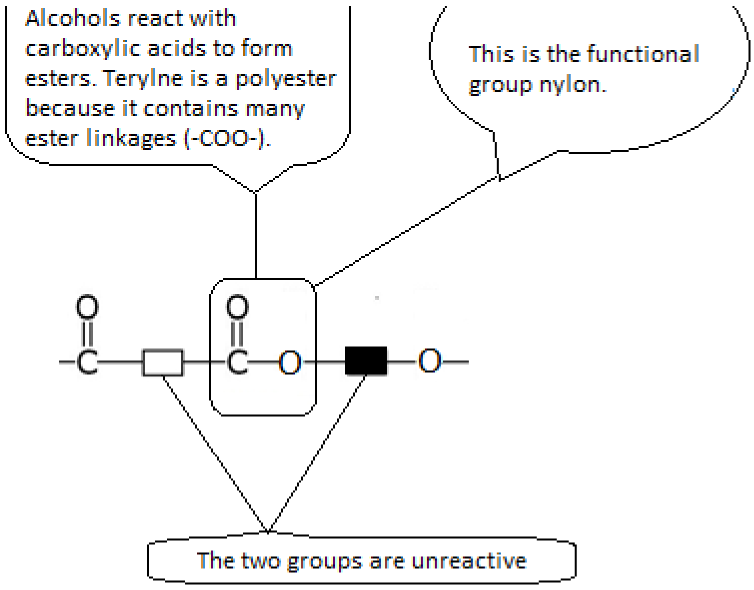
Diol

These monomers react to form Terylene and water:



Part of the Terylene polymer

The reactive group or functional group in Terylene is shown below.



5.21 understand that condensation polymerisation produces a small molecule, such as water, as well as the polymer.

(Follow 5.20)

 # Difference between condensation and addition polymerisation

Addition	Condensation
Unsaturated monomers are used to produce a polymer.	Each monomer have two functional group.
Elimination of water doesn't take place.	Elimination of water takes place.
Unsaturated monomers become saturated.	The monomers join among them by a linkage.
Addition polymerisation is a fast process.	Condensation polymerisation is comparatively a slower process.
Examples: polythene, poly(propene), polystyrene	Examples: nylon, polyester.

d) The industrial manufacture of chemicals

5.22 understand that nitrogen from air, and hydrogen from natural gas or the cracking of hydrocarbons, are used in the manufacture of ammonia

Ammonia (NH_3) is manufactured by using nitrogen from air and hydrogen from natural gas.

$N_2(g) + H_2(g) \rightarrow NH_3\,(g)$

5.23 describe the manufacture of ammonia by the Haber process, including the essential conditions:

i. a temperature of about 450°C
ii. a pressure of about 200 atmospheres
iii. an iron catalyst

Ammonia is manufactured by combining nitrogen and hydrogen in an important industrial process called the Haber process.

The Haber process:

- Raw materials: nitrogen (from the air)
  hydrogen (made from natural gas)
- The proportions: 1 volume of nitrogen to 3 volumes of hydrogen
- The temperature: $450^{o}C$
- The pressure: 200 atmospheres
- The catalyst: iron

The raw materials for the process are hydrogen and nitrogen. Hydrogen is produced industrially from cracking oil, and nitrogen from liquefaction of the air. The two gases are combined directly in a ratio of 3:1, and are passed over an iron catalyst at $450^{o}C$ and 200 atmospheres.

nitrogen + hydrogen ⇌ symbol ammonia

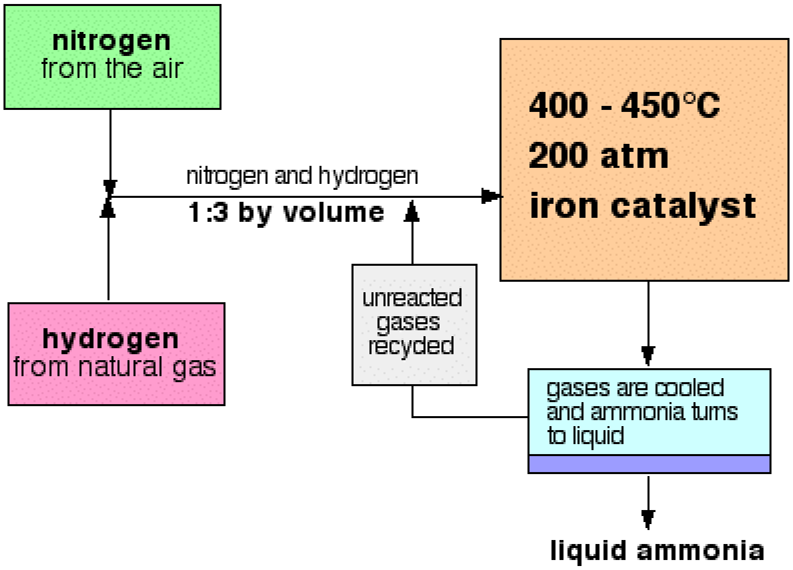
$N_2(g) + 3H_2(g) \rightleftharpoons$ *symbol* $2NH_3(g)$

As shown in the equation, the reaction is reversible with the forward reaction producing ammonia and the back reaction decomposing ammonia back into hydrogen and nitrogen gases. In order to favour the forward reaction, to produce more ammonia, high pressures are needed, as this will help to convert four volumes of hydrogen and nitrogen into two volumes of ammonia. The temperature of $450^{o}C$ is a compromise, because the reaction is also exothermic. This means it gives out heat and prefers a low temperature, but if the temperature drops too low, the rate becomes too slow. A moderate temperature of $450^{o}C$ is therefore used.

The catalyst in the Haber Process is the transition metal iron, with promoters of iron oxide. This speeds up both the forward and backward reactions, so the point of equilibrium between the two reactions is reached more quickly. When the ammonia has been produced, it is liquefied 'out', by reducing the temperature to $-34^{o}C$, and the unconverted nitrogen and hydrogen gases are recycled to avoid wastage.

5.24 understand how the cooling of the reaction mixture liquefies the ammonia produced and allows the unused hydrogen and nitrogen to be recirculated

The products from the reactant are sent through a cooling mechanism, this is at a temperature that condenses ammonia, but not hydrogen and nitrogen. Liquid ammonia is then collected but hydrogen and nitrogen float right back into the reactor.



Only 50% ammonia is produced in the converter. The reaction mixture is cooled, the ammonia condenses as liquid and as ammonia has lower boiling point than nitrogen and hydrogen which will be tapped off and will be collected in a storage tank.

5.25 describe the use of ammonia in the manufacture of nitric acid and fertilisers

Manufacture of Nitric acid:

Industrially, nitric acid is made by the catalytic oxidation of ammonia over heated platinum. Oxidising ammonia produces oxides of nitrogen which can then be dissolved in water to produce nitric acid.

Initially nitrogen(II) oxide will be formed from the catalytic oxidation of ammonia using the transition metal platinum at a temperature of 800°C to 1000°C.

$4NH_3\,(g) + 5O_2\,(g) \rightarrow 4NO(g) + 6H_2O\,(g)$

The nitrogen(II) oxide is rapidly cooled before combining with oxygen (from excess air) to form nitrogen(IV) oxide.

$2NO(g) + O_2(g) \rightleftharpoons 2NO_2\,(g)$

The nitrogen(IV) oxide is now allowed to react with water to form nitric acid.

$4NO_2\,(g) + O_2\,(g) + 2H_2O\,(l) \rightarrow 4HNO_3\,(aq)$

Most of the nitric acid made is used to make the all-important fertilisers such as ammonium nitrate.

Manufacture of Fertilizers:

The main use of ammonia is in the manufacture of fertilizers. Approximately 75% of all ammonia produced is converted into various ammonium compounds like ammonia sulphate ($(NH_4)_2SO_4$), ammonium nitrate (NH_4NO_3) and urea (NH_2CONH_2). These compounds are called nitrogenous fertilizers. They are solids for ease in handling and water soluble so that they seep into the soil to be absorbed by the roots of the plant.

$NH_3\,(aq) + HNO_3\,(aq) \rightarrow NH_4NO_3\,(aq)$

$2NH_3\ (aq) + H_2SO_4\ (aq) \rightarrow ((NH_4)_2SO_4)\ (aq)$

5.26 recall the raw materials used in the manufacture of sulfuric acid

- Sulphur (sulphur is found in rocks and some natural gasses)
- Oxygen from the air.

5.27 describe the manufacture of sulfuric acid by the contact process, including the essential conditions:

- **a temperature of about 450°C**
- **a pressure of about 2 atmospheres**
- **a vanadium(V) oxide catalyst**

I) First Sulphur dioxide is produce:

The raw materials are sulphur and air(oxygen). To produce sulphur dioxide, **either** burn sulfur in air:

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Or heat sulphide ores strongly in air:

$$4FeS_2(s) + 11O_2(g) \rightarrow 2Fe_2O_3(s) + 8SO_2(g)$$

II) Purification of Sulphur dioxide

The sulphur dioxide is then purified, by removing impurities like arsenic compounds which would otherwise poison the catalyst. It is then passed through an electrostatic dust precipitator, which, as its name implies, charges dust particles which are then removed by being attracted to oppositely charged plates.

III) Sulphur trioxide is produced:

Sulphur dioxide and air are then washed, dried and passed over a vanadium(V) oxide catalyst at 450^{o}C and 2-3 atmospheres. The reaction is reversible but at these temperature and pressures, 98% conversion to sulphur trioxide is achieved:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Because of the forward reaction is exothermic, there would be a higher percentage conversion of sulfur dioxide into sulfur trioxide at a low temperature. However, at a low temperature the rate of reaction would be very slow. 450^{o}C is a compromise.

There are 3 gas molecules on the left-hand side of the equation, but only 2 on the right. Reactions in which the numbers of gas molecules decrease are favored by high pressures. In this case, though, the conversion is good at low pressures that it isn't economically worthwhile to use higher ones.

The catalyst, vanadium(V) oxide, has no effect on the percentage conversion, but helps to speed up the reaction. Without the catalyst, the reaction would be extremely slow.

IV) Making the sulphuric acid

In principal you can react sulfur trioxide with water to make sulphuric acid. In practice, this produce an uncontrolled fog of concentrated sulfuric acid. Instead the sulfur trioxide is absorbed in concentrated sulfuric acid to give fuming sulfuric acid (oleum).

$$H_2SO_4(l) + SO_3(g) \rightarrow H_2S_2O_7(l)$$

This is converted into twice as much concentrated sulfuric acid by careful addition of water.

$H_2S_2O7(l) + H_2O(l) \rightarrow 2H_2SO_4(l)$

5.28 describe the use of sulfuric acid in the manufacture of detergents, fertilisers and paints

i. The production of fertilizers such as ammonium sulphate, potassium sulphate, calcium superphosphate, etc.

   Ammonia sulphate can be prepared by reacting ammonia and sulphuric acid.
   $NH_3\ (g) + H_2SO_4\ (l) \rightarrow (NH_4)_2SO_2$

   Sulphur phosphate is the main source of phosphorus in agriculture. To make super phosphate, calcium phosphate is treated with concentrated sulphuric acid, a product containing two calcium sulphate is form.
   $Ca_3(PO_4)_2\ (s) + 2H_2SO_4\ (aq) + 4H_2O\ (l) \rightarrow \text{Super Phosphate}\ [Ca(H_2PO_4)_2\ (s) + 2CaSO_4.2H_2O]$

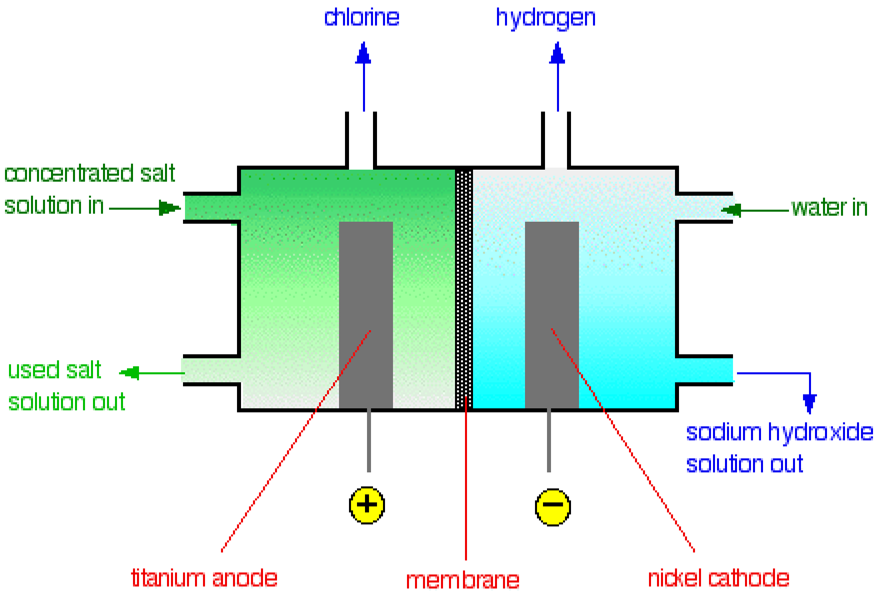
ii. The manufacture of non-soapy detergents: modern detergents are organic compounds ‘sulphonated’ with concentrated sulphuric acid.

   In the manufacture of detergent, hydrocarbons are treated with concentrated sulphuric acid to form an organic acid. The organic acid is then neutralized with sodium hydroxide solution to produce a detergent.
   Hydrocarbon + Concentrated sulphuric acid → Organic acid
   Organic acid + Sodium hydroxide solution → Detergent

iii. In paint manufacture, sulfuric acid is used in extracting the white pigment titanium dioxide, TiO_2, from titanium ores.
iv. The making of artificial silks like rayon: here, the fine threads in the alkaline cellulose solution are neutralized by passing them through a bath of sulphuric acid.
v. The cleaning of metals by removing the surface oxide coating: this is called pickling and is important in preparing articles for electroplating.
vi. It’s used as an electrolyte inside batteries for cars: most car batteries are made up of lead plates in a sulphuric acid electrolyte; occasionally, the electrolyte needs to be ‘topped up’ with distilled water; this is because small amounts of hydrogen and oxygen gases are given off by the chemical changes inside the battery, and therefore the sulphuric acid loses water and becomes too concentrated.

5.29 describe the manufacture of sodium hydroxide and chlorine by the electrolysis of concentrated sodium chloride solution (brine) in a diaphragm cell

The concentrated salt solution can be electrolysed to produce three useful chemicals – sodium hydroxide, chlorine and hydrogen. The electrolysis can be carried out in a diaphragm cell.



At the titanium anode, chloride ions are discharged to produce chlorine gas.

$2Cl^{-}(aq) \rightarrow Cl_2(g) + 2e^{-}$

At the steel cathode, it is too difficult sodium ions, so hydrogen ions from the water are discharged instead to produce hydrogen gas.

$2H^{+}(aq) + 2e^{-} \rightarrow H_2(g)$

More and more water keeps splitting up to replace the hydrogen ions as soon as they discharged. Each time a water molecule splits up it produces a hydroxide ion as well. That means that there will be a build-up of sodium ions and hydroxide ions in the right-hand compartment-sodium hydroxide solution is formed. This is contaminated with uncharged sodium chloride.

The sodium hydroxide solution is concentrated by evaporating it. During this process, most of the sodium chloride crystallizes out as solid salt. This can be separated, redissolved in water and recycled back through the cell again.

FAQ

What happens if chlorine comes into contact with sodium hydroxide or hydrogen?

The cell is designed to keep the products apart. If chlorine comes into contact with sodium hydroxide solution, it reacts to make bleach – a mixture of sodium chloride and sodium chlorate(I) solution. If chlorine comes into contact with hydrogen it produces a mixture which would explode violently on exposure to sunlight or heat to give hydrogen chloride.

Why is it important that the liquid level in the left hand compartment of a diagram cell is higher than that in the right hand side?

This is to make sure that the flow of liquid is always from left to right, preventing any of the sodium hydroxide solution formed finding its way back to where chlorine formed.

5.30 write ionic half-equations for the reactions at the electrodes in the diaphragm cell

(Follow 5.29)

5.31 describe important uses of sodium hydroxide, including the manufacture of bleach, paper and soap; and of chlorine, including sterilising water supplies and in the manufacture of bleach and hydrochloric acid.

Uses of sodium hydroxide include:

- The purification of bauxite to make aluminium oxide ('alumina') as a part of the manufacture of aluminium.
- Paper making – the sodium hydroxide helps break the wood down into pulp.
- Soap making – sodium hydroxide reacts with animal and vegetable fats and oils to make compounds, such as sodium stearate, that are present in soap.
- Making bleach – bleach is formed when sodium hydroxide and chlorine react together in the cold; it is a mixture of sodium chloride and sodium chlorate(I) solution.
  $2NaOH(aq) + Cl_2(g) \rightarrow NaCl(aq) + NaOCl(aq) + H_2O(l)$

Uses of chlorine include:

- Sterilising water to make it safe to drink
- Making hydrochloric acid (by controlled reaction with hydrogen)
- Making bleach, weedkillers, pesticides
- Making PVC (nearly 1/3 of it is used for this)
- Making solvents with perchlroethylene for degreasing and dry-cleaning
- Making paints and dye stuffs
- Making pharmaceuticals
- Making titanium dioxide, a white pigment used in paints, ceramics, cosmetics and paper.
- Killing bacteria in swimming pools