CHEMISTRY

"O" Level Notes

Syllabus Code 5070

Name: _____

My aim for O Level is _____

Class: _____

Author
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GCE O LEVEL SYLLABUS 2010

CHEMISTRY (5070)

Name: _____ Date: _____

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CHAPTER 1 – EXPERIMENTAL CHEMISTRY

1.1 Experimental Design

Volumes of Liquids

SI unit: **cubic metre (m^3)**

Large volume measurement: **decimetres (dm^3)**

$$1 \, dm^3 = 1\,000 \, cm^3$$

Daily life measurement: **millilitres (ml)** *or* **litres(l)**

$$1 \, litre = 1\,000 \, ml$$

Apparatus for measuring liquids depends on:

- The volume being measured
- How accurate the measurement needs to be

- *Beaker* hold approximate volume of 100 cm^3 and 250 cm^3.
- *Conical flask* hold approximate volume of 100 cm^3 and 250 cm^3.
- *Measuring cylinder* has accuracy to 1 cm^3.
  - Reading to be taken nearer to the meniscus (bottom line).
  - If reading is 23 cm^3, should not write 23.0 cm^3 as the '0' means accurate to 0.1 cm^3.
- *Burette* has long scale of 0 – 50 cm^3, accurate to 0.1 cm^3.
  - Liquid level to be measured before and after tap opening. The difference of volume gives the liquid volume poured off.
- *Bulb pipette* measures exact volumes such as 20.0, 25.0 or 50.0 cm^3, not odd volumes such as 31.0 cm^3.

| Conical flask | Beaker | Bulb Pipette | Burette | Measuring cylinder |
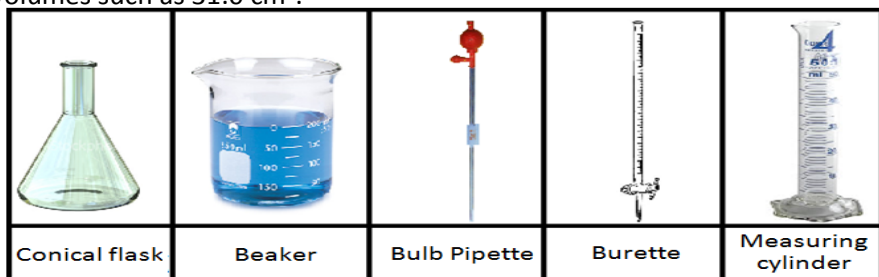
gas syringe

Volumes of Gases

Measured with gas syringe, up to 100 cm^3

Temperature

Measured with thermometer. 2 types are:

a) Mercury-in-glass

b) Alcohol-in-glass

SI Unit: **Kelvin (K)**

Daily life measurement: **degree Celcius (^{o}C)**

$$K = {}^oC + 273$$

Time

SI Unit: **seconds (s)**

Other Units: **minutes (min)/hour (h)**

Measured with:

(a) Clock

(b) Digital stopwatch

| Analogue Stopwatch | Digital Stopwatch |

Mass

Mass – the measure of amount of matter in a substance

SI Unit: **kilogram (kg)**

Other Units: **grams (g)/milligrams (mg)**

Large volume measurements: **tonnes (t)**

$$1 \, tonne = 1\,000 \, kg$$

Measured with:

(a) Electric "top-pan" balance

(b) Triple beam balance

| Triple beam balance | Electric top-pan balance |
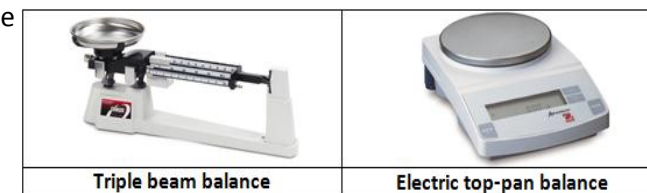
1.2 Methods of Purification and Analysis

Pure substance – single substance not mixed with anything else
E.g: white sugar, copper sulfate crystals, distilled water

Mixture – contains two or more substances. Its quantity is more on Earth.
E.g: seawater (salt, water & dissolved solids), milk (fats & dissolved solids)

2.2 Obtaining Pure Substances

Purification – The seperation process of mixtures into pure substances by using physical methods without chemical reactions.

Filtration

Filtration – separates insoluble solid from a liquid.

- Mixture is poured through a filter with tiny holes made of paper.
- Large solid particles cannot pass through the pores and trapped in it as residue while tiny liquid particles pass through as filtrate.
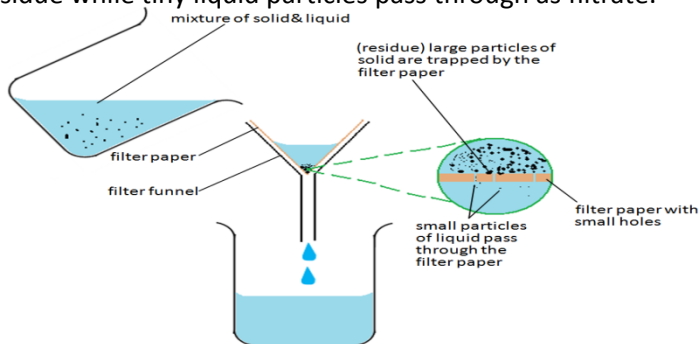
Crystallisation & Evaporation to Dryness

Crystallisation – separation of dissolved solid from a solution as well-formed crystals

Evaporation to Dryness – seperation of dissolved solid from a solution as crystals of salt by evaporating all the liquid off.

FULL OVERVIEW PROCESS ON THE RIGHT:

1. The impure solid is dissolved in a solvent.
2. The solution is heated to evaporate most of the solvent. To turn the process to evaporation to dryness, continue heating until salt is obtained.
3. The hot solution is allowed to cool. The solid appears as cool crystals.
4. The cold solution is poured off to obtain the crystals. The crystals may be dried by pressing them between sheets of filter paper.

Why crystallisation occur?

- Solubility of most solutes decrease as temperature decrease, when solution cools, solution can't hold more solute (*saturated*) so the extra solute separates as pure crystals.

Sublimation

Sublimation – separation of a mixture of solids which one of it sublimes (by heating the solid mixture to turn one of the substance into vapour without going through liquid state).
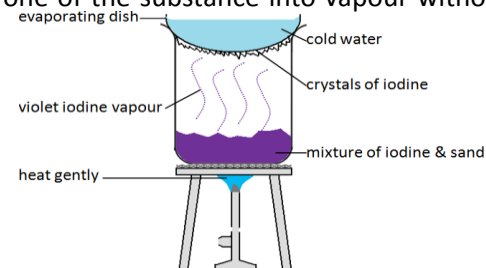
When mixture of iodine and sand is heated, iodine sublimes (turns into vapour directly) then cools and crystallise when it reaches cold water area

Examples of sublimable solids: CO_2 (s), dry $FeCl_3$ (s), dry $AlCl_3$ (s)

Simple Distillation

Simple Distillation – separation of pure liquid from a solution by condensing vaporised liquid

Condensed pure liquid – *distillate*

Diagram and Distillation Graph
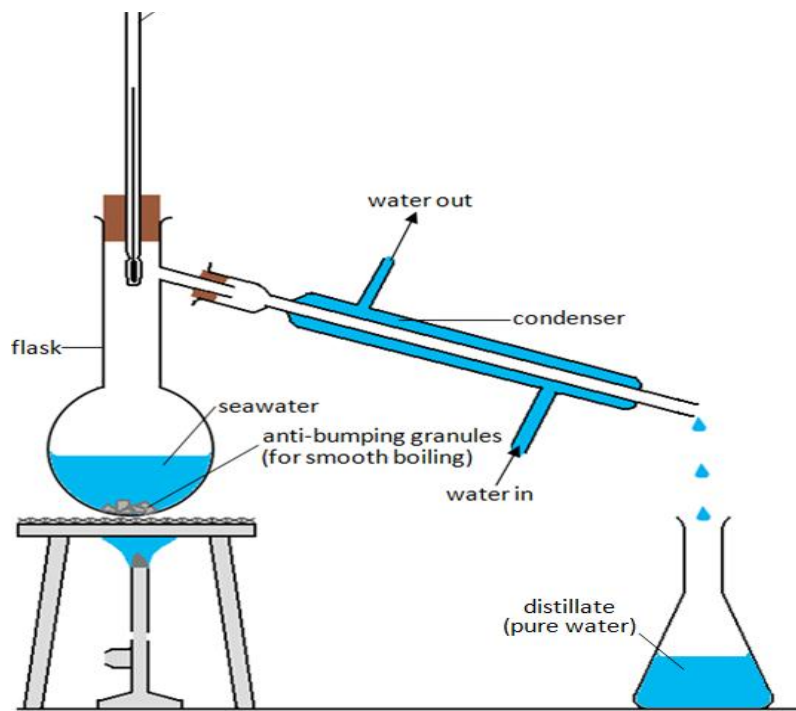
Process of Distillation:

Solution is heated, and steam (pure vapour) is produced. The steam is cooled in condenser to form pure liquid. Solute remains in the flask.

Fractional Distillation

Fractional Distillation – separates mixture of miscible (soluble) liquids with widely differing boiling points.

Use of **fractionationg column** separates them

Process of Fractional Distillation: E.g. ethanol and water

Mixture of ethanol and water is placed in flask and heated. Ethanol with lower boiling point boils and vaporises first and reach fractionating column then cools and condenses into ethanol as it passes through condenser. Temperature will stay constant until all ethanol is distilled. Water will distil the same way after all ethanol is distilled.
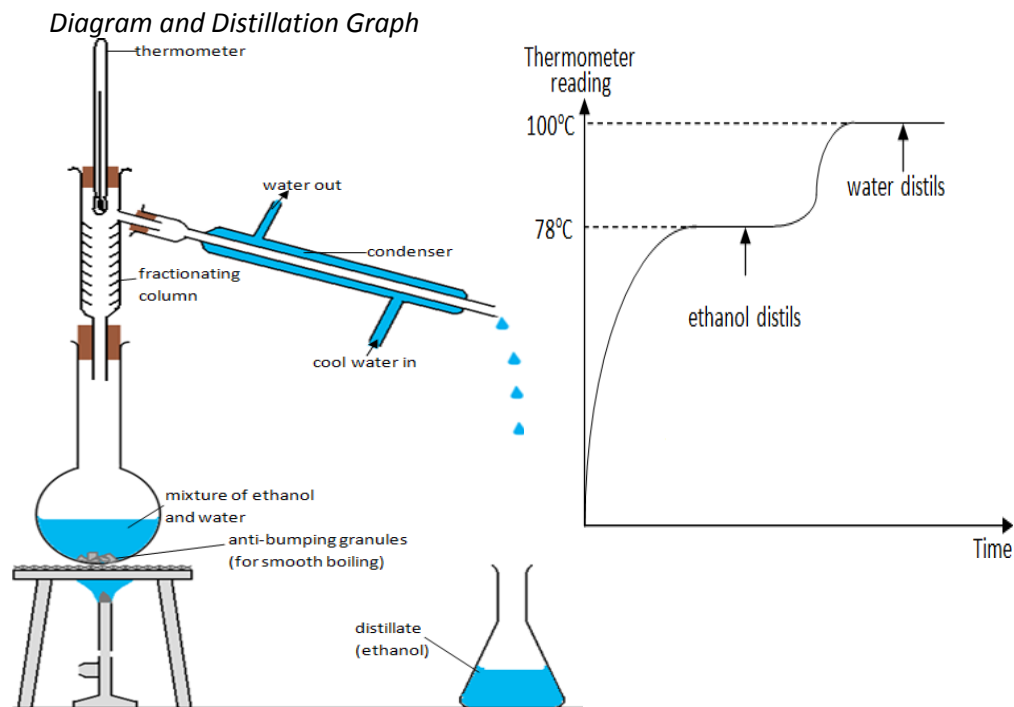
Uses of fractional distillation:

- Separates pure oxygen and pure nitrogen from liquefied air
- Separates substances in petroleum (crude oil) into fractions
- Separates alcohol to produce alcoholic drinks

Reverse Osmosis

Reverse Osmosis – separates a solution (e.g. seawater) by pressurizing the mixture against a membrane which separates the solute and the solvent
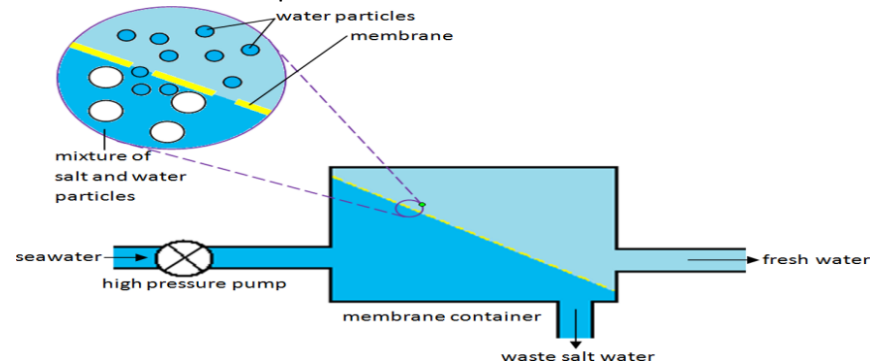
- Seawater is pumped under great pressure into a closed container onto a membrane forcing water particles but salt particles to pass through. Some salt particles may still pass through.

Use of Separating Funnel

Separating Funnel is used to separate immiscible liquids
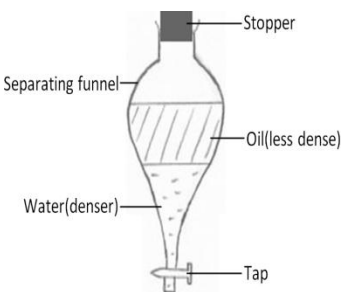
- two liquids insoluble to each other will create two layers of overlying liquids of each type. To separate, take the stopper off and turn the tap on to run the denser liquid at the bottom off the funnel and leave the less dense liquid in the funnel by turning the tap off and reset the stopper at its original position.

Chromatography

Chromatography – a method of separating and identifying mixtures.

The need for Chromatography
- Separates and identify mixtures of coloured substances in dyes
- Separates substances in urine, drugs & blood for medicinal uses
- To find out whether athletes have been using banned drugs

Separating Mixtures of Coloured Substances

Obtain a dye sample then put a drop of the sample on a pencil line drawn on the filter paper then dip the paper into a solvent with the level below the spot. The dye will dissolve in solvent and travel up the paper at different speed. Hence they are separated.

Identifying Mixturees of Coloured Substances

In the diagram on the right, drop of sample dye is placed on pencil line. The result shows that:

- The sample dye is made of 3 colours.
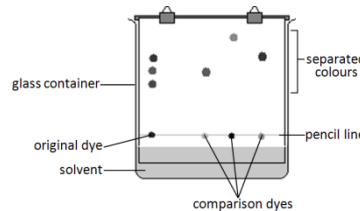
- 2 comparison dyes are of one of the compositions of the original dye as the spots are of same colour and distance.
- a comparison dye isn't part of sample.

Separating and Identifying Mixtures of Colourless Substances

To do this a locating agent is to be sprayed on filter paper.

Locating Agent – a substance that reacts with substances (e.g. sugars) on paper to procuce a coloured product.

R_f Values

To identify unknown dye in the diagram at the very top:

R_f value = $\dfrac{x}{y}$

Where x = distance moved by the substance and;

y = distance moved by the solvent

Checking the Purity of Substances

- Pure substances have FIXED MELTING AND BOILING POINTS.
- ❖ Pure water boils at $100^{o}C$ and melts at $0^{o}C$.

- Impure substances have NO FIXED MELTING AND BOILING POINTS. They melt and boil at a RANGE OF TEMPERATURES
- ❖ e.g. starts boil at $70^{o}C$, completes boil at $78^{o}C$
  - ➤ Also, it can VARY melting and boiling points of pure substances.
- ❖ e.g. pure water boil at $100^{o}C$, but with salt is at $102^{o}C$

1.3 Identification of Ions and Gases

Refer to Insert 1. Everything lies there.

END OF CHAPTER 1

CHAPTER 2 – THE PARTICULATE NATURE OF MATTER

2.1 Kinetic Particle Theory

Matter – anything that has mass and takes up space. Three forms – solids, liquids, gas.

SOLIDS
- fixed volume
- fixed shape
- incompressible
- do not flow

LIQUIDS
- fixed volume
- no fixed shape – takes the shape of container
- incompressible
- flow easily

GASES
- no fixed volume
- no fixed shape
- compressible
- flow in all direction

The Kinetic Particle Theory of Matter
- particles are too small to be seen directly
- there are spaces between particles of matter; the amount of space varies between each states
- the particles are constantly move; each state moves in different speed

DIFFUSION

Diffusion is the spreading and mixing of particles in gases and liquids.

Diffusion of gases

Bromine drops are placed into a jar. Another jar full of air is placed on top of jar with bromine, separated with cover. Cover is removed and bromine evaporates, filling both jars with dense reddish-brown bromine vapour.

Explanation:

Bromine particles move from lower jar into spaces between air particles in upper jar. At the same time, air particles move down from upper jar to mix with bromine particles in lower jar. Eventually, bromine and air particles are mixed completely.
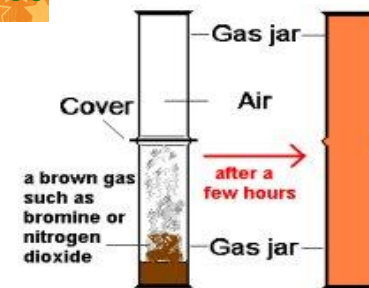
Diffusion of liquids
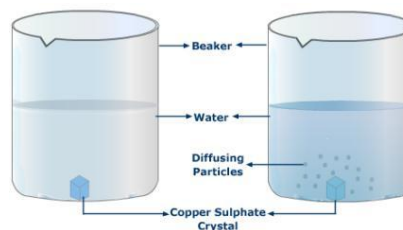
$CuSO_4$ crystals placed in beaker of water, blue particles of the crystals is spread throughout the water to form uniformly blue solution.

Factors Affecting Rate of Diffusion

- **Temperature**
  The higher the temperature, the more particles of matter absorb energy making them move faster, the higher the rate of diffusion; the lower the temperature, the slower the rate of diffusion
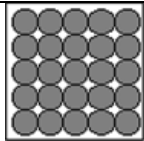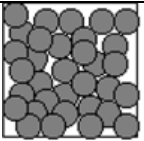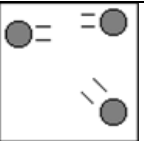
- **Mass of particles**
  Greater mass, the slower it diffuses; Smaller mass, the faster it diffuses
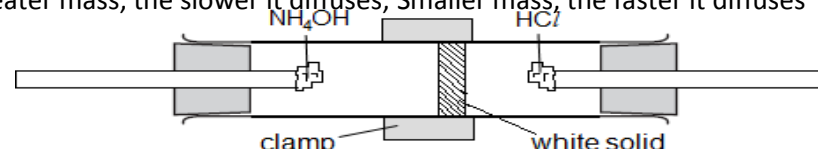
A cotton soaked in aqueous ammonia and another soaked in hydrochloric acid are placed on opposite sides of the tube. NH_4OH vapor and HCl vapor diffuses in the tube and a compound is produced inside the tube closer to HCl soaked cotton as the particles are heavier. The greater mass, the slower particles diffuse. The smaller mass, the faster particles diffuse.

Particulate Models of Matter

Solid	Liquid	Gas
Particles in solid: - Are packed close together in orderly arrangement - Have little empty space between them - Can vibrate but cannot move freely about their fixed position	**Particles in liquid:** - Are packed closely but not orderly arranged - Have little empty space between them but more than in solids - Are not held fixed but free to move throughout liquid	**Particles in gas:** - Are far apart and in random arrangement - Are free to move anywhere in the container

Differences between properties of matter and particles in them.

1. Matter can be coloured (e.g. sulphur is yellow) but particles are not.
2. Substances feels hot/cold but particles don't get hot/cold. The temperature is due to speed of movement of particles. If hot, particles move fast.
3. Matter expands when heated but particles don't. They increase distance between particles during expansion.

Changes of State

Melting
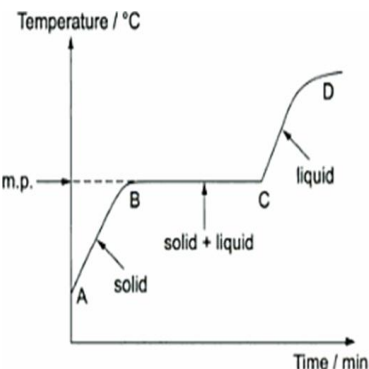
Melting is change from solid to liquid by absorbing heat to break force of attraction holding particles together. The temperature at which solid melts is *melting point*.

From the graph:

A-B: the temperature of solid increases to melting point.

B-C: the temperature remains constant as heat is absorbed to break forces of attraction instead for raising temperature. Solid and liquid are present.

C-D: liquid heats as heat energy increases temperature.

Freezing

Freezing is the change of liquid to solid by cooling down of liquid.

Freezing point is the temperature at which liquid freezes.
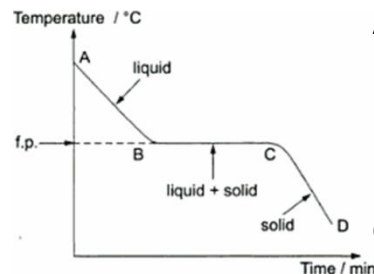
A-B: liquid temperature decreases to freezing point.

B-C: heat energy is released as particles slow down to take up fixed and orderly position of a solid. The temperature remain constant release of energy compensates for loss of heat to surroundings.

C-D: solid cools to the temperature of surroundings.

Boiling

Boiling is the change of liquid to gas by absorbing heat to break the forces holding them together.

Boiling point is the temperature at which liquid boils.
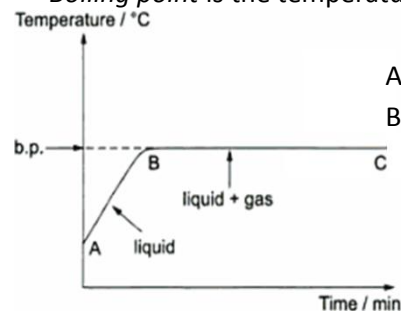
A-B: liquid temperature rises to boiling point.

B-C: heat energy is absorbed by particles to break the attractive forces so that they move freely and far apart as gas particles. That's why the temperature remain constant

Evaporation

Evaporation is change of liquid to gas without boiling, occurs below boiling point on water surface. It gives cooling effect – heat energy absorbed from surroundings.

Condensation

Condensation is the change of gas to liquid. Heat energy is given out as gas particles slow down and move closer to one another to form liquid.

Sublimation is the change of solid to gas without melting. Heat is absorbed.

2.2 Atomic Structure

Atoms contain PROTONS, NEUTRONS, and ELECTRONS

Protons have positive charge while neutrons has neutral charge but same mass as protons. Since an atom is electrically neutral, electrons has to carry a negative charge and the amount of electrons is the same as the amount of protons.

Particle	Symbol	Relative mass	Charge
Proton	p	1	+ 1
Neutron	n	1	0
Electron	e^-	$\dfrac{1}{1836}$	- 1

Protons and neutrons are located in nucleus. These make up **nucleon number**.

Electrons move around nucleus in an orbit called **electron shells**.
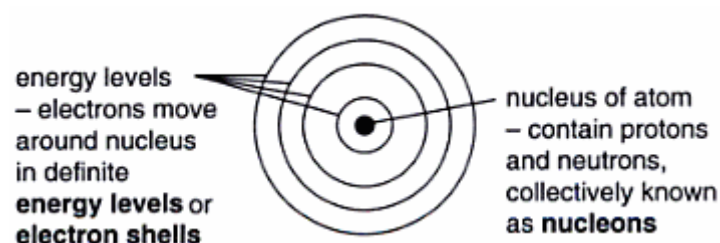
PROTON NUMBER is the number of protons in an atom.

NUCLEON NUMBER is the number of protons and neutrons in nucleus of an atom. Therefore, to find the number of neutrons, we subtract proton number from nucleon number, i.e.: **Nucleon number – Proton number = Neutrons**

ELECTRONS have the same number as protons to balance the charges.

$$\text{nucleon number} \longrightarrow {}_b^a X \longleftarrow \text{symbol of element}$$
$$\text{proton number} \longrightarrow$$
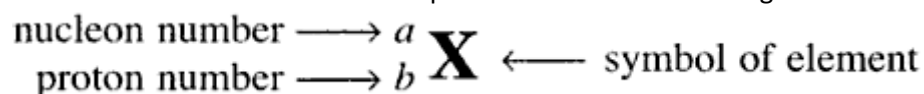
ISOTOPES are atoms of the same element with different number of neutrons. Therefore, their **nucleon number is different**.

E.g. Hydrogen atoms has 3 isotopes, ${}_1^1H$, ${}_1^2H$ and ${}_1^3H$. Structurally, it's drawn:

ELECTRONIC CONFIGURATION

Electrons are placed in orbits. First shell contains maximum 2 electrons. Second shell and so and so for has maximum of 8 electrons.
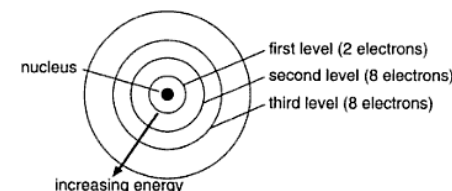
To write electronic configuration we write as *n.n.n....* where first *n* denotes the first shell, second the second shell and so and so for.

E.g. Sulfur has electronic configuration of 2.8.6
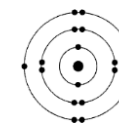
The valence electrons is the number of electrons of the outermost shell. Sulphur has 6 valence electrons.

Relation with Periodic Table

Elements in same horizontal row: **Period**

Elements in same vertical column: **Group**

Group 1 has 1 valency, Group 2 has 2 valency, Group 3 has 3 valency and so on.

Group 0 has full valency which makes it having **stable electronic configuration**.

Down the period the number of shells increases.

2.3 Structure and Properties of Materials

Elements

Element is a substance that cannot be broken down into simpler substances by chemical nor physical methods.

Classifying Elements
- **Classifying by state**. E.g. some elements are solids, some liquids, some gases.
- **Classifying by metals and non-metals**. E.g. most elements are metals, semi-metals are *metalloids* (having properties of metals & non-metals), some are non-metals
- **Classifying by periodicity**.From left-right elements change from metal to non-metal

COMPOSITION OF ELEMENTS

Elements are made of atoms

Atom is smallest unit of an element, having properties of that element.

Molecule is group of two or more atoms chemically joined together, e.g. chlorine molecule has 2 chlorine atoms

Chemical formula shows the number and kinds of atoms in a molecule, e.g. chlorine molecule has formula Cl_2, where Cl is chlorine symbol and the subscript number (2) shows that there are 2 atoms in a chlorine gas molecule.

Compounds

Compound is substance containing 2 or more elements chemically joined together e.g. Magnesium is an element; oxygen is an element – they can only be burnt to form magnesium oxide compound.

COMPOSITION OF COMPOUNDS

Ions or molecules make up compounds

Ions are atoms having electrical charge

E.g. $NaCl$ made up of 2 ions; positively charged Na, negatively charged Cl.

Mixtures

Mixture contains 2 or more substances not chemically joined together.

e.g. seawater is made up of water and $NaCl$ (salt); oxygen in air varies.

	Compound	Mixture
1	Component substances cannot be separated by physical methods; a chemical method is required for separation.	Component substances can be separated by physical methods.
2	Its physical and chemical properties are different from that of its elements.	Its physical and chemical properties are the same as its component substances.
3	Fixed composition by mass.	Variable composition by mass.
4	Has fixed melting and boiling points.	Has variable melting and boiling points.

2.4 Ionic Bonding

Ionic bonding is the transfer of electrons from one atom to another to become achieve an inert gas configuration, forming ions.

Ionic bonds are formed between **METALLIC** and **NON- METALLIC ATOMS** ONLY.
- Metals lose electrons to form positive ions (cations)
- Non-metals gain electrons to form negative ions (anions)

The formation of ions is resulted from transfer of atoms from one atom to another atom(s), which the ions produced are of opposite charges, and unlike charges attract, causing them to be held together with a strong force.

E.g. Formation of NaCl

sodium atom, Na (2.8.1)   chlorine atom, Cl (2.8.7)   sodium ion, Na^+ (2.8)   chloride ion, Cl^- (2.8.8)
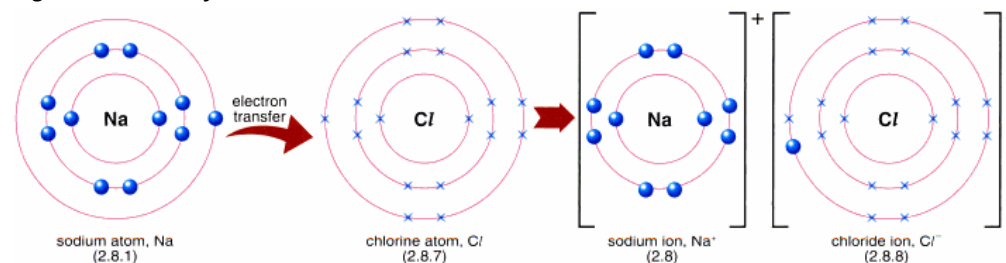
Sodium atom loses an electron by transferring the electron to chlorine atom, making both stable. The loss of electron forms cation, Na^+, and the gain of electron forms anion, Cl^-. The opposite charges acquired by both ions attract to each other, forming a strong ionic bond of $NaCl$.

E.g. Formation of MgF$_2$

magnesium atom, Mg (2.8.2) — fluorine atoms, F (2.7) — magnesium ion, Mg^{2+} (2.8) — fluoride ions, F$^-$ (2.8)
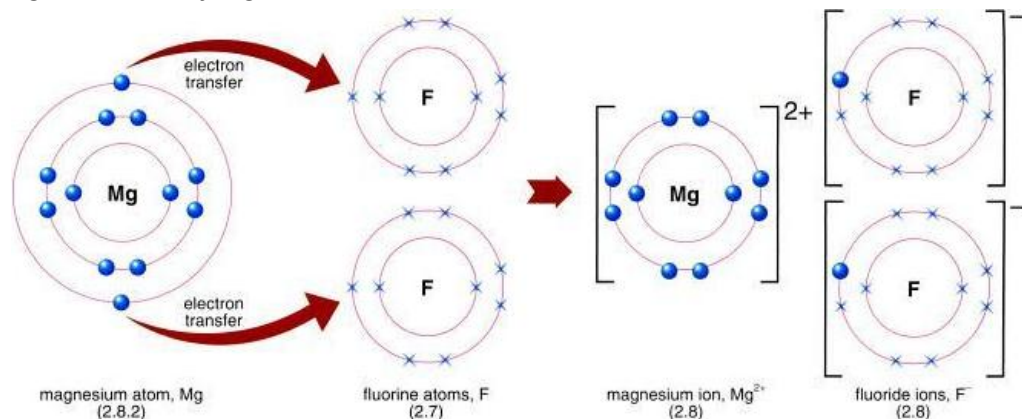
Sodium atom loses two electrons by transferring the electrons to fluorine atoms, one each, making both stable. The loss of electron forms cation, Mg^{2+}, as it loses 2 electrons, and the gain of electron forms anion, F$^-$. The opposite charges acquired by both ions attract to each other, forming a strong ionic bond of MgF$_2$.

Deducing formula ionic compounds

We can know the charge of elements by looking at groups of periodic table. Group I to group III elements has charge of +1, increasing to +3, going to the right. Group V to group VII elements has charge of -3, decreasing to -1, going to the right.

E.g. Aluminium sulfate

We have to balance the charges to make a stable bond

Ions present: Al^{3+} SO$_4^{2-}$

           Al^{3+} SO$_4^{2-}$

                    SO$_4^{2-}$
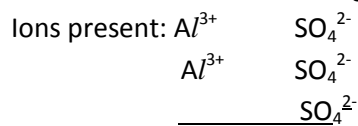
Total change: 6+   6-

Therefore, the formula is Al_2(SO$_4$)$_3$

1. The symbol of metal ion should always be first, e.g. NaCl
2. Polyatomic ion should be placed in brackets, e.g. Fe(NO$_3$)$_2$

STRUCTURE AND PROPERTIES OF IONIC BONDS

Structure

Ionic substances appear as giant lattice structures which the ions are held together by electrostatic force between oppositely charged ions. To find the formula of ionic bond, say sodium chloride bond, by looking at lattice structure, we count the ratio of amount of metal ions to non-metal ions. E.g. in sodium chloride, the ratio Na:Cl is 1:1, therefore the ionic formula is NaCl.

= Na$^+$ = Cl^-
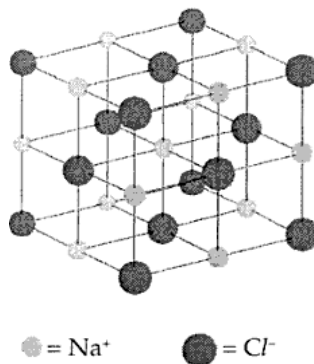
Properties

1. Ionic compounds are hard crystalline solids with flat sides and regular shapes because the ions are arrnged in straight rows in strong ionic bonds.
2. Ionic compounds have very high melting points and boiling points.
3. The strong forces holding ionic compounds prevents them to evaporate easily. Hence, ionic compounds have no smell.
4. Solid ionic compounds don't conduct electricity but they do when they are aqueous or molten. This is because in liquid/aqueous state the ions which conduct electricity are free to move. In solids, these ions are fixed in place.
5. Ionic compounds are soluble in water but insoluble in organic compounds. This is because the ions attract water molecules which distrupts the crystal structure, causing them separate & go into solution. Vice versa is when in organic solvent.

2.5 Covalent Bonding

Covalent bonding is the sharing a pair of electrons to gain electronic configuration of an inert gas, usually for molecules.

Covalent bonds occur between **NON-METALLIC ATOMS** ONLY.

In covalent bond, **WE TRY TO SUBTITUTE THE SHORT OF ELECTRONS OF TWO/MORE ATOMS BETWEEN EACH OTHER TO FORM THE 2 OR 8 VALENCE ELECTRONS**. **THE SHARED ELECTRONS APPEAR IN PAIRS!**

E.g. H_2 molecule

hydrogen atom, H   hydrogen atom, H   hydrogen molecule, H_2
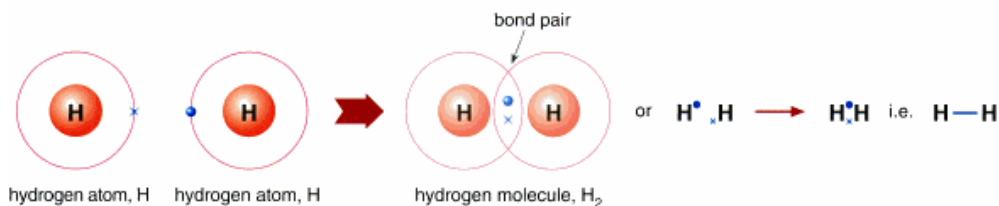
Hydrogen atom has one valency. To become stable with hydrogen atom, it needs one more electron, just like helium which has 2 valency. When 2 hydrogen atoms join, they share their electrons, on which, the share becomes 2 electrons, which is now a noble gas configuration, being shared between these 2 atoms. We write the bond as H – H single bond, which means they share an electron pair (2 electrons).

E.g. Cl_2 molecule

chlorine atom, Cl chlorine atom, Cl chlorine molecule, Cl_2
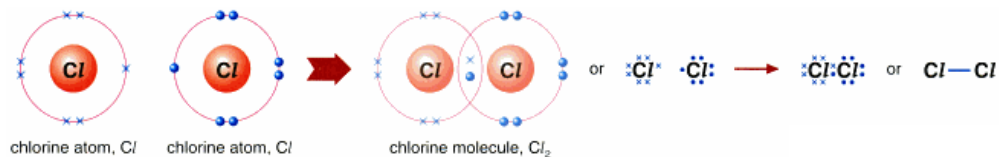
C*l* atom has 7 valency and needs one electron, each, to form a noble gas configuration between two C*l* atoms. Hence they share an electron EACH to hence share 2 electrons between the atoms. Hence, each C*l* atom now has 8 valency which is a noble gas configuration.

E.g.O_2 molecule

oxygen atom, O   oxygen atom, O   oxygen molecule, O_2
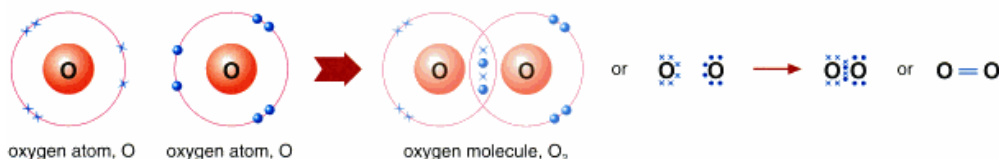
An O atom has 6 valency and needs 2 electrons, each, to form a noble gas configuration. Hence, EACH SHARE THE AMOUNT OF ELECTRONS EACH SHORT OF, in this case – 2 electrons, to form stable molecule. The contribution hence now become 4 electrons and what left on each oxygen atom is 4 electrons. We combine each 4 electrons on oxygen atom with the 4 electrons shared and hence we get 8 valency for each oxygen atom – a noble gas configuration!

E.g. H_2O molecule
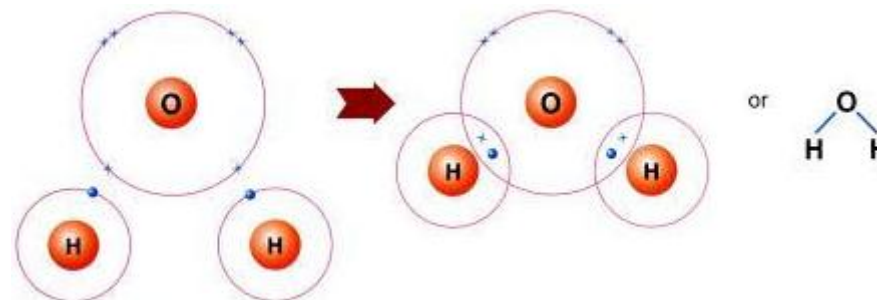
Apart from oxygen sharing between oxygen atoms, it can have electrons with other atoms. Oxygen needs 2 electrons and when bonded with hydrogen, which need an atom each, they combine to provide 2 electrons on both sides of oxygen bonded with hydrogen atoms. Each hydrogen with oxygen atom form a single bond: O – H.

E.g. CO_2 molecule
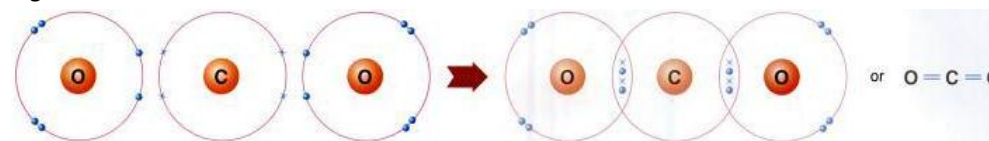
Carbon needs 4, oxygen needs 2. We share two from oxygen part, WHICH HAS THE SMALLEST NUMBER OF SHORT ELECTRONS, TO SHARE THE AMOUNT OF ELECTRONS THAT ATOM NEEDS, to form 4 shared atoms. Now oxygen is stable but carbon needs 2 more, which we now know they can get from another oxygen atom. The atoms are now stable and since each bond has 2 pairs of electrons, we call this double bond: C = O.

A pair of shared electrons between 2 atoms forms **SINGLE BOND**, X – Y.
Two pairs of shared electrons between 2 atoms forms **DOUBLE BOND**, X = Y.
Three pairs of shared electrons between 2 atoms forms **TRIPLE BOND**, X $\equiv$ Y.

This information is important when you want to know the bond forces between atoms in exothermic/endothermic reactions.

STRUCTURE AND PROPERTIES OF COVALENT BONDS

Structure

Giant Covalent Bond

Diamond

Diamond has carbon atoms bonded with another carbon atoms in a tetrahedral arrangement which each carbon atom uses all its valence electrons to form 4 single covalent bonds with other 4 carbon atoms.

Silicon Dioxide

Silicon dioxide, SiO_2, has silicon atoms bonded with another oxygen atoms in a tetrahedral arrangement which each silicon atom uses all its valence electrons to form 4 single covalent bonds with other 4 oxygen atoms.

Graphite

Graphite has flat layers of carbon atoms bonded strongly in hexagonal arrangement in which the layers are bonded to each other weakly.
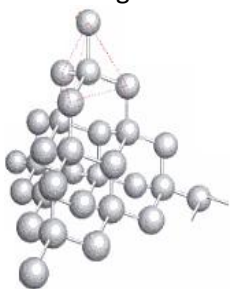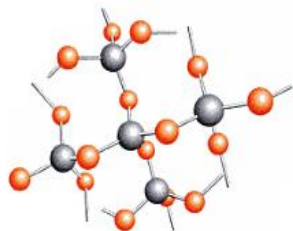
arrangement of atoms in one layer   arrangement of layers

strong covalent bond

weak force between layers

Diamond   Silicon Dioxide   Graphite

Properties
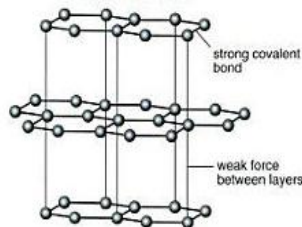
1. It is a hard solid because it consists of many strong covalent bonds between atoms. This property makes it suitable as abrasives.
2. It has *very high* melting and boiling points.
3. It does not conduct electricity (except graphite) because there are no free electrons in covalent bonds since they are used to form bonds; hence electrons are in fixed positions. To conduct electricity, there must be free electrons.
4. All covalent structures are insoluble in water.

2.6 Metallic Bonding

Metallic bonding is bonding within atoms of metals caused by attractive force between positively charged metal ions and negatively charged free electrons. The atoms are packed closely together in **giant lattice structures**.

BOND FORMING

Each atom in metal gives up valence electrons to form positive ions. There are free electrons moving between the spaces and positive metal ions are attracted to the sea of electrons which hold the atoms together.

STRUCTURE AND PROPERTIES OF METALLIC BONDS

1. Metals can be bent (ductile) and can be stretched (malleable) because the layers of atoms in metals slide over each other when force is applied but will not break due to attractive force between electrons and metal ions.
2. Metals conduct electricity as it has free electrons which carries current.
3. Metals conduct heat as it has free electrons which gains energy when heated and moves faster to collide with metal atoms, releasing heat in collisions.
4. Metals have high melting and boiling points because the bonds between metals is very strong. Hence very high heat energy needed to break the bonds.

2.7 Simple Molecular Substances

1. Simple molecular substances are usually liquids/gases at r.t.p. because the molecules are not tightly bonded like in solids, hence free to move.
2. They have low melting and boiling points because the force of attraction is weak that they can be easily broken by heat.
3. Since they have low boiling points, they evaporate easily.
4. They don't conduct electricity because they don't have free electrons/ions which helps to conduct electricity.
5. Most of these are insoluble in water but soluble in organic solvent.

END OF CHAPTER 2

CHAPTER 3 – FORMULAE, STOICHIOMETRY AND THE MOLE CONCEPT

3.1 Relative Atomic Mass

Comparing Atomic Masses with the Carbon Atom

To compare to a carbon atom, a carbon-12 atom is used. The mass of the isotope is 12 times greater than hydrogen atom so $\frac{1}{12}$ of carbon-12 atoms is equivalent to the mass of one hydrogen atom.

Relative Atomic Mass - the average mass of one atom of the element (averaging isotopes) when compared with $\frac{1}{12}$ mass of a carbon-12 atom.

$$\therefore A_r = \frac{Average\ Mass\ of\ one\ atom\ of\ the\ element}{\frac{1}{12} \times the\ mass\ of\ one\ atom\ of\ carbon-12}$$

In short is:

$$A_r = \frac{\frac{total\ mass\ of\ all\ isotopes}{number\ of\ isotopes\ formed}}{\frac{1}{12} \times 12}$$

The Relative Atomic Masses are already stated on the periodic table above each chemical formula.

3.2 Relative Molecular Mass and Relative Formula Mass

Using A_r, we calculate Relative Masses of molecules and ionic compounds

Relative Molecular Mass

Molecules contains atoms joined together, e.g. Cl_2

Average mass (molecular mass) of Cl_2= add relative masses of both atoms.

Relative Molecular Mass – the average mass of one molecule of substance (averaging isotopes) when compared with $\frac{1}{12}$ mass of a carbon-12 atom.

$$In\ short:\ M_r = \frac{total\ mass\ of\ all\ atoms}{\frac{1}{12} \times 12}$$

Relative Formula Mass– same as relative molecular mass but for ions only

Relative Formula Mass– total A_r of all atoms in formula of ionic compound

e.g. Relative formula mass of $MgSO_4$?

$$M_r = 24 + 32 + 4(16) = 120$$

3.3 Percentage Composition

e.g. Determine which oxides of iron of Fe_2O_3 or Fe_3O_4 has more iron.

Solution next page

$M_r(Fe_2O_3) = 2(56) + 3(16) = 160$

Percentage of Fe in $Fe_2O_3 = \frac{A_r(Fe_2)}{M_r(Fe_2O_3)}$ x 100 %

$\qquad = \frac{2(56)}{160}$ x 100%

$\qquad = 70\%$

$M_r(Fe_3O_4) = 3(56) + 4(16) = 232$

Percentage of Fe in $Fe_2O_3 = \frac{A_r(Fe_3)}{M_r(Fe_3O_4)}$ x 100 %

$\qquad = \frac{3(56)}{232}$ x 100%

$\qquad = 72\%$

$\therefore$ Fe_3O_4 has more iron composition than that of Fe_2O_3.

Calculating the Mass of an Element in a Compound

Use the example of Fe_2O_3 in the example above. The percentage mass of iron in iron(III) oxide is 70%. Therefore to calculate mass of iron in a 200g compound of iron(III) oxide is (0.7 x 200)g = 140g

e.g. Determine the mass of iron in 200g of Fe_2O_3.

$M_r(Fe_2O_3)= 2(56) + 3(16) = 160$

Mass of Fe in $Fe_2O_3 = \frac{A_r(Fe_2)}{M_r(Fe_2O_3)}$ x 200g

$\qquad = \frac{2(56)}{160}$ x 200g

$\qquad = 140g$

Calculating the Mass of Water in a Compound

Compound with water mass is 'hydrated' and has H_2O in their formula.

e.g. Calculate water mass in 12.5g hydrated copper sulfate, $CuSO_4 \cdot 5H_2O$

Mass of $5H_2O$ in $CuSO_4 \cdot 5H_2O = \frac{A_r(5\ H_2O)}{M_r(CuSO_4 \cdot 5\ H_2O)}$ x mass of sample

$\qquad = \frac{5(18)}{250}$ x 12.5g

$\qquad = 4.5g$

MOLE

3.4 Counting Particles

Unit for particles = **mole**

Symbol = **mol**

1 mol = 6 x 10^{23} atoms

3.5 Moles of Particles

Calculating the Number of Moles

$$n = \frac{Number\ of\ particles}{6 \times 10^{23}}$$

e.g 1: How many molecules in 6×10^{24} molecules of water, H_2O?

$$n = \frac{3 \times 10^{24}}{6 \times 10^{23}}$$

$$= 5\ mol$$

e.g 2: Calculate the number of molecules in 0.25 mole of CO_2. Hence, how many atoms are present?

$$0.25 mol = \frac{Number\ of\ particles}{6 \times 10^{23}}$$

Number of particles = 0.25 mol x 6×10^{23}

$$= 1.5 \times 10^{23}\ molecules$$

Number of atoms = total number of atoms in CO_2 x noumber of particles

$$= 3 \times 1.5 \times 10^{23}$$

$$= 4.5 \times 10^{23}\ atoms$$

3.6 Molar Mass

Molar mass – the mass of one mole of any substances

For substances consisting of atoms

It is the A_r of the element in grams. Eg. $A_r(C) = 12$, molar mass = 12g

For substances consisting of molecules

It is the A_r of the substance in grams. Eg. $A_r(H_2O) = 18$, molar mass = 18g

For substances consisting of ions

It is the A_r of substance in grams. Eg. $A_r(NaCl) = 58.5$, molar mass = 58.5g

Calculations Using Molar Mass

$$n = \frac{mass\ (m)}{molar\ mass\ (M_r)}$$

e.g. Find the mass of 0.4 mol of iron atom.

$$n = \frac{m}{M_r}$$

$$m = n \times M_r$$

$$m = 0.4 \times 56 = 22.4\ g$$

e.g. Argon Fluorohydride gas, HArF, first known noble gas compound, has molar mass of 60g. Find the number of moles Argon atom in 6.66g of HArF.

$$n\ (HArF) = \frac{6.66}{60}$$

$$= 0.111\ mol$$

$$n\ (Ar) = 0.111\ mol \times 1\ Ar\ atom\ in\ HArF$$

$$= 0.111\ mol$$

3.7 Different Kinds of Chemical Formulae

Ethene formula is C_2H_6

Molecular Formula – shows the actual formula and kinds of atoms present, e.g. C_2H_6

Empirical Formula – shows the simplest whole number ratio of the atoms present, e.g. C_2H_6, ratio 1:3, therefore C_1H_3, simply CH_3

Structural Formula – shows how atoms are joined in the molecule. It can be represented by ball-and-stick model or diagrammatically.
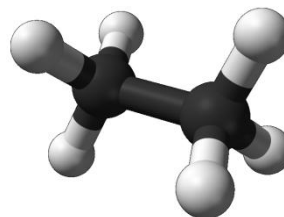
Ball-and-Stick   *Diagrammatic*
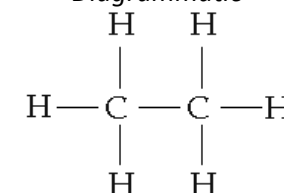
Calculating the Empirical Formula of a Compound

Find the empirical formula of an oxide of magnesium consisting of 0.32g of oxygen and 0.96g of magnesium.

Step 1: find the number of moles of the 2 elements.

$$n(Mg) = \frac{0.96g}{24g} \qquad n(O) = \frac{0.32g}{16g}$$

$$= 0.04\ mol \qquad\qquad = 0.02\ mol$$

Step 2: Divide the moles by the smallest number.

$$Mg = \frac{0.04}{0.02} \qquad O = \frac{0.04}{0.02}$$

$$= 2 \qquad\qquad = 1$$

Therefore, the empirical formula is Mg_2O

Calculating the Empirical Formula from Percentage Composition

An oxide of sulphur consists of 40% sulphur and 60% oxygen.

Take the total 100% to be 100g.

Step 1: find the number of moles of the 2 elements.

$n(S) = \frac{40g}{32g}$ $\qquad$ $n(O) = \frac{60g}{16g}$

$\quad$ = 1.25 mol $\qquad$ = 3.75 mol

Step 2: Divide the moles by the smallest number.

$S = \frac{1.25}{1.25}$ $\qquad$ $O = \frac{3.75}{1.25}$

$\quad$ = 1 $\qquad$ = 3

Therefore, the empirical formula is SO_3

From Empirical formula to Molecular Formula

Find the molecular formula of propene, CH_2, having molecular mass of 42.

Molecular formula will be C_nH_{2n}

Relative molecular mass = 12n(from carbon A_r) + 2n(2 x hydrogen A_r) = 14n

14 n = 42 $\qquad$ Therefore, C_3H_6

$n = \frac{42}{14} = 3$

3.8 Molar Volume of Gases

The Avogadro's Law

Equal volume of gases at same temperature and volume contain equal number of particles or molecules.

Molar Volume of Gas – volume occupied by one mole of gas

All gases at room temperature and pressure (r.t.p.) = $24dm^3$

$$1dm^3 = 1000cm^3$$

Formulae:

$$\text{Number of moles of a gas (n)} = \frac{volume\ of\ gas\ (V)}{molar\ volume\ of\ gas\ (M_r)}$$

Volume of a gas = Number of moles (n) x Molar volume (M_r)

e.g. What is the number of moles of $240cm^3$ of Cl_2 at r.t.p.?

$$n = \frac{volume\ of\ gas\ (V)}{molar\ volume\ of\ gas\ (M_r)}$$

$$= \frac{240}{24\,000} \text{ (or } \frac{0.24}{24} \text{ using } dm^3)$$

$$= 0.01 \text{ mol}$$

Molar Volume and Molar Mass

Gases have same volume but not necessarily same mass

Example: Hydrogen -> 2g, Carbon Dioxide -> 44g

e.g. Find the volume of 7g of N_2 at r.t.p.

Step 1: Find the number of moles from the mass of nitrogen

$$n = \frac{7g}{28g} = 0.25 \text{ mol}$$

Step 2: Find the volume of nitrogen, now with formula of gas

$$0.25 \text{ mol} = \frac{volume\ of\ gas\ (V)}{24}$$

Volume of gas = 0.25 mol x 24

$$= 6 \text{ } dm^3 \text{ (or } 6000cm^3)$$

3.9 Concentration of Solutions

Concentration of solution tells the number of solute in a volume of solution

$$\text{Concentration (C)} = \frac{number\ of\ grams\ \left(if\ \frac{g}{dm^3}\right) or\ moles\ \left(if\ \frac{mol}{dm^3}\right) of\ solute}{Volume\ of\ solution\ (in\ dm^3)}$$

Calculating the Amount of Solute

$$\text{Moles of solute (n)} = \text{Concentration } \left(\frac{mol}{dm^3}\right) \text{ x Volume of solution } (dm^3)$$

e.g. What is the mass of solute in $600cm^3$ of 1.5 $\frac{mol}{dm^3}$ NaOH solution?

Volume of solution in dm^3 = 0.60 dm^3

$$n = 1.5 \text{ x } 0.60$$

$$= 0.9 \text{ mol}$$

Number of moles of NaOH = $\frac{m}{M_r}$

$$0.9 = \frac{m}{40}$$

$$m = 0.9 \text{ x } 40$$

$$= 36g$$

CALCULATIONS USING CHEMICAL EQUATIONS

3.10 Constructing Chemical Equations

E.g. 1: Reaction Between Hydrogen and Oxygen

Word Equation: Oxygen + Hydrogen → Water

To write the **chemical equation**, we use symbols of atoms/molecules:

$$O_2 + H_2 \rightarrow H_2O$$
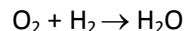
BUT THIS IS IMBALANCED! A BALANCED EQUATION **MUST** HAVE THE SAME NUMBER OF ATOMS OF EACH ELEMENTS ON BOTH SIDES! THEREFORE...

$$O_2 + H_2 \rightarrow H_2O$$

O	H	H
O	H	H
		O
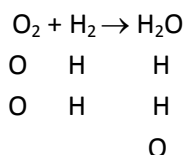
From above, we know that H_2O is short 1 oxygen atom. Therefore we multiply product by 2 first. Note: all atoms in molecules are automatically multiplied by 2.

$$O_2 + H_2 \rightarrow 2H_2O$$

Ø	H	H
Ø	H	H
		H
		H
		Ø
		Ø
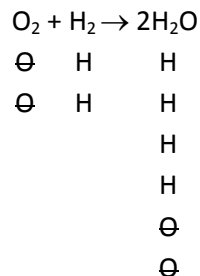
Now we can cancel off oxygen atoms. However, hydrogen atoms from reactant is short 2 atoms. Therefore, we multiply the hydrogen molecule by 2 so that the short is balanced. The equation is fully balanced when we are able to cancel off all atoms of that element on both sides.

$$O_2 + 2H_2 \rightarrow 2H_2O$$

Ø	H̶	H̶
Ø	H̶	H̶
	H̶	H̶
	H̶	H̶
		Ø
		Ø
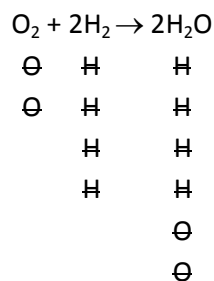
3.11 Calculations from Equations

Reacting Masses

In every equation, each atom is rational to each other. Suppose we want to find moles of X atoms that reacted to form 0.25 mole of Y atoms. We always put the atom we want to find as numerator and the denominator being the atom we know.

E.g. $\qquad X + 2Z \rightarrow 2Y$

Find the ratio first:

$$\frac{\text{No of moles of } X}{\text{No of moles of } Y} = \frac{1}{2}$$

Then multiply the ratio by no. of moles of Y to find the reacting mole of X.

$$\frac{1}{2} \times 0.25 = 0.125 \text{ mole}$$

Therefore 0.125 mole of X reacted with 0.25 mole of Y. To find the reacting mass of X, e.g. Y is given as 35g, we just multiply the mole by the mass of Y as they are always in ratio:

$$0.125 \times 35 = 4.375 \text{ g}$$

Reacting Masses and Volumes

First, find the ratio of moles and multiply the mole of the gas volume you want to find with the volume of gas at room temperature ($24dm^3$)

Example

$MgCl_2$ is formed by reacting Mg and HC*l* according to equation:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(s) + H_2(g)$$
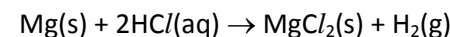
Find the amount of hydrogen gas, in cm^3, formed when 14.6g of HC*l* is reacted.

$$\text{Ratio:} \frac{\text{No.of moles of } H_2}{\text{No of moles of HC}l} = \frac{1}{2}$$

$$m(HCl) = \frac{14.6}{36.5}$$

$$= 0.4 \text{ mol}$$

Multiply ratio by mole of HC*l* $= \frac{1}{2} \times 0.4 = 0.2$ mol

Multiply mole by molar volume of gas at r.t.p. $= 0.2 \times 24 \text{ dm}^3 = 4.8 \text{ dm}^3$

$$1dm^3 = 1000cm^3$$

$$\therefore 4.8dm^3 \times 1000 = 4\,800 \text{ cm}^3$$

4 800cm^3 of gas is formed

CHEMICAL ANALYSIS

3.12 Introductory Chemical Analysis

Analysis is finding out what a substance or product is made of

Chemical analyst is the person who does chemical analysis

2 kinds of chemical analysis:

- **Qualitative analysis**

  is the identification of elements/compounds present in an unknown substance

- **Quantitative analysis**

  Is the meaurement of concentration of elements/compunds in unknown substance

3.13 Use of Physical Tests to Identify Substances

- *Colour* – some substances have distinctive colours.
  - Ammonium compounds and compounds in Groups I and II are white solids that dissolve in water to form colourless solutions
  - Copper(II) compounds are blue/green (except CuO is black)
  - Iron(II) compounds are pale green, iron (III) compounds are red or yellowish
  - Chlorine gas is greenish-yellow. Most other gases are colourless

- *Smell*
  - Gases like oxygen, hydrogen and carbon dioxide are odourless
  - Others like chlorine, ammonia and sulphur dioxide have characteristic smells

- *Solubility in Water*

  Some substances like $AgCl$ and $CaSO_4$ are insoluble while other does

- *pH*

  If a substance is pH 1 or 2, all alkaline and weakly acidic substances couldn't be the substance.

3.14 Volumetric Analysis

Is a measure of concentrations of acids/alkalis in solutions

Acid-alkali Titrations in Volumetric Analysis

It needs: - a *standard solution*: a solution of known concentration, and

- a solution of unknown concentration

Detecting the End Point

End point is the point at which neutralisation of acid and alkali is complete

- Sharp indicators (phenolphtalein and methyl orange) are used to detect end point effectively
- Litmus and universal indicators isn't used as the changes of end point isn't sharp

A Typical Acid-alkali Titration

The diagram shows how titration is used to find concentration of H_2SO_4 using NaOH

1. Draw up 25.0 cm³ of the sulfuric acid into a pipette.
2. Place the 25.0 cm³ of the sulfuric acid into a conical flask (called the titration flask).
3. Add a few drops of phenolphthalein to the sulfuric acid.
4. Fill the burette with the standard sodium hydroxide solution. Then run the alkali into the titration flask. The indicator changes colour from colourless to pink when all the sulfuric acid has reacted. Record the volume of sodium hydroxide used.

Using above example, to find the concentration of H_2SO_4 is given on the next page

Example:

30.0 cm^3 of 0.100 mol/dm^3 NaOH reacted completely with 25.0 cm^3 of H_2SO_4 in a titration. Calculate the concentration of H_2SO_4 in mo mol/dm^3 mol/dm^3 given that:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$
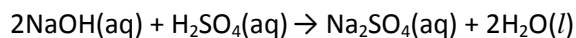
Step 1: Find the reacting mole of NaOH

$$n(NaOH) = \text{Concentration x Volume in mol/dm}^3$$

$$= 0.100 \times \frac{30.0}{1000} \text{ mol}$$

Step 2: Write the chemical equation for the reaction

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$
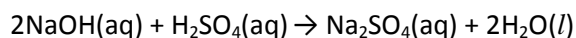
Step 3: Find the ratio of number of moles of H_2SO_4 to number of moles of NaOH

$$\frac{n(H_2SO_4)}{n(NaOH)} = \frac{1}{2}$$

Step 4: Use ratio to find number of moles of H_2SO_4 that reacted

$$n(H_2SO_4) = \frac{1}{2} \times \text{number of moles of NaOH}$$

$$= \frac{1}{2} \times 0.100 \times \frac{30.0}{1000} \text{ mol}$$

$$= 0.0015 \text{ mol}$$

Step 5: Find the concentration of H_2SO_4 in mol/dm^3

$$\text{Concentration} = \frac{n(H_2SO_4)}{\text{Volume of } H_2SO_4 \text{ in dm}^3}$$

$$= 0.0015 \text{ mol} \times \frac{1000}{25.0}$$

$$= 0.06 \text{ mol/dm}^3$$

Other Titrations

To find the concentration of a solution of $FeSO_4$ using $KMnO_4$ is as below

suck with a pipette filter

pipette

$FeSO_4(aq)$

H_2SO_4

1. Draw up 25.0 cm^3 of FeSO$_4$ into pipette.

2. Place the 25.0cm^3 of the FeSO$_4$ into conical flask.

3. Add approximately 20cm^3 of dilute sulphuric acid to the conical flask (to enable KMnO$_4$(aq) to react with FeSO$_4$)

4. Fill the burette with the standard KMnO$_4$(aq). Then run it into the conical flask. The KMnO$_4$(aq) is purple in colour and immediately decolorised by FeSO$_4$(aq). As KMnO$_4$ is run in, the purple colour dissapears until, at the end point, one drop of KMnO$_4$ causes a permanent pink colour. This completes the titration. Record the volume of KMnO$_4$(aq) used.

Example:

25.0 cm^3 of $FeSO_4$(aq), H_2SO_4 acidified, needs 27.5 cm^3 of 0.020 mol/dm^3 $KMnO_4$ for reaction in titration. Calculate the concentration of $FeSO_4$(aq)

Step 1: Find the reacting mole of $KMnO_4$

$$n(KMnO_4) = \text{Concentration x Volume in mol/dm}^3$$

$$= 0.020 \times \frac{27.5}{1000} \text{ mol}$$

Step 2: Write the chemical equation for the reaction

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

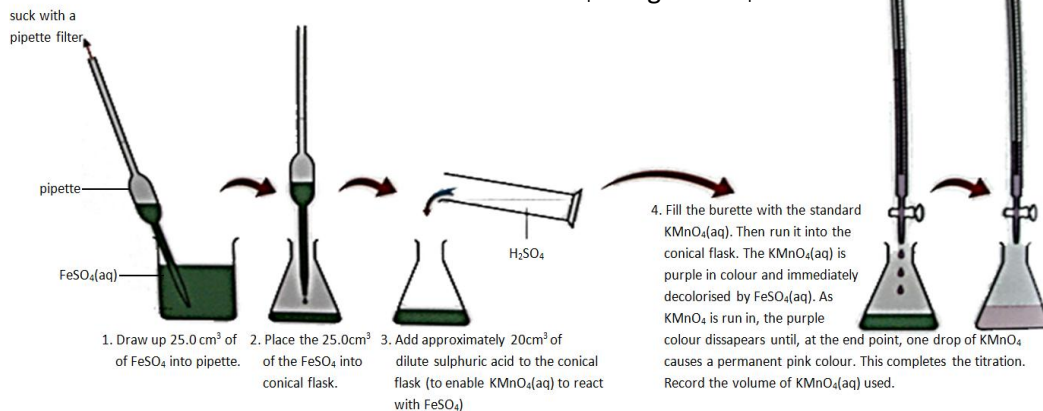**Step 3:** Find the ratio of number of moles of $FeSO_4$ to number of moles of $KMnO_4$

$$\frac{n(FeSO_4)}{n(KMnO_4)} = \frac{10}{2} = \frac{5}{1}$$

Step 4: Use ratio to find number of moles of $FeSO_4$ that reacted

$$n(FeSO_4) = 5 \times \text{number of moles of NaOH}$$

$$= 5 \times 0.020 \times \frac{27.5}{1000} \text{ mol}$$

$$= 0.00275 \text{ mol}$$

Step 5: Find the concentration of $FeSO_4$ in mol/dm^3

$$\text{Concentration} = \frac{n(FeSO_4)}{\text{Volume of } FeSO_4 \text{ in dm}^3}$$

$$= 0.00275 \text{ mol} \times \frac{1000}{25.0}$$

$$= 0.11 \text{ mol/dm}^3$$

3.15 Uses of Titrations in Analysis

Identification of Acids and Alkalis

Example:

An acid has formula of H_2XO_4. One mole of H_2XO_4 reacts with 2 moles of NaOH. A solution of the acid contain 5.0 g/dm^3 of H_2XO_4. In titration, 25.0cm^3 of acid reacted with 25.5cm^3 of 0.1 mol/dm^3 NaOH. Calculate the concentration of acid in mol/dm^3 and find X of the acid and its identity

$$n(NaOH) = \text{Concentration x Volume in mol/dm}^3$$

$$= 0.01 \times \frac{25.5}{1000} \text{ mol}$$

Ratio of H_2XO_4 to NaOH:

$$\frac{n(H_2XO_4)}{n(NaOH)} = \frac{1}{2}$$

$$n(H_2XO_4) = \frac{1}{2} \times 0.01 \times \frac{25.5}{1000}$$

$$\therefore \text{Concentration} = \frac{n(H_2XO_4)}{\text{Volume of } H_2XO_4 \text{ in } dm^3}$$

$$= \frac{1}{2} \times 0.01 \times \frac{25.5}{1000} \times \frac{1000}{25.0}$$

$$= 0.051 \text{ mol/dm}^3$$

Since 1 dm^3 of H_2XO_4 contains 0.051 mol and 5 g of H_2XO_4. $\therefore$ 0.051 mol of H_2XO_4 has a mass of 5g of H_2XO_4 and 1 mole of H_2XO_4 has a mass of $\frac{5.0 \text{ g}}{0.051 \text{ mol}}$ = 98 g

Hence, M_r of X = 98 − 2(1) − 4(16) = 32.

$\therefore$ X is sulphur and H_2XO_4 is sulphuric acid

Percentage Purity of Compounds

$$\text{Percentage purity} = \frac{\text{Mass of actual substance/}dm^3}{\text{Mass of impure substance/}dm^3}$$

Example:

5 g of impure sulphuric acid is dissolved in 1 dm^3 of water. 25.0 cm^3 of the solution required 23.5 cm^3 of 0.1 mol/dm^3 NaOH for complete titration reaction. Calculate the percentage purity of the acid.

$$n(NaOH) = \text{Concentration} \times \text{Volume in mol/dm}^3$$

$$= 0.1 \times \frac{23.5}{1000} \text{ mol}$$

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$
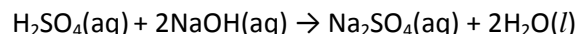
Ratio of H_2SO_4 to NaOH according to equation, then find mole according to ratio

$$\frac{n(H_2SO_4)}{n(NaOH)} = \frac{1}{2}$$

$$n(H_2SO_4) = \frac{1}{2} \times 0.1 \times \frac{23.5}{1000}$$

$$\therefore \text{Concentration} = \frac{n(H_2SO_4)}{\text{Volume of } H_2SO_4 \text{ in } dm^3}$$

$$= \frac{1}{2} \times 0.1 \times \frac{23.5}{1000} \times \frac{1000}{25.0}$$

$$= 0.047 \text{ mol/ dm}^3$$

Hence mass of H_2SO_4 in 1 dm^3 = 0.047 x $M_r(H_2SO_4)$

$$= 0.047 \times 98g$$

$$= 4.61g$$

Hence percentage purity = $\frac{4.61}{5.00} \times 100$ = 92.2%

Formulae of Compounds

Example:

Solution Y contains 30.0 g/dm^3 of $FeSO_4 \cdot xH_2O$. In a titration, 25.0 cm^3 of Y reacted with 27.0 cm^3 of 0.02 mol/dm^3 $KMnO_4$. In the reaction, 5 moles of $FeSO_4 \cdot xH_2O$ react with 1 mole $KMnO_4$. Calculate the concentration of Y in mol/dm^3 and the value of x.

$$n(KMnO_4) = \text{Concentration} \times \text{Volume in mol/dm}^3$$

$$= 0.02 \times \frac{27.0}{1000} \text{ mol}$$

Ratio of $FeSO_4 \cdot xH_2O$ to $KMnO_4$ according to question; find mole according to ratio

$$\frac{n(FeSO_4 \cdot xH_2O)}{n(KMnO_4)} = \frac{5}{1}$$

$$n(FeSO_4 \cdot xH_2O) = 5 \times 0.02 \times \frac{27.0}{1000}$$

$$\therefore \text{Concentration} = \frac{n(FeSO_4 \cdot xH_2O)}{\text{Volume of } FeSO_4 \cdot xH_2O \text{ in } dm^3}$$

$$= 5 \times 0.02 \times \frac{27.0}{1000} \times \frac{1000}{25.0}$$

$$= 0.108 \text{ mol/dm}^3$$

Hence 0.108 mol $FeSO_4 \cdot xH_2O$ = 30.0 g $FeSO_4 \cdot xH_2O$

Therefore 1 mole $FeSO_4 \cdot xH_2O$ has a mass of $\frac{30.0}{0.108}$ g = 278 g

Therefore $M_r(FeSO_4 \cdot xH_2O)$ = 278, Hence $x = \frac{278 - 152}{18}$ = 7

Number of Reacting Moles in an Equation

Example:

In a titration, 25.0 cm^3 of 0.04 mol/dm^3 H_2O_2 reacted with 20.0 cm^3 of 0.02 mol/dm^3 $KMnO_4$. Find the values of x and y given the equation:

$$xH_2O_2 + yKMnO_4 + \text{acid} \rightarrow \text{products}$$
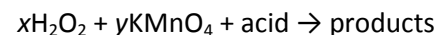
$$n(H_2O_2) = \text{Concentration} \times \text{Volume in mol/dm}^3$$

$$= 0.04 \times \frac{25.0}{1000} \text{ mol} = 0.001 \text{ mol}$$

$$n(KMnO_4) = \text{Concentration} \times \text{Volume in mol/dm}^3$$

$$= 0.02 \times \frac{20.0}{1000} \text{ mol} = 0.0004 \text{ mol}$$

Therefore 1 mole $KMnO_4$ react with $\frac{0.001}{0.0004}$ = 2.5 moles of H_2O_2

Hence ratio of x:y is 2.5:1 = 5:2 (round off)

Therefore, x=5 and y=2

END OF CHAPTER 3

CHAPTER 4 – ELECTROLYSIS

4.1 Introductory Electrolysis

Electrolysis is the decomposition of compound using electricity
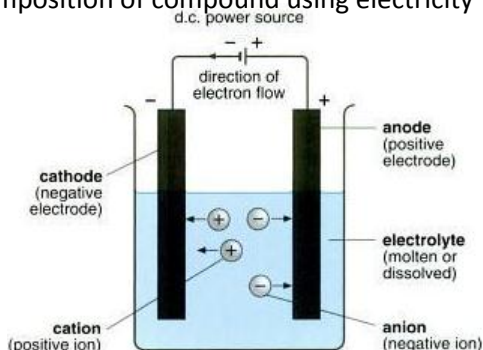
Electrolyte is an ionic compound which conducts electric current in molten or aqueous solution, being decomposed in the process.

Electrode is a rod or plate where electricity enters or leaves electrolyte during electrolysis. Reactions occur at electrodes.

Discharge is the removal of elctrons from negative ions to form atoms or the gain of electrons of positive ions to become atoms.

Anode is positive electrode connected to positive terminal of d.c. source. Oxidation occurs here. **Anode** loses negative charge as electrons flow towards the battery, leaving anode positively charged. This causes anion to discharge its electrons here to replace lost electrons and also, negativecharge are attracted to positive charge.

Cathode is negative electrode connected to negative terminal of d.c. source. Reduction occurs here. **Cathode** gains negative charge as electrons flow from the battery towards the cathode, making cathode negatively charged. This causes cation to be attracted and gains electrons to be an atom.

Anion is negative ion. It's attracted to anode.

Cation is positive ion. It's attracted to cathode.

4.2 Electrolysis of Molten Compounds

Molten/aqueous ionic compounds conduct electricity because **ions free to move**. In solid state, these ions are **held in fixed position** within the crystal lattice. Hence solid ionic compounds *do not conduct electricity*.

When **molten binary compound** is electrolysed, **metal** is formed on **cathode** while **non-metal** is formed on **anode**.

Electrolysis of Molten $PbBr_2$

To make molten lead(II) bromide, $PbBr_2$, we strongly heat the solid until it melts. To electrolyse it, pass current through the molten $PbBr_2$.

Ions Present

Pb^{2+} and Br^-

Reaction at Anode

Br^- loses electrons at anode to become Br atoms. Br atoms created form bond together to make Br_2 gas.

$$2Br^-(aq) \rightarrow Br_2(g) + 2e^-$$
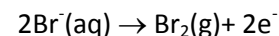
Reaction at Cathode

Pb^{2+} gains electrons at cathode to become Pb atoms becoming liquid lead (II).

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

Overall Equation

$$PbBr_2(l) \rightarrow Pb(l) + Br_2(g)$$
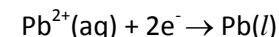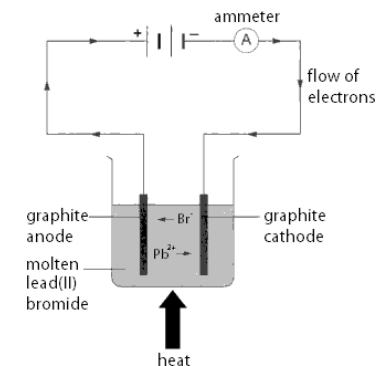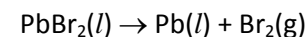
Below are other compounds that can be electrolysed. The theory's same as $PbBr_2$.

Molten electrolyte	Cathode product	Anode product
Calcium chloride ($CaCl_2$)	Calcium, Ca	Chlorine, Cl_2
Sodium chloride ($NaCl$)	Sodium, Na	Chlorine, Cl_2
Aluminium(III) oxide (Al_2O_3)	Aluminium, Al	Oxygen, O_2
Sodium Iodide (NaI)	Sodium, Na	Iodine, I_2

4.3 Electrolysis of Aqueous Solution

Aqueous solutions contain additional H^+ and OH^- ions of water, totalling 4 ions in the solution – 2 from electrolyte, 2 from water. Only 2 of these are discharged.

Electrolysis of aqueous solutions use the theory of **selective discharge**.

Ease of discharge	Cation	Product at cathode	Anion	Product at anode
Difficult ↓ Easy	K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Ni^{2+} Pb^{2+} H^+	} Hydrogen } from } water } • Nickel • Lead Hydrogen	SO_4^{2-} NO_3^- Cl^- Br^- I^- OH^-	} Oxygen from } water Chlorine Bromine Iodine Oxygen
	Cu^{2+} Ag^+	Copper Silver		

At cathode

- In CONCENTRATED solutions of nickel/lead compound, nickel/lead will be discharged instead of hydrogen ions of water which is less reactive than nickel/lead.
- In VERY DILUTE solutions, hydrogen, copper and silver ions are preferrable to be discharged, according to its ease to be discharged.
- Reactive ions (potassium, sodium, calcium, magnesium, aluminium) will NEVER BE DISCHARGED in either concentrated or dilute condition. Instead, hydrogen ions from water will be discharged at cathode.

At anode

- In CONCENTRATED solutions, iodine/chlorine/bromine ions are preferrable to be discharged, although it's harder to discharged compared to hydroxide ions.
- In VERY DILUTE solutions containing iodide/chloride/bromide ions, hydroxide ions of water will be discharged instead of iodide/chloride/bromide, according to ease of discharge.
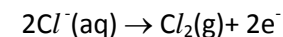- Sulphate and nitrate are NEVER DISCHARGED in concentrated/dilute solutions.

Concentrated Solutions

Electrolysis of Concentrated NaCl

Ions Present

Na^+, H^+, OH^- and Cl^-

Reaction at Anode

Cl^- loses electrons at anode to become Cl atoms, although OH^- is easier to discharge. Cl atoms created form covalent bond together to make Cl_2 gas.

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

Reaction at Cathode

H^+ gains electrons at cathode to become H atoms becoming hydrogen gas.

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

Overall Equation

$$2HCl(l) \rightarrow H_2(l) + Cl_2(g)$$
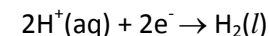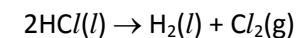
Note: any cation and anion left undischarged in solution forms new bonds between them. E.g. in above, leftovers Na^+ and OH^- combine to form NaOH.

Very Dilute Solutions

Electrolysis of Dilute H_2SO_4

Ions Present

H^+, OH^- and SO_4^{2-}

Reaction at Anode
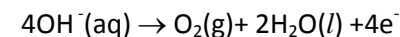
OH^- loses electrons at anode to become O_2 and H_2O.

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

Reaction at Cathode

H^+ gains electrons at cathode to become H atoms becoming hydrogen gas.

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

Overall Equation

Both equations must be balanced first. The cathode equation is short 2 electrons. Hence, we should first even them by multiplying cathode equation by 2.

$$(2H^+(aq) + 2e^- \rightarrow H_2(g)) \times 2$$
$$= 4H^+(aq) + 4e^- \rightarrow 2H_2(g)$$

Now we can combine the equations, forming:

$$4H^+(aq) + 4OH^+(aq) \rightarrow 2H_2(g) + O_2(g) + 2H_2O(l)$$
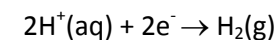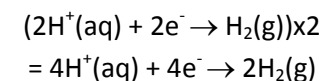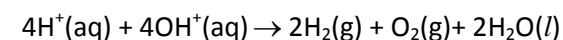
$4H^+$ and $4OH^+$ ions, however, combine to form $4H_2O$ molecules. Hence:

$$4H_2O(l) \rightarrow 2H_2(g) + O_2(g) + 2H_2O(l)$$
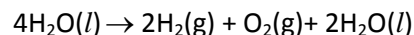
H_2O molecules are formed on both sides. Therefore, they cancel the coefficients:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$
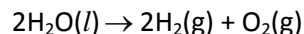
Since only water is electrolysed, the sulfuric acid now only becomes concentrated.

4.4 Electrolysis Using Different Types of Electrodes

Inert Electrodes are electrodes which do not react with electrolyte or products during electrolysis. Examples are platinum and graphite.

Active Electrodes are electrodes which react with products of electrolysis, affecting the course of electrolysis. Example is copper.

Electrolysis of $CuSO_4$ Using Inert Electrodes(e.g. carbon)

Ions Present

Cu^{2+}, H^+, OH^- and SO_4^{2-}

Reaction at Anode

OH^- loses electrons at anode to become O_2 and H_2O.

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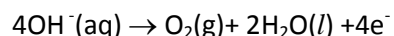
Reaction at Cathode

Cu^{2+} gains electrons at cathode to become Cu atoms becoming liquid copper. Hydrogen ions are not discharged because copper is easier to discharge.
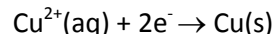
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

Overall Equation

Both equations must be balanced first. The cathode equation is short 2 electrons. Hence, we should first even them by multiplying cathode equation by 2.

$$(Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)) \times 2$$
$$= 2Cu^{2+}(aq) + 4e^- \rightarrow 2Cu(s)$$

Now we can combine the equations, forming:

$$2Cu(OH)_2(aq) \rightarrow 2Cu(s) + O_2(g) + 2H_2O(l)$$
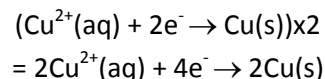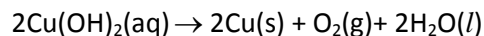
Since copper ions in solution are used up, the blue colour fades. Hydrogen and sulphate ions left forms sulphuric acid.

Electrolysis of $CuSO_4$ Using Active Electrodes(e.g. copper)

Ions Present

Cu^{2+}, H^+, OH^- and SO_4^{2-}

Reaction at Anode

Both SO_4^{2-} and OH^- gets attracted here but not discharged. Instead, the copper anode discharged by losing electrons to form Cu^{2+}. So, the electrode size decreases.

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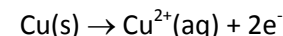
Reaction at Cathode

Cu^{2+} produced from anode gains electrons at cathode to become Cu atoms becoming copper. Hence, the copper is deposited here and the electrode grows.

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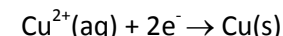
Overall Change

There is no change in solution contents as for every lost of Cu^{2+} ions at cathode is replaced by Cu^{2+} ions released by dissolving anode. Only the cathode increases size by gaining copper and anode decreases size by losing copper. We can use this method to create pure copper on cathode by using pure copper on cathode and impure copper on anode. Impurities of anode falls under it.
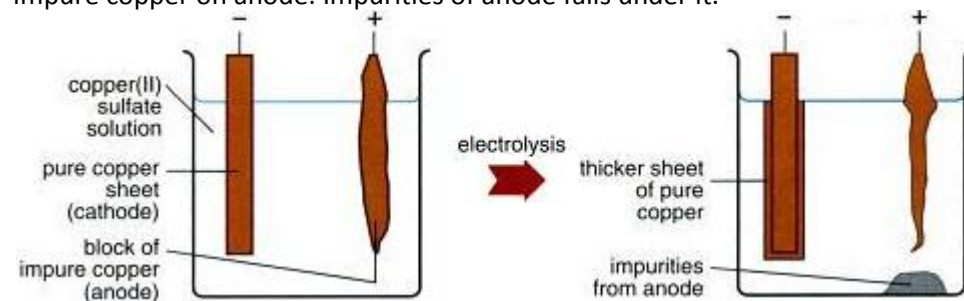
4.5 Electroplating

Electroplating is coating an object with thin layer of metal by electrolysis. This makes the object protected and more attractive.

Object to be plated is made to be cathode and the plating metal is made as anode. The electrolyte MUST contain plating metal cation.

Plating Iron object with Nickel

Reaction at Anode

Ni^{2+} discharged from anode into solution. So, the electrode size decreases.

$$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$$
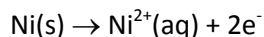
Reaction at Cathode

Ni^{2+} produced from anode gains electrons at cathode to become Ni atoms becoming nickel. Hence, the nickel is deposited here and the electrode grows.

$$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$$
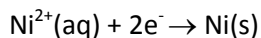
Overall Change

There is no change in solution contents while iron object receives nickel deposit.

Uses of Electroplating

Plating Metal	Uses
Chromium	Water taps, motorcar bumpers, bicycle parts
Tin	Tin cans
Silver	Silver sports trophies, plaques, ornaments, cutleries
Nickel	For corrosion-resistant layer
Gold	Watches, plaques, cutleries, water taps, ornaments
Rhodium	Silverware, jewellery, watches, ornaments
Copper	Printed circuit boards, trophies, ornaments

4.6 Creation of Electric Cells by Electrolysis

An electric cell consists of 2 electrodes made of 2 metals of different reactivity. The cathode is made of **more reactive metal**. This is because they have more tendency of losing electrons. The anode is made of **less reactive metal**. The more further apart the metals in reactivity series, the higher voltage is created.
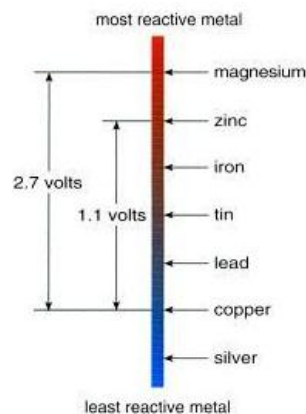
How the Voltage is Produced

Use an example of zinc and copper as electrodes and sodium chloride as electrolyte
As zinc is more reactive, it is cathode while copper is anode.

At cathode, Zn atoms in anode loses electrons to form Zn^{2+}

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$
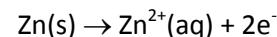
Zn^{2+} goes into solution while the electrons lost makes the zinc negative. The electrons flow against conventional current towards copper anode. Both Na^+ & H^+ ions in solution are attracted to the copper anode due to electrons in it but only H^+ ions discharged, due to selective discharge, to form hydrogen gas.

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$
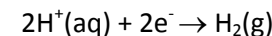
Hence the overall ionic equation is:

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$
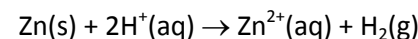
which comes from the equation:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$
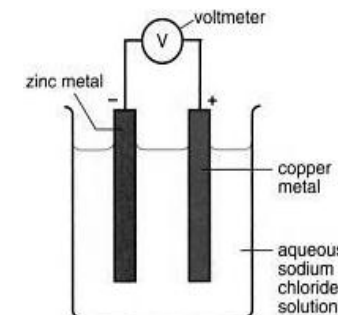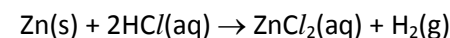
END OF CHAPTER 4

CHAPTER 5 – ENERGY FROM CHEMICALS

5.1 Exothermic Reaction

Exothermic change is one which heat energy is given out. This is to form bonds between the reactants which needs less energy in them.

Reaction is written as:

$$Reactants \rightarrow Products + heat$$

(or)
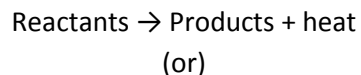
$$Reactants \rightarrow Products [\Delta H = -n \text{ kJ}]$$, where n is amount of heat energy released

Examples of exothermic changes:

1. **Changes of State**

   When gas condenses to water or water freezes to solid, heat is given out.

   Condensation of steam to water

   $$H_2O(g) \rightarrow H_2O(l) + heat$$
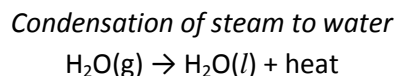
2. **Combustion reactions**

   All combustion (burning) reactions are exothermic.

   Burning of hydrogen in air

   $$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) + heat$$
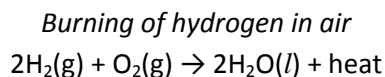
3. **Dissolving of anhydrous salts/acids in water**

   Dissolving solid salt to aqueous solution of the salt gives out heat

   Dissolving of Na_2CO_3 in water (or $CuSO_4$)

   $$Na_2CO_3(s) \rightarrow Na_2CO_3(l) + heat$$

   Dissolving of concentrated acid in water

   $$HCl(aq) + H_2O(l) \rightarrow \text{less concentrated } HCl(aq) + heat$$
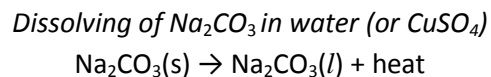
4. **Neutralization**

   When acid and alkali react it gives out heat due to combining of H^+ ions from acid and OH^- ions from alkali to form water

   $$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) + heat$$
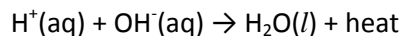
5. **Metal Displacement**

   Magnesium reacting with copper(II) sulphate

   $$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(s) + Cu(s) + heat$$

5.2 Endothermic Reaction

Endothermic change is one which heat energy is absorbed. This is to break bonds between the reactants which needs more energy in them.

Reaction is written as:

$$Reactants + heat \rightarrow Products$$

(or)
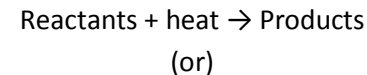
$$Reactants \rightarrow Products [\Delta H = +n \text{ kJ}]$$, where n is amount of heat energy absorbed

Examples of endothermic changes:

1. **Changes of states**

   When solid melts to water & boils to steam, heat is absorbed to break the bond.

   Condensation of steam to water

   $$H_2O(s) + heat \rightarrow H_2O(l)$$
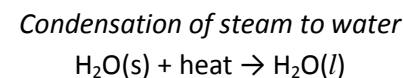
2. **Photolysis**

   Reaction of light sensitive silver chloride in camera reel in light

   $$2AgBr(s) + heat \rightarrow 2Ag(s) + Br_2(g)$$
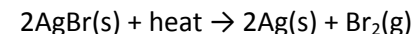
3. **Dissolving of Ionic Compounds**

   Ionic compounds such as NH_4Cl, KNO_3, $CaCO_3$ absorb heat from surroundings.

   $$NH_4Cl(s) + heat \rightarrow NH_4Cl(aq)$$

   $$CuSO_4(s) + heat \rightarrow CuSO_4(aq)$$
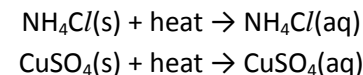
4. **Photosynthesis**

   Light energy is absorbed by plants to produce starch.

5. **Decomposition by heat**

   Many compounds require heat for decomposition, e.g. $CaCO_3$ to CO_2 + CaO

   $$CaCO_3(s) + heat \rightarrow CO_2(g) + CaO(s)$$
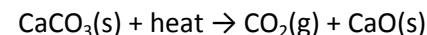
6. **Acid + Bicarbonates (HCO_3)**

   $$NaHCO_3(s) + H_2SO_4(aq) + heat \rightarrow NaSO_4(aq) + CO_2(g) + H_2O(l)$$
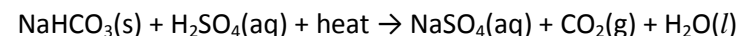
Acid Spill Treatment on Body

We don't neutralize spilled acid on body as it produces heat. Instead we dilute the solution with water, although it also produces heat, but is less than neutralizing it.

5.3 Heat of Reaction

The amount of energy given out or absorbed during a chemical reaction is **enthalpy change**. The symbol is **ΔH** measured in **kilojoules(kJ)**.

Examples of reactions from back page:

Exothermic reaction:

$$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s) \text{ [ΔH = −378 kJ]}$$
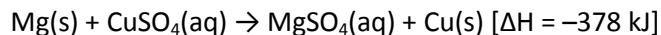
378 kJ of heat energy is given out when 1 mol of Mg react with 1 mol $CuSO_4$ to produce 1 mol of $MgSO_4$ and 1 mol of Cu.

Endothermic reaction:

$$CaCO_3(s) \rightarrow CO_2(g) + CaO(s) \text{ [ΔH = +222 kJ]}$$
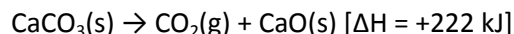
222 kJ of heat energy is absorbed when 1 mol of $CaCO_3$ decompose to 1 mol of CO_2 and 1 mol of CaO.

5.4 Heat Energy and Enthalpy Change in Reaction

When bonds made, heat energy is given out, it's exothermic and ΔH is negative

When bonds broken, heat energy is absorbed, it's endothermic and ΔH is positive

Question:

Hydrogen bromide is made by reacting H_2 gas with Br_2 gas. Calculate the heat change of the reaction given the equation and bond energy table below.

$$H_2(g) + Br_2(g) \rightarrow 2HBr(g)$$
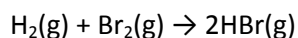
Covalent Bond	Bond energy (kJ/mol)
H − H	436
Br − Br	224
H − Br	366

Bonds of H_2 and Br_2 molecules must be broken first to make HBr. Heat energy is absorbed to break these bonds by endothermic reaction.

$$H − H + Br − Br \rightarrow H \ H \ Br \ Br$$
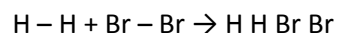
Broken bonds are used to make H − Br bonds of HBr. Heat energy is released.

$$H \ H \ Br \ Br \rightarrow 2H − Br$$
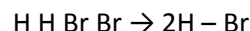
Heat change can be calculated by:

ΔH = heat released in making bonds + heat absorbed in breaking bonds

Exothermic ΔH = the bond energy of 2 H − Br bonds

$$= 2(366)$$

$$= − 732 \text{ kJ}$$

Endothermic ΔH = the bond energy of 1 H − H bond + 1 Br − Br bond

$$= 436 + 224$$

$$= + 660 \text{ kJ}$$

$$ΔH = − 732 + 660$$

$$= − 72 \text{ kJ}$$

Therefore more heat is given out in making bond than absorbed in breaking bond. The overall change is to give out heat and it's exothermic with ΔH negative.

Exothermic graph:

When heat is given out, the solution becomes warm and later the temperature goes back to room temperature.
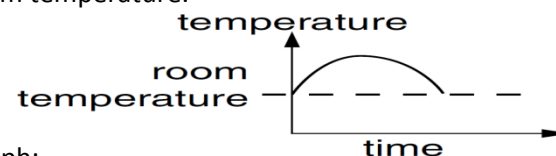
Endothermic graph:

When heat is absorbed from the surrounding of reactant, the solution becomes cooler and later the temperature goes back to room temperature.
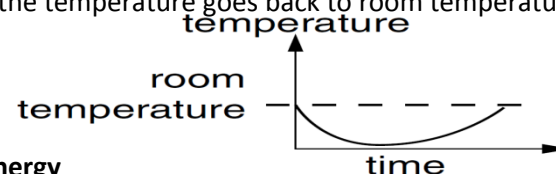
5.5 Activation Energy

Activation energy is the minimum energy needed to start a reaction.

It is the energy needed to break the reactant bonds before new bonds are formed.
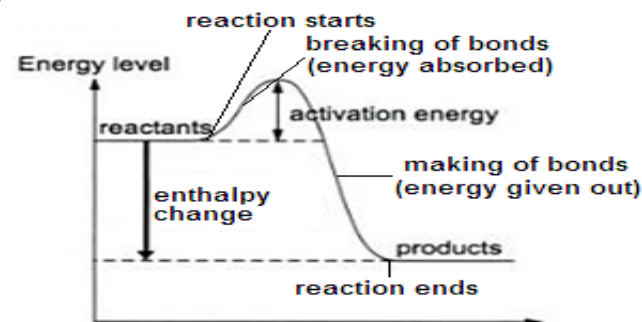
Reactions occur because of collision of particles and sufficient kinetic energy is needed to provide activation energy to break the bonds and start the reaction by providing extra energy from a heat source.

Exothermic and Endothermic Reaction Graph

In exothermic reaction, enough energy given out in the reaction of parcticles to provide activation energy therefore less energy is needed to form products.

In endothermic reaction, insufficient energy is given out when bonds are made to provide activation energy for reaction to continue. More energy is needed to form products and heat must be continually added to fulfill energy requirement.

EXOTHERMIC REACTION   ENDOTHERMIC REACTION
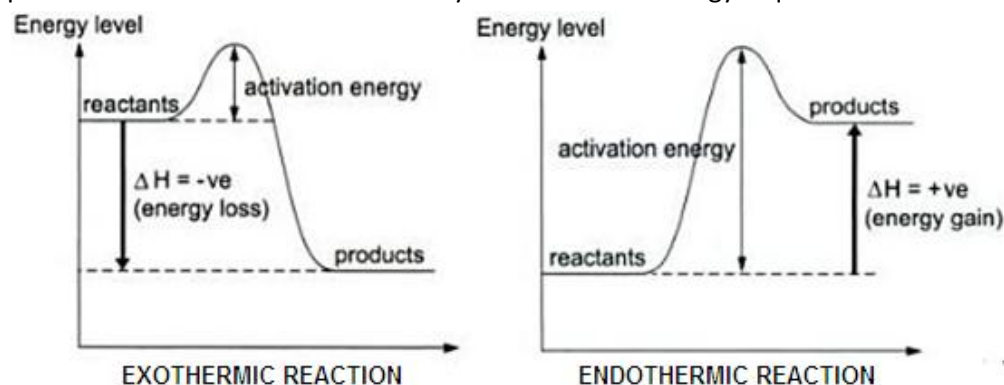
5.6 FUELS

The combustion of fuels gives out large amount of energy in industries, transport & homes. These fuel mainly methane from coal, wood, oil, natural gas & hydrogen. Combustion in air provides energy and gives out heat. Hence, exothermic reaction.

Hydrogen as a Fuel

Hydrogen provides twice as much as heat energy per gram than any other fuel and burns cleanly in air to form steam. They are mainly used as rocket fuel.

Production of Hydrogen

Hydrogen is produced either by electrolysis of water or by cracking of hydrocarbon

By cracking of hydrocarbon:

First, methane (hydrocarbon) and steam are passed over a nickel catalyst to form hydrogen and carbon monoxide.

$$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$$
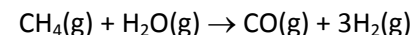
The by-product carbon monoxide is not wasted. It is reacted with more steam to form carbon dioxide and hydrogen.

$$CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$$
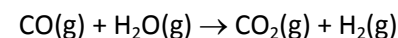
Now you get more hydrogen.

By electrolysis:

Water is electrolysed according to equation:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$
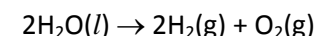
However, electrolysis is costly.

Creation of the Fuel

In Engines:

The hydrogen created is reacted with oxygen to form steam and heat energy

$$2H_2(g) + O_2(g) \rightarrow H_2O(g) + heat$$
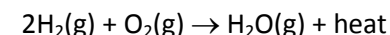
This heat is needed to thrust the vehicle forward. However, we don't use heat energy for our daily appliances. Instead we use electrical energy and to make electrical energy from hydrogen, we use fuel cell.

A **fuel cell** converts chemical energy directly into electrical energy.

How Fuel Cells Work
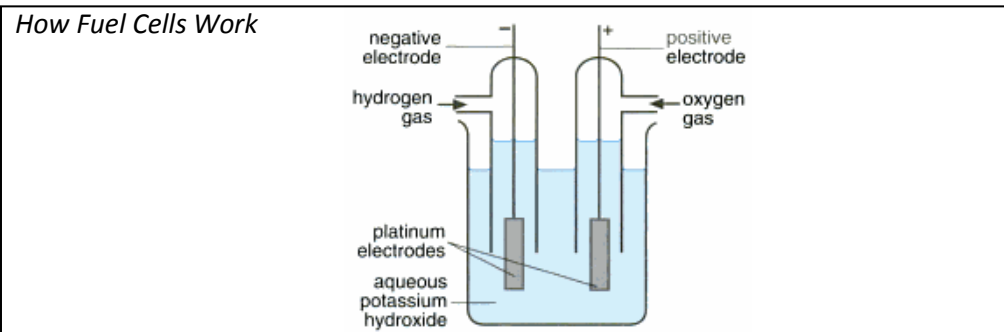
Hydrogen reacts with hydroxide ions into electrolyte on the platinum catalyst on electrode to make the electrode negatively-charged.

$$H_2 + 2OH^- \rightarrow 2H_2O + 2e^-$$
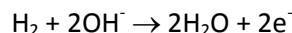
Electrons flows past the load and to the other electrode. That negatively-charged electrode is now anode. Hydroxide ions constantly deposit electrons here to make water. While then, the other electrode is now cathode.

Oxygen reacts with water created on from hydrogen on the cathode to gain electrons from it:

$$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$$
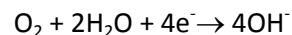
If we combine the ionic equations, we still get water as product of hydrogen and oxygen, but the energy produced is now electrical energy:

$$2H_2(g) + O_2(g) \rightarrow H_2O(g) + \text{electrical energy}$$
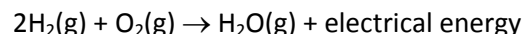
Advantages of Fuel Cells

- Electrical energy can be generated continuously if there's continuous fuel supply
- The by-product of fuel cells is steam, which do not pollute the environment
- Chemical energy is efficiently converted to electrical energy. Hence there is minimal loss of energy.

Disadvantages of Fuel Cells

- Hydrogen-oxygen fuel cells are very expensive, hence limiting their use.

Our Main Fuel Resource – PETROLEUM

Petroleum is a mixture of hydrocarbons, which are compounds made up of carbon and hydrogen only.

Crude oil, freshly extracted from underground, undergo **refining** – a process where oil undergoes **fractional distillation** to be separated into its fractions.

First, crude oil is heated up to 350^oC and the vapours rise up a tower, divided with trays on some certain heights for the fractions to be collected. The fractionating column is cooler on top, hence upper trays collects fractions of low boiling points while the lower ones, being hotter, collect those with higher boiling points.

Fraction	Approximate boiling	Remarks and uses
Gas fraction	Below 40 °C $C_1 - C_7$	Further distilled and treated chemically to produce • gaseous heating fuel, • unsaturated hydrocarbons for plastic manufacture, and • solvents (e.g. petroleum ether).
Gasoline	30 °C – 180°C $C_6 - C_{12}$	Treated to a process of reforming. Additives added. Used as petrol for cars.
Naphtha	110 °C – 195 °C $C_5 - C_{16}$	Used as a feedstock for the chemical industry and for producing high octane gasoline via catalytic reforming.
Kerosene (or paraffin)	150 °C – 240 °C $C_{11} - C_{16}$	Treated to remove unsaturated compounds. Used as jet fuel, in heating and lighting (e.g. oil stoves). Some is cracked to form petrol.
Gas oil (diesel)	220 °C – 250 °C $C_{15} - C_{25}$	Used as fuel for diesel engines or cracked to form petrol.
Heavy oil	250 °C – 350 °C $C_{20} - C_{70}$	Used directly as fuel for furnaces and ships. Distilled under a vacuum to produce • lubricating oils, • waxes, polishes, greases, etc.
Residues	Above 350 °C $> C_{70}$	Used to make bitumen (for roads) pitch and waxes.

Figure 5.13 Fractions and their uses

PHOTOSYNTHESIS AND ENERGY

Plants take in carbon dioxide and water in presence of chlorophyll and synthesize them in the presence of sunlight to produce glucose and release oxygen:

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$
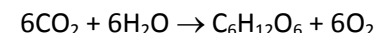
Plants get their energy by using the glucose formed. Scientists believe that we can use the stored energy in glucose as combustible fuels.

First, glucose fermented to make ethanol by microorganisms such as yeast. This is *fermentation*. The glucose is usually derived from corn plant or sugar cane.

$$C_6H_{12}O_6 \rightarrow 2C_2H_6O + 2CO_2$$

Then, water is removed from ethanol by fractional distillation by heating it up until 78^oC (boiling point of ethanol). Some water might still be present as the boiling point is close to ethanol. The ethanol produced is then mixed with fuel to be combusted to produce energy. This is *biofuel*, and it's a renewable energy source.

END OF CHAPTER 5

CHAPTER 6 – CHEMICAL REACTIONS

6.1 Speed of Reaction

Measuring Speed of Reaction

It is the speed for a reactant to be used up or product to be formed.

2 ways to find out speed of reaction

1. Measuring time for reaction to complete

Speed of reaction is *inversely proportional* to time taken; the shorter the time needed for reaction to complete, the faster the speed of reaction is.
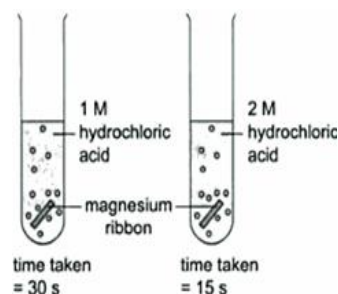
Speed of reaction = $\frac{1}{\text{time taken}}$

Speed of reaction A = $\frac{1}{30 \text{ s}}$ = 0.333/s

Speed of reaction B = $\frac{1}{15 \text{ s}}$ = 0.667/s

Therefore reaction B is faster than reaction A as time taken for B is shorter

Number of times B faster than A = $\frac{0.667}{0.333}$ = 2 times

2. Measuring the amount of product produced in a period of time or measuring the amount of reactant remain in a period of time.

Can be measured by plotting change in volume of gas evolved, mass of reaction mixture as reaction proceeds and change of pressure of gas formed.

- *Measuring the amount of gas evolved.*

Consider reaction of limestone with acid to produce carbon dioxide. A syringe is used to help in measurement of gas produced in volume every time interval. A graph of volume of gas against time is plotted.
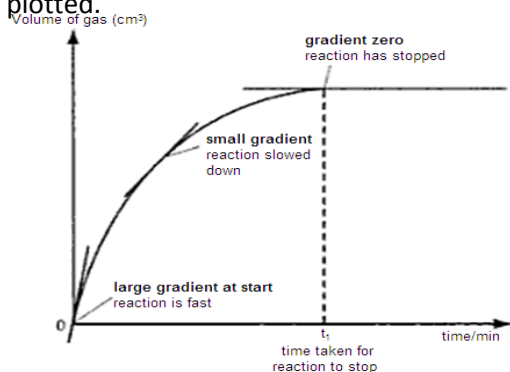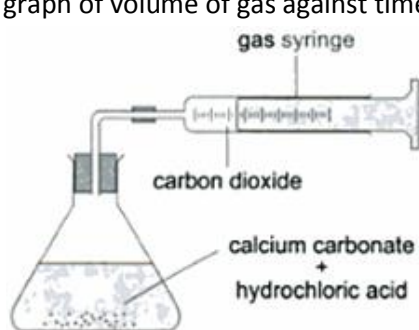
- Gradient largest at start indicating speed at its greatest.
- Gradient decreases with time – speed decreases with time.
- Gradient becomes zero, speed is zero. The reaction has finished.

- *Measuring change in mass of reaction mixture.*
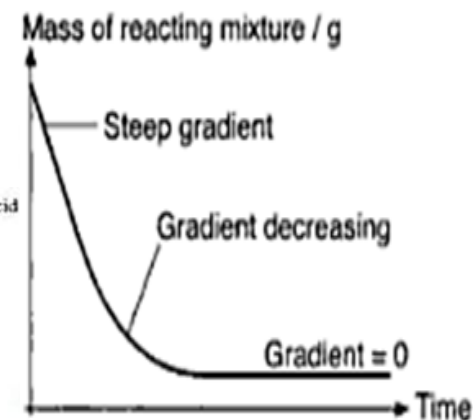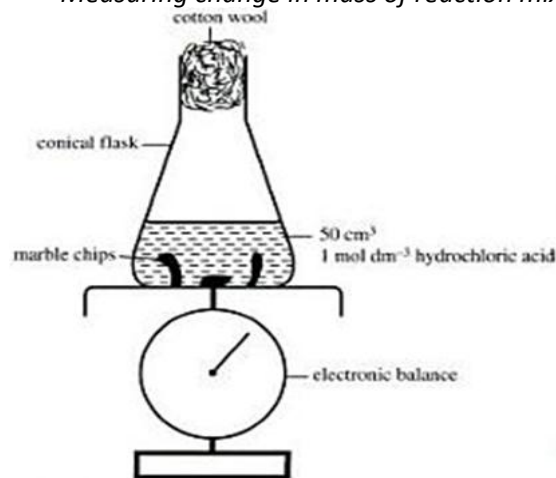
Marble is reacted with acid in a flask with cotton wool stucked at top to prevent splashing during reaction but it allows gas to be free. The reading on balance is plotted on a graph on every time interval.

Factors Affecting Speed of Reaction

1. **Particle Size of Reactant**

When large marble is reacted with acid and compared to reaction of fine marble solids being reacted with acid and the graph of volume of gas against time is plotted, it's found that the reaction involving finer marble chips produces gas faster than the one with larger marble chunk as the graph of finer chips is steeper. The volume of gas at the end is the same for both reactions. Therefore, reactions of solids with liquid/gas is faster when the solids are of smaller pieces.
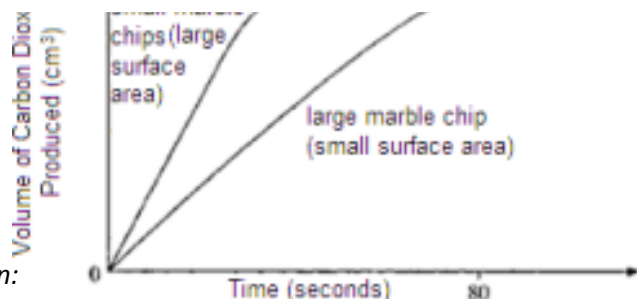
Explanation:

Reactions occur when particles collide. Small particles creates larger surface area for more collisions between reacting particles which increases speed of reaction.

Explosions: chemical reactions occuring extremely rapid rate producing heat+gas

- Coal dust burn faster than large pieces as it has larger surface area. In coal mines, when air contains too much coal dust, explosion can occur from a single spark or match. Water is sprayed into the air to remove coal dust.
- Flour in mills can ignite easily due to large surface area.

2. Concentration of Reactant

In the increase of concentration means there are more solute particles per unit volume of the solution which favours for more effective collision resulting in an increase in speed of reaction.

3. Pressure of Reactant

LOW PRESSURE   HIGH PRESSURE

Only gaseous reactions are affected as gas is compressible. At higher pressure, molecules are forced to move closely together, hence increasing the particles per unit volume of gas and effectively increases the collision between reacting molecules so the speed of reaction increases.

High pressure is used in industrial processes (e.g. Haber Process Plant) so that the reaction goes faster.

4. Temperature of Reaction

Speed of reaction increases when temperature increases. Particles don't always react upon collision but just bounce as they don't have enough activation energy to react. With increase in temperature, particles absorb the energy and having enough activation energy, they move faster and collide more effectively per second. Therefore, speed of reaction is increased.

Usually, speed of reaction doubles for every $10^{o}C$ rise in temperature.

5. Effect of Catalyst

What are catalysts?

They are chemical substances which alters speed of reaction without itself being used at the end of a reaction. It can be reused and only small amount of catalyst is needed to affect a reaction.

- transition metals (e.g. Titanium, Nickel, Iron, Copper) are good catalysts
- most catalyst catalyse one kind of reaction (except titanium)

Reaction	Catalyst
Production of sulphur by contact process	Vanadium(V) oxide, V_2O_5
Production of ammonia by Haber Process	Iron, Fe
Production of hydrogen by cracking of hydrocarbons	Aluminium oxide, Al_2O_3 Silicon dioxide, SiO_2
Production of margarine by reacting hydrogen with vegetable oil	Nickel, Ni
Production of plastics	Titanium(IV) chloride, $TiCl_4$
Converting CO into CO_2 in catalytic converters	Titanium, Ti Rhodium, Rh

Catalysts lower the need of energy to break bonds so activation energy is lower. Consequently, bond breaking occurs easily and more often when particles collide

Factors Affecting Speed of Catalysed Reactions:

Speed of catalysed reactions can be increased by:

- increasing temperature
- increasing concentration of solutions
- increasing pressure of gas reactions

Catalyst provide "alternative path" which results in lower activation energy.
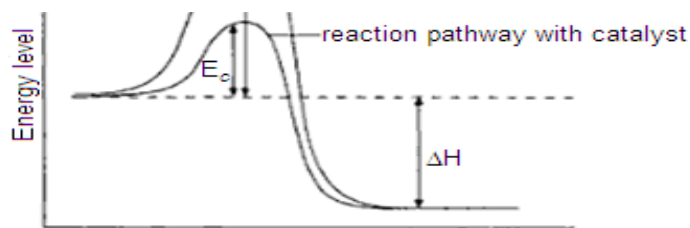
E_c is activation energy with catalyst
E_w is activation energy without catalyst

Enzymes

Enzymes are biological catalysts

Characteristics of enzymes:
- They are very specific. One enzyme catalyse one type of reaction.
- Enzymes are sensitive to temperature. They work best at 40^oC. Too high or too low temperatures destroy enzymes.
- Enzymes are sensitive to pH. They function within narrow range of pH.

Industrial uses of enzymes:
- They are added to detergents from bacteria, and also to make tough meat tender. These enzymes can be found in papaya fruit.
- Yeast convert sugars into alcohol and carbon dioxide by fermentation. Beer, wine and soy sauce are made this way.
- Fungical enzymes can be used to make antibiotics such as penicillin.

6.2 Redox

Oxygen in Reduction-Oxidation reaction

Oxidation is the *gain of oxygen* by a substance
Reduction is the *loss of oxygen* by a substance

$$Pb(s) + Ag_2O \ (aq) \rightarrow PbO(aq) + 2Ag \ (aq)$$
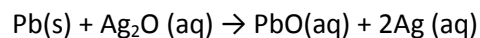
Pb is oxidized as it gains oxygen from Ag_2O to form PbO. Ag_2O is oxidizing agent.
Ag_2O is reduced as it loses oxygen to Pb to form Ag. Pb is reducing agent.

Oxidizing agent is a substance which causes *oxidation of another substance*
Reducing agent is a substance which causes *reduction of another substance*

Hydrogen in Reduction-Oxidation reaction

Oxidation is the *loss of hydrogen* by a substance
Reduction is the *gain of hydrogen* by a substance

$$H_2S(g) + Cl_2(g) \rightarrow 2HCl(g) + S(g)$$
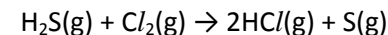
H_2S is oxidized as it loses hydrogen to Cl_2 to form S. Cl_2 is oxidizing agent.
Cl_2 is reduced as it gains hydrogen from H_2S to form HCl. H_2S is reducing agent.

Electrons in Reduction-Oxidation reaction

Oxidation is the *loss of electrons* by a substance
Reduction is the *gain of electrons* by a substance

$$2Na(s) + Cl_2(g) \rightarrow 2Na^+Cl^-(s)$$
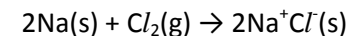
Na is oxidized as it loses electron to Cl_2 to form Na^+ ions. Cl_2 is oxidizing agent.
$$2Na(s) \rightarrow 2Na^+(s) + 2e^-$$
Cl_2 is reduced as it gains electron from Na to form Cl^- ions. Na is reducing agent.
$$Cl_2(g) + 2e^+ \rightarrow 2Cl^-(s)$$
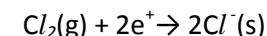
Redox reactions relating electron transfer:
1. Reaction of metal + dilute acid
2. Displacement reactions

Oxidation State in Reduction-Oxidation reaction

Oxidation State is the charge an atom would have if it existed as an ion
To work out oxidation state, the rules are:
- Free elements have oxidation state zero, e.g. Cu, Fe, N_2
- Oxidation of an ion is the charge of the ion, e.g. Na^+ = +1, Cu^{2+}=+2, O^{2-} = -2
- The oxidation state of some elements in their compounds is fixed, e.g.
  Group I Elements = +1
  Group II Elements = +2
  Hydrogen in most compounds = +1
  Iron or copper can have either +1, +2, +3, so it's not fixed
- Oxidation states of the elements in a compound adds up to zero, e.g.

$NaCl$: $(+1) + (-1) = 0$

K_2O: $(+1)$ x $2 + (-2) = 0$

Al_2O_3: $(+3)$ x $2 + (+2)$ x $3 = 0$

- Sum of oxidation states of elements in an ion is equal to charge on the ion, e.g.

OH^-: $(-2) + (+1) = -1$

Examples:

Work out the oxidation states of the <u>underlined</u> elements in these compounds:

(a) <u>C</u>O$_2$

(oxidation state of C) + (-2) x 2 = 0

(oxidation state of C) + (-4) = 0

∴ Oxidation state of C = +4

(b) K<u>Mn</u>O$_4$

(+1) + (oxidation state of Mn) + (-2) x 4 = 0

(oxidation state of Mn) + (+1) + (-8) = 0

(oxidation state of Mn) + (-7) = 0

∴ Oxidation state of Mn = +7

(c) <u>Fe</u>(NO$_3$)$_2$

(oxidation state of Fe) + (-1) x 2 = 0

(oxidation state of Fe) + (-2) = 0

∴ Oxidation state of Fe = +2

Note: Transition metals and some common elements may have different oxidation states in different compounds.

Examples of elements with variable oxidation states

Oxidation state	-2	-1	0	+1	+2	+3	+4	+5	+6	+7
Manganese			Mn		$MnCl_2$		MnO_2			$KMnO_4$
Chromium			Cr		$CrCl_2$	$CrCl_3$			$K_2Cr_2O_7$	
Iron			Fe		$FeCl_2$	$FeCl_3$				
Sulphur	FeS		S				SO_2		H_2SO_4	
Carbon			C		CO		$CaCO_3$			

Some compounds with possible variable oxidation states have roman numeral as a guide about their oxidation state, e.g.

- Iron(II) chloride has formula $FeCl_2$ and iron oxidation state +2
- Potassium(VI) dichromate has formula $K_2Cr_2O_7$ and potassium oxidation state +6
- Manganese(IV) oxide has formula MnO_2 and manganese oxidation state +4

Oxidation is the *increase of oxidation state* by a substance

Reduction is the *decrease of oxidation state* by a substance

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Note:

- Losing electrons means gain in oxidation state
- Gaining electrons means loss in oxidation state

Examples:

Metals with acids

$$Cu(s) + HCl(aq) \rightarrow CuCl_2(aq) + H_2(g)$$
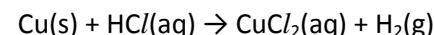
Cu is oxidized as it gains oxidation state from 0 to +2. Cu is reducing agent

H^+ ions in HCl reduced as it loses oxidation state from +1 to 0. H^+ ions are oxidising agent

Halide (Halogen) Displacement Reactions

$$Cl_2(aq) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(aq)$$
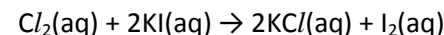
I^- ions in KI oxidized as it gains oxidation state from -1 to 0. I^- ions is reducing agent

Cl_2 is reduced as it loses oxidation state from 0 to -1. Cl_2 is oxidizing agent

Test for Oxidising/Reducing Agents

Oxidizing agents		
Name of compound	**Formula**	**Applications**
Potassium dichromate(VI)	$K_2Cr_2O_7$	Test for reducing agent; orange $K_2Cr_2O_7$ reduces to green Cr^{3+} ions
Potassium manganate(VII)	$KMnO_4$	Test for reducing agent; purple $KMnO_4$ reduces to colourless Mn^{2+} ions
Chlorine	Cl_2	Oxidizes Br^- to Br_2 and I^- to I_2; green-yellow Cl_2 reduces to colourless Cl^- ions

Reducing agents		
Name of compound	**Formula**	**Applications**
Potassium Iodide	KI	Test for oxidizing agent; colourless I^- ions oxidizes to brown I_2
Carbon monoxide	CO	Reduces metal oxide to metal in heat
Hydrogen	H_2	Reduces copper(II) oxide to copper
Sulfur dioxide	SO_4	used as bleach and preservative
Metals (highly reactive)	Na, Mg, etc.	Displaces less reactive metals

Not Redox!

- **Decomposition of carbonates by heat**: $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$

  The oxidation state of each element don't change. This is not a redox reaction.

- **Neutralization**: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O (l)$

  The oxidation state of each element don't change. This is not a redox reaction.

- **Precipitaion reactions**: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

  The oxidation numbers of silver and chloride ions unchanged. This is not redox.

6.3 Reversible Reactions

Reversible reactions are denoted by the sign "$\rightleftharpoons$" where the arrow $\rightarrow$ denotes **forward reaction**, where reactants react to form products, and the arrow $\leftarrow$ denotes **backward reaction** where products decompose to reform reactants. The reactions occur at the same time. E.g. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Effect of Temperature on Reversible Reactions

With higher temperature, the condition is now favored to break the bonds of the product formed (The bonding of products requires low temperatures). Thus, the products decompose to its constituents, leading to backward reaction.

Effect of Pressure on Reversible Reactions

Increase in pressure encourages forward reaction because the higher pressure the more reactants collide to react.

Dynamic equilibrium

Dynamic Equilibrium is the state when the rate of forward reaction is the same as the rate of backward reaction. Both reactants are reacted and products decompose at the same rate. Hence, there is no overall change in the amounts of reactants and products.

When we remove the products, it will also encourage forward reaction as the reaction would try to achieve equilibrium. Similar thing happens when we remove the reactants, that the decomposition of products is encouraged to reach the point

END OF CHAPTER 6

CHAPTER 7 - THE CHEMISTRY AND USES OF ACIDS, BASES AND SALTS

7.1 The Characteristics of Acids and Bases

Common Acids

Acids in daily life:

- Ethanoic acid – found in vinegar and tomato juice
- Citric acid – found in citrus foods like lemons, oranges and grapefruit
- Lactic acid – found in sour milk and yoghurt, and in muscle respiration
- Tartaric acid – found in grapes
- Tannic acid – found in tea and ant's body
- Formic acid – found in bee stings
- Hydrochloric acid – found in stomach juices

Laboratory acids: 3 common laboratory acids

- Hydrochloric acid (HCl)
- Sulphuric acid (H_2SO_4)
- Nitric acid (HNO_3)

Dilute acids – solution containing small amount of acid dissolved in water

Concentration acids – solution containing large amount of acid dissolved in water

Properties of Dilute Acids

- **Acids have a sour taste**

- **Acids are hazardous**

  acids are irritants (they cause skin to redden and blister)

- **Acids change the colour of indicators**

  Acids turn common indicator litmus – blue litmus to red

- **Acids react with metals**

  Acids react with metals to produce hydrogen gas. The gas is tested with a burning splint which shows hydrogen burns with a 'pop' sound.

  $$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$$
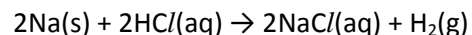
- **Acids react with carbonates and hydrogencarbonates (bicarbonates)**

  Carbon dioxide is to be formed. To test this, the gas produced is bubbled into limewater which forms a white precipitate.

  Carbonates:

  $$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O \ (l)$$

  Bicarbontes:

  $$NaHCO_3(s \text{ or } aq) + HCl(aq) \rightarrow NaCl_2(aq) + CO_2(g) + H_2O \ (l)$$
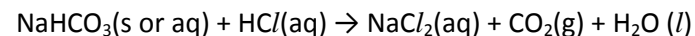
- **Acids react with metal oxides and hydroxides**

  Metal oxides & hydroxides react slowly with warm dilute acid to form salt+water

  $$Cu(OH)_2(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + 2H_2O \ (l)$$
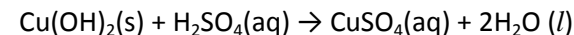
Storage of Acids

Acids are stored in claypots, glass or plastic containers as sand, glass and plastic do not react with acids. If it's stored in metal container, metal would react with acids

Uses of Acids

- Sulphuric Acid - Used in car batteries
  - Manufacture of ammonium sulphate for fertilisers
  - Manufacture of detergents, paints, dyes, artificial fibres & plastics

- Hydrochloric acid can remove rust (iron(III) oxide) which dissolves in acids

- Acids are used in preservation of foods (e.g. ethanoic acid)

7.2 Acids and Hydrogen Ions

The Need for Water in Acids

Acids are covalent compounds and do not behave as acids in the absence of water as water reacts with acids to produce H^+ ions, responsible for its acidic properties.

e.g. Citric acid crystals doesn't react with metals and doesn't change colours of indicators; citric acid in water reacts with metals and change turns litmus red.

Hydrogen Ions

Hydrogen gas is formed by acids as $H^+(aq)$ ions are present in acid solutions

- This means when a solid/gas acid dissolved in water, they produce H^+ ions in it

Chemical eqation: $HCl(s) \xrightarrow{water} HCl(aq)$

Ionic Equation: $HCl(s) \xrightarrow{water} H^+(aq) + Cl^-(aq)$

Note that for ionic equation only aqueous solutions are ionised

- However when dissolved in organic solutions, they don't show acidic properties

When metals react with acids, only the hydrogen ions react with metals, e.g.:

Chemical equation: $2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$

Ionic equation: $2Na(s) + 2H^+(aq) \rightarrow 2Na^+(aq) + H_2(g)$

Basicity of an acid is maximum number of H^+ ions produced by a molecule of acid

Some Acids With Their Basicity		
Acids	**Reaction with water**	**Basicity**
Hydrochloric acid	$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$	monobasic
Nitric acid	$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$	monobasic
Ethanoic acid	$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$	monobasic
Sulphuric acid	$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$	dibasic

The fizz of drinks

Soft drink tablets contains solid acid (e.g. citric acid, $C_6H_8O_7$) & sodium bicarbonate

- When tablet is added to water, citric acid ionises and the H^+ produced reacts with sodium bicarbonate to produce carbon dioxide gas, making them fizz

Strong and Weak Acids

Strong Acid - acid that completely ionises in water. Their reactions are irreversible.

E.g. H_2SO_4, HNO_3, HCl

$$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$
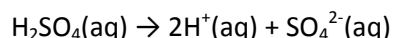
In above H_2SO_4 has completely been ionized in water, forming 3 kinds of particles:

- H^+ ions
- SO_4^{2-} ions
- H_2O molecules

Strong acids react more vigorously with metals than weak acids – hydrogen gas bubbles are produced rapidly

Weak Acids - acids that partially ionise in water. The remaining molecules remain unchanged as acids. Their reactions are reversible. E.g. CH_3COOH, H_2CO_3, H_3PO_4

$$H_3PO_4(aq) \rightleftharpoons 3H^+(aq) + PO_4^{2-}(aq)$$

Weak acids react slowly with metals than strong acids – hydrogen gas bubbles are produced slowly.

Comparing Strong and Weak Acids with Concentrated and Dilute Acids

CONCENTRATION	STRENGTH
Is the amount of solute (acids or alkalis) dissolved in 1 dm^3 of a solution	Is how much ions can be disassociated into from acid or alkali
It can be diluted by adding more water to solution or concentrated by adding more solute to solution	The strength cannot be changed

Comparing 10 mol/dm^3 and 0.1 mol/dm^3 of hydrochloric acids and 10 mol/dm^3 and 0.1 mol/dm^3 of ethanoic acids

- 10 mol/dm^3 of ethanoic acid solution is a concentrated solution of weak acid
- 0.1 mol/dm^3 of ethanoic acid solution is a dilute solution of weak acid
- 10 mol/dm^3 of hydrochloric acid solution is a concentrated solution of strong acid
- 0.1 mol/dm^3 of hydrochloric acid solution is a dilute solution of strong acid

Bases and Alkalis

Bases are oxides or hydroxides of metals

Alkalis are bases which are soluble in water

Laboratory Alkalis

- Sodium Hydroxide, $NaOH$
- Aqueous Ammonia, NH_4OH
- Calcium Hydroxide, $Ca(OH)_2$

All alkalis produces hydroxide ions (OH^-) when dissolved in water. Hydroxide ions give the properties of alkalis. They don't behave as acids in absence of water.

Alkalis are therefore substances that produce hydroxide ions, $OH^-(aq)$, in water.

Properties of Alkalis

- **Alkalis have a slippery feel**

- **Alkalis are hazardous**
  - Dilute alkalis are irritants
  - Concentrated alkalis are corrosive and burn skin (**caustic**(i.e. burning) **alkalis**)

- **Alkalis change the colour of indicators**
  Alkalis turn common indicator litmus – red litmus to blue

- **Alkalis react with acids**
  The reaction is called **neutralisation**

- **Alkalis react with ammonium compounds**
  They react with heated solid ammonium compounds to produce ammonia gas
  $$(NH_4)_2SO_4(s) + Ca(OH)_2(aq) \rightarrow CaSO_4(aq) + 2NH_3(g) + 2H_2O(l)$$

- **Alkalis react with solutions of metal ions**
  Barium sulphate, $BaSO_4(aq)$, contains $Ba^{2+}(aq)$ ions
  $$Ca(OH)_2(aq) + BaSO_4(aq) \rightarrow Ba(OH)_2(s) + CaSO_4(aq)$$
  The solid formed is precipitate – the reaction is called *precipitate reaction*

Strong and Weak Alkalis

Strong Alkalis - base that completely ionises in water to form $OH^-(aq)$ ions. Their reactions are irreversible. E.g. NaOH, KOH, $Ca(OH)_2$
$$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$$
Weak Alkalis - base that partially ionise in water. The remaining molecules remain unchanged as base. Their reactions are reversible. E.g. NH_3
$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$
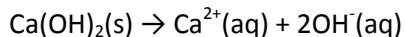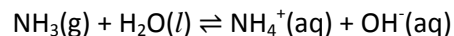
Uses of Alkalis

- Alkalis neutralise acids in teeth (toothpaste) and stomach (indigestion)
- Soap and detergents contain weak alkalis to dissolve grease

- Floor and oven cleaners contain NaOH (strong alkalis)
- Ammonia (mild alkalis) is used in liquids to remove dirt and grease from glass

Indicators and pH

Indicators are substances that has different colours in acidic and alkaline solutions
Common indicators:
➢**Litmus**
➢**Methyl orange**
➢**Phenolphtalein**
The table shows the change of colours made by some indicators

Indicator	Colour in acids	colour changes at pH	Colour in alkalis
Phenolphtalein	Colourless	9	Pink
Methyl orange	Red	4	Yellow
Litmus	Red	7	Blue
Screened methyl orange	Red	4	Green
Bromothymol blue	Yellow	7	Blue

The pH Scale

A measure of acidity or alkalinity of a solution is known as **pH**

- pH 7 is neutral – in pure water
- solutions of less than pH 7 are acidic. The solutions contain hydrogen ions. The smaller pH, the more acidic the solution is and more hydrogen ions it contains.
- solutions of more than pH 7 are alkaline. The solution contains hydroxide ions. The biger pH, the more alkaline the solution and more hydroxide ions it contains.
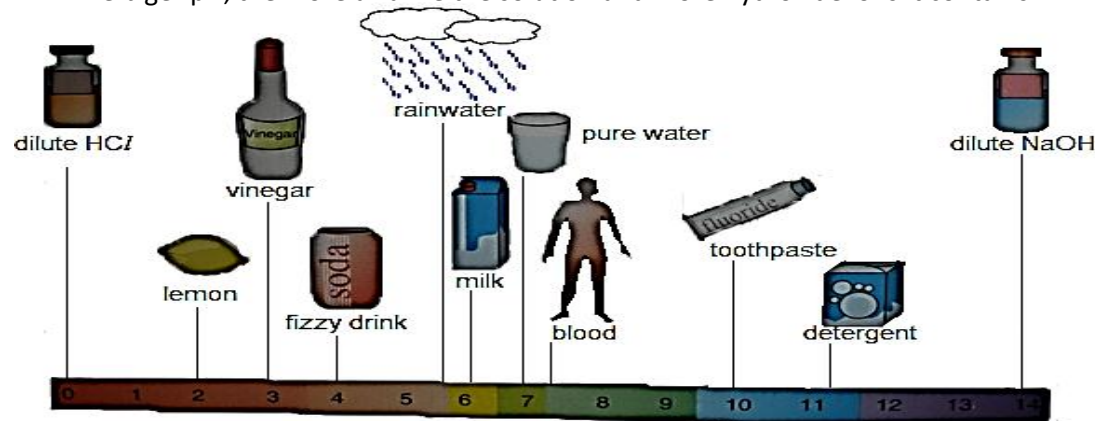
Measuring pH of a Solution

1. **Universal indicators**

   It can be in paper or solution form. Universal paper can be dipped into a solution then pH found is matched with the colour chart. It gives approximate pH value.

2. **pH meter**

   A hand-held pH probe is dipped into solution and meter will show the pH digitally or by a scale. Measures pH water in lakes, water, and streams accurately

3. **pH sensor and computer**

   A probe is dipped into solution and will be sent to computer through interface used to measure pH of solution. The pH reading is displayed on computer screen.

pH Around Us

- Substances in body involved in good digestion have different pH values
- Blood to heart and lungs contains CO_2 making blood slightly acidic
- Acids are used in food preservations (ethanoic acid to preserve vegetables; benzoic acid used in fruit juices, jams and oyster sauce)
- pH affects plant growth – some plants grow in acidic soil; some need alkaline soil
- When hair is cleaned with shampoo which is alkali to dissolve grease, hair can be damaged unless it's rinsed or acid conditioner is used to neutalise excess alkali

Ionic Equations

Ionic equation is equation involving ions in aqueous solution, showing formation and changes of ions during the reaction

Rule to make ionic equations:

- Only formulae of ions that change is included; ions don't change = omitted
- Only aqueous solutions are written as ions; liquids, solids and gases written in full

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Reaction Between Metals and Acids

For example, reaction of sodium with hydrochloric acid

$$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$$

Its ionic equation is written as:

$$2Na(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow 2Na^+(aq) + 2Cl^-(aq) + H_2(g)$$
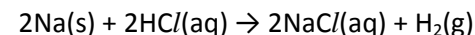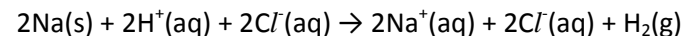
Since 2 Cl^-(aq) ions don't change, they're not involved in reaction. As ionic equation is used to *show changes in reactions*, we omit Cl^-(aq) ions. So we're left with:

$$2Na(s) + 2H^+(aq) \rightarrow 2Na^+(aq) + H_2(g)$$
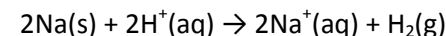
Reaction Between Soluble Ionic Compounds and Acids

e.g. Reaction of sodium hydrogencarbonate with hydrochloric acid

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O\ (l)$$

Its ionic equation is:

$$Na^+(aq) + H^+(aq) + CO_3^{2-}(aq) + H^+(aq) + Cl^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + CO_2(g) + H_2O(l)$$
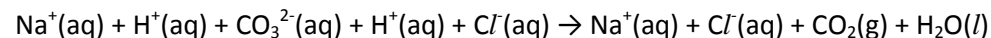
Since Na^+(aq) and Cl^-(aq) ions don't change, we omit them, leaving:

$$H^+(aq) + CO_3^{2-}(aq) + H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$
$$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$
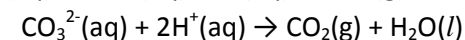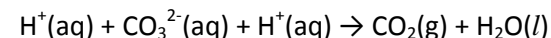
Reaction Between Insoluble Ionic Compounds and Acids

e.g. Reaction between iron(II) oxide and sulphuric acid

$$FeO(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2O(g)$$

Its ionic equation is:

$$FeO(s) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow Fe^{2+}(aq) + SO_4^{2-}(aq) + H_2O(g)$$
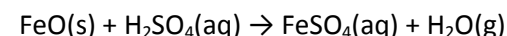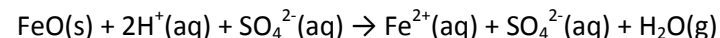
Note: FeO is written in full as it's solid (although it's an ionic compound)

Since SO_4^{2-}(aq) ions don't change, we omit SO_4^{2-} ions, leaving:

$$FeO(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2O(g)$$
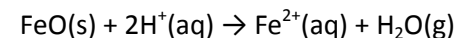
E.g. Reaction between calcium carbonate and hydrochloric acid

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

Its ionic equation is:

$$CaCO_3(s) + 2H^+(aq) + 2Cl^-(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + CO_2(g) + H_2O(l)$$
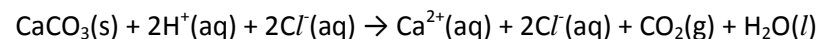
Since 2 Cl^-(aq) ions don't change, we omit Cl^- ions, leaving:

$$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$$

Reactions Producing Precipitate

E.g. Reaction between calcium hydroxide and barium sulphate

$$Ca(OH)_2(aq) + BaSO_4(aq) \rightarrow Ba(OH)_2(s) + CaSO_4(aq)$$

Its ionic equation is written as:

$$Ca^{2+}(aq) + 2OH^-(aq) + Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Ba(OH)_2(s) + Ca^{2+}(aq) + SO_4^{2-}(aq)$$
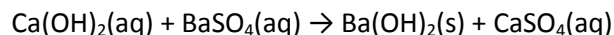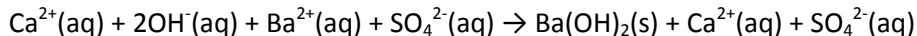
Since $Ca^{2+}(aq)$ and $SO_4^{2-}(aq)$ ions don't change, we omit them, leaving:

$$Ba^{2+}(aq) + 2OH^-(aq) \rightarrow Ba(OH)_2(s)$$
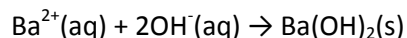
Displacement Reactions

E.g. Reactions between magnesium with zinc sulphate

$$Mg(s) + ZnSO_4(aq) \rightarrow MgSO_4(aq) + Zn(s)$$

Its ionic equation is written as:
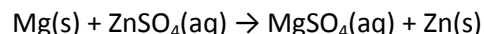
$$Mg(s) + Zn^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq) + Zn(s)$$

Since $SO_4^{2-}(aq)$ ions don't change, we omit them, leaving:

$$Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$$
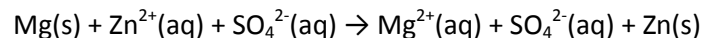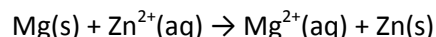
Neutralization is the reaction between acid and base to form salt and water only.

From ionic equation, we know that the reaction only involves H^+ ions from acids with OH^- ions from alkali to form water .

E.g. NaOH + H_2SO_4 forms Na_2SO_4 + H_2O

$$H_2SO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(g)$$

Ionic equation is:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(g)$$
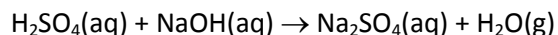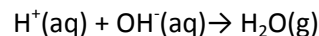
Plants don't grow well in acidic soil. Quicklime (calcium hydroxide) is added to neutralise the acidity of soil according to equation:

$$Acid(aq) + Ca(OH)_2(aq) \rightarrow Ca(acid\ anion)(aq) + H_2O(g)$$
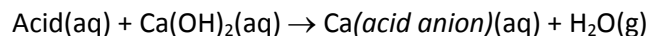
Reaction between Base and Ammonium Salts

E.g. Reaction between NaOH and NH_4OH

$$NaOH(aq) + NH_4Cl(aq) \rightarrow NaCl(aq) + NH_3(g) + H_2O(g)$$

Ionic equation:

$$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(g) + H_2O(g)$$
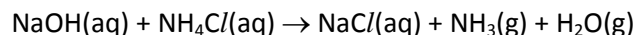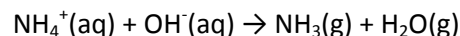
OXIDES

Acidic Oxide	Basic Oxide	Amphoteric Oxide	Neutral Oxide
Oxides of non-metals, usually gases which reacts with water to produce acids, e.g. CO_2, NO_3, P_4O_{10}, SO_2	Oxides of metals, usually solid which reacts with water to produce alkalis, e.g. CaO, K_2O, BaO	Oxides of transition metals, usually solid, which reacts with acids/alkalis to form salt and water, e.g. Al_2O_3, FeO, PbO	Oxides that don't react with either acids/alkalis, hence do not form salts, e.g. H_2O, CO, NO

7.3 Preparation of Salts

Soluble and Insoluble Salts

Soluble	Insoluble
All Nitrates	-
All Supates	$BaSO_4$, $CaSO_4$, $PbSO_4$
All Chlorides	$PbCl$, $AgCl$
Potassium, Sodium, Ammonium salts	-
K_2CO_3, Na_2CO_3, NH_4CO_3	All Carbonates
K_2O, Na_2O	All Oxides

Preparation of Insoluble Salts

Insoluble salts, e.g. $BaSO_4$, $CaSO_4$, $PbSO_4$, $PbCl$, $AgCl$ and most carbonates, can be prepared by reacting compound containing the wanted cation with another compound containing the wanted anion. This is *precipitation reaction*.

E.g. Preparation of $BaSO_4$

First $BaCl$, since it contains wanted barium ion, is reacted with H_2SO_4, since it contains wanted sulphate ion, to produce solid $BaSO_4$ & aqueous KCl. $BaSO_4$ then separated from KCl by filtration, leaving filtrate KCl & $BaSO_4$ left on filter paper. Salt is washed with water to completely remove KCl & filter paper is squeezed with another filter paper to dry $BaSO_4$.
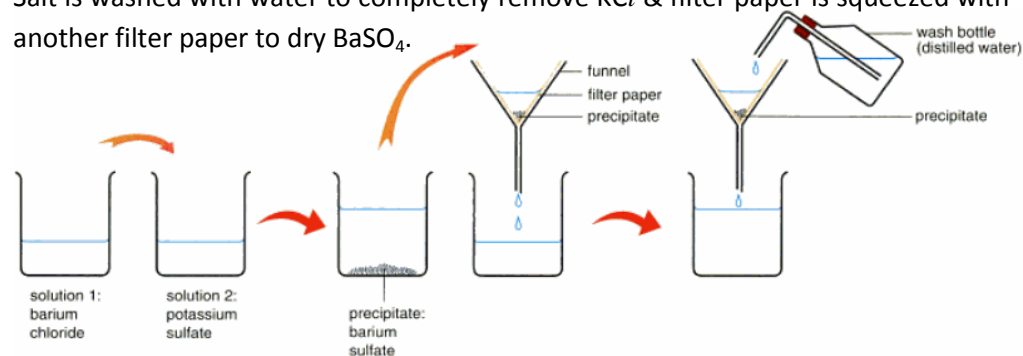
Preparation of Soluble Salts

By Neutralization

$25.0cm^3$ acid, as standard solution, is placed in conical flask using pipette. Add few drops of indicator & titrate with alkali from burette until indicator changes colour, showing all acid has just reacted. Volume of alkali added is measured. Prepare new $25.0cm^3$ acid again with pipette & add same volume of alkali as before to prevent excess alkali/acid because both reactant & product are aqueous. Next, the product is evaporated to dryness to obtain the salt.

pipette

sulfuric acid

phenolphthalein indicator

burette

sodium hydroxide solution

sulfuric acid + indicator

solution of sodium sulfate

1. Using a pipette, 25.0 cm³ of sulfuric acid is placed in a conical flask. A few drops of indicator are added (e.g. phenolphthalein).
2. Dilute sodium hydroxide is placed in a burette and then released into the conical flask until the indicator changes colour (which means all the acid has just reacted). The volume of alkali added from the burette to the acid is measured.

3. The experiment is now repeated. 25.0 cm³ of sulfuric acid is placed in a titration flask (as before), but no indicator is added. The sodium hydroxide is placed in a burette and the same volume of this alkali is added to the flask as before. The flask then contains a solution of sodium sulfate without excess acid or alkali.
4. The sodium sulfate is obtained by evaporating most of the water and crystallising the salt.

By Reacting Metal with Acid

zinc powder

sulfuric acid

reacting zinc powder

glass rod

funnel

filter paper

excess zinc

reaction mixture from step 1
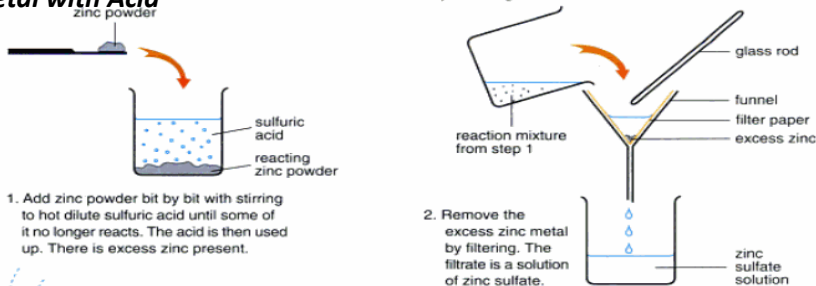
1. Add zinc powder bit by bit with stirring to hot dilute sulfuric acid until some of it no longer reacts. The acid is then used up. There is excess zinc present.

2. Remove the excess zinc metal by filtering. The filtrate is a solution of zinc sulfate.

zinc sulfate solution

evaporating dish

solution of zinc sulfate from step 2

crystals of zinc sulfate

funnel
filter paper
crystals of zinc sulfate

filter paper

3. Pour the solution of zinc sulfate into an evaporating dish and heat to evaporate most of the water. This produces a hot saturated solution of zinc sulfate. Allow the solution to cool.

4. Filter the zinc sulfate crystals and dry them by squeezing them between sheets of filter paper.

Only metals like zinc and magnesium, which moderately react with dilute acids, are used.

E.g. Reacting Zn with H_2SO_4 to prepare $ZnSO_4$

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$$
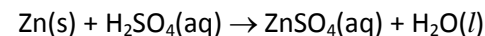
Zn is added to H_2SO_4 until in excess to ensure no more H_2SO_4 is present. Then the mixture is filtered off to separate Zn from $ZnSO_4$. The filtrate ($ZnSO_4$) is then placed in evaporating dish to evaporate most of water then it's cooled after $ZnSO_4$ crystals are formed. The crystals then filtered and squeezed between filter papers to dry.

By Reacting Insoluble Base with Acid

E.g. Reacting MgO with Acids

$$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l)$$
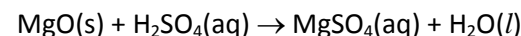
The same method as reaction of acid with metal is used, so refer to diagram and above explanation, substituting reactants and products.

By Reacting Carbonate with Acid

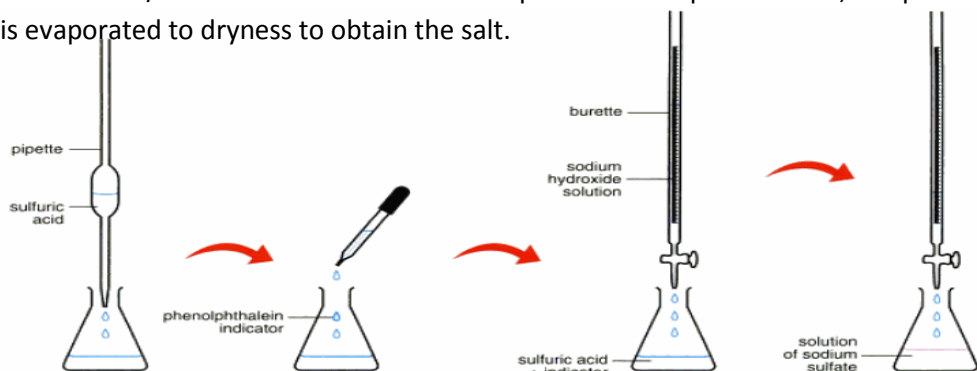E.g. Reacting $CaCO_3$ with Acids

$$K_2CO_3(s) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + CO_2(g) + H_2O(l)$$
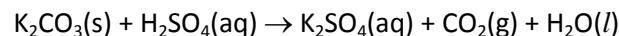
The same process is used as reaction of acid with metal, just that carbon dioxide is produced. Carbon dioxide can be tested by bubbling it into limewater which will turn limewater colourless to milky.

7.4 Properties and Uses of Ammonia

Ammonia and its Uses

Are produced from nitrogen reacted with hydrogen

For producing: fertilisers, nitric acid, nylon, dyes, cleaners and dry cell

The Manufacture of Ammonia: The Haber Process

The Process

Nitrogen and hydrogen are mixed together in ratio 1:3, where nitrogen is obtained from air and hydrogen is obtained from natural gas, and passed over iron catalyst.

unreacted N_2 and H_2 recycled

N_2 (1 volume)

purifier and dryer

N_2, H_2

compressor (200 atm)

catalyst chamber 450 °C iron catalyst

condensation

liquid ammonia

H_2 (3 volumes)
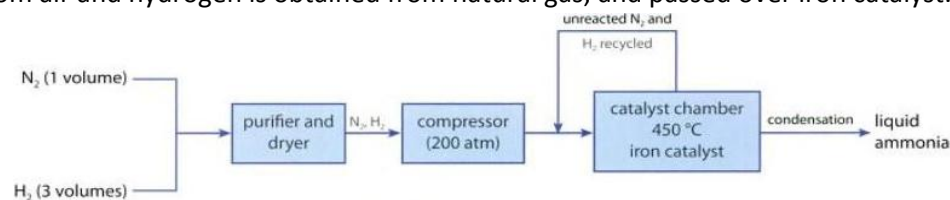
Since the reaction is reversible so H_2 and O_2, reproduced from decomposition of produced NH_3, are passed over the catalyst again to produce ammonia.

Conditions for Manufacturing Ammonia
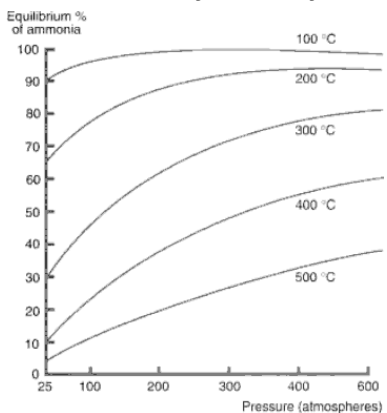
Graph shows that to have high yield of ammonia we should have:
1. Higher pressure
2. Lower temperature

But in practice, we use lower pressure of 200 atm and higher temperature of 450^oC. This is because:
- Using low temperature is too slow to reach equilibrium
- Using high pressure involves safety risk and higher cost

Ammonia as Fertilizers

Plants need nitrogen as one of component for growth and ammonium fertilizers contain Nitrogen for that.

% content of nitrogen in ammonium fertilizers

E.g. Ammonium sulphate, $(NH_4)_2SO_4$, and urea, $(NH_2)_2CO$, are 2 kinds of fertilizers. Deduce, in terms of nitrogen content, which of these fertilizers best for plants.

$$\% \text{ mass} = \frac{\text{total mass of element}}{\text{total mass of compound}} \times 100$$

$(NH_4)_2SO_4 = \frac{2 \times N}{2(N+(4 \times H)) +S+(4 \times O)} \times 100$ $(NH_2)_2CO = \frac{2 \times N}{2(N+(2 \times H)) +C+O} \times 100$

$\qquad = \frac{2 \times 14}{2(14+(4 \times 1))+ 32+(4 \times 16)} \times 100$ $= \frac{2 \times 14}{2(14+(2 \times 1)) +12+16} \times 100$

$\qquad = \frac{28}{132} \times 100$ $= \frac{28}{60} \times 100$

$\qquad = 21.2\%$ of N $= 46.7\%$ of N

Therefore, $(NH_2)_2CO$ is a better fertilizer since it contains more nitrogen

---

Problems with Ammonia

Eutrophication is the inrease in organic content of water when fertilizers leach into soil and washed into rivers and streams.

When excess fertilizers washed away by rain, nitrate ions in it gets into rivers and helps aquatic plants like algae to grow swiftly. When too much algae, water turns murky and sunlight would not penetrate into water to help their growth which in turn lead to deaths of algae. Decay of this organic matter uses up oxygen, hence killing aquatic animals. Then even more algae dies and even more animals die

Water pollution results from runoff of fertilizer use, leaching from septic tanks, sewage and erosion of natural deposits.

Nitrate ions from nitrogen in soil leaches down the soil into groundwater due to its solubility. Since groundwater is our drink source, when humans drink this water, they will get seriously ill and babies may suffer breathlessness to death.

7.4 Sulphuric Acid

About Sulphuric Acid

It is a colourless oily liquid with density slightly higher than water and high boiling point of 338°C. It's soluble in water and emits heat when dissolved.

The Contact Process

1. Sulphur is reacted with oxygen to produce sulphur dioxide, SO_2

$$S(g) + O_2(g) \rightarrow SO_2(g)$$
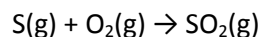
2. SO_2 gas is purified from impurities by passing it through dust settlers and washed with water then dried with concentrated H_2SO_4. If there's impurities, the catalyst will be "poisoned" and the reaction will be less effective.

3. Sulphur dioxide is further reacted with oxygen to produce sulphur trioxide.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$
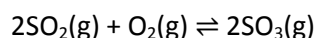
Since this is a reversible reaction, low temperature and high pressure is needed to enhance forward reaction.
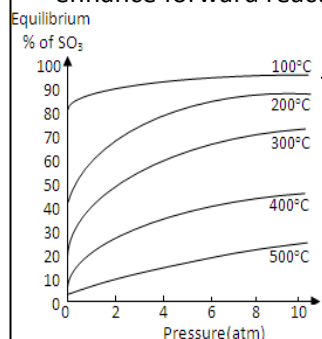
According to graph:
- Low temperatures yields high percentage of sulphur trioxide at equilibrium

Dynamic equilibrium is the state when forward and backward reaction occurs at same speed and the concentration of reactant and product is equal. The reactions do not stop and some reactants and products always remain.

Although lower temperature is required for high yield, 450°C is instead used in the process as the reaction will be too slow with low temperatures.

Pressure of 2-3 atm is used in practice.

A catalyst vanadium(V) oxide, V_2O_5, is used to increase the yield of sulphur trioxide. Platinum can catalyse more efficiently but it is too expensive.

4. Sulphur trioxide is cooled & reacted in concentrated H_2SO_4 to produce oleum, $H_2S_2O_7$. SO_3 is not reacted with water right now as mist forms at this temperature.

$$SO_3 + H_2SO_4 \rightarrow H_2S_2O_7$$
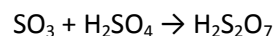
5. Oleum is diluted with water to produce sulphuric acid.

$$H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$$
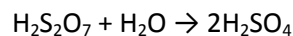
Uses of Sulphuric Acid

- Making of fertilizers such as superphosphate and ammonium sulphate
- Making detergents
- Cleaning surfaces of iron and steel surface before galvanization or electroplating
- To manufacture plastics and fibres
- As electrolyte in car batteries
- In refining of petroleum
- In production of dyes, drugs, explosives, paints, etc.

Uses of Sulphur Dioxide

Sulphur dioxide is a colourless, toxic gas with suffocating smell, denser than the air and dissolves in water. Sulphur dioxide is emitted from electric power plants and smelting operations to produce copper, zinc, lead and nickel from sulfide ores and it's a major contributor to acid rain. It is used:

- In sulfite manufacturing used in digestion & bleaching wood pulp to make paper
- As food preservatives such as dried fruit and fruit juices. It's not used to preserve meat as it destroys vitamin B_1
- To bleach straw

END OF CHAPTER 7

CHAPTER 8 – THE PERIODIC TABLE

The Progenitor Periodic Table

First periodic table made by Dimitri Mendeleev in 1869.

The Modern Periodic Table

Differences to Mendeleev's:

- 115 elements while for Mendeleev's is 69.
- Mendeleev arranged the elements according to *relative atomic mass* while today we arrange according *proton number*.

Period – horizontal row of elements in periodic table

Group – vertical column of elements in periodic table numbered from I to 0

Elements between Group II and Group III – ***transition metals***

8.1 Patterns in the Periodic Table

1) Electronic Structure

Elements in same group has the same number of valence shell electrons which the amount is the same as the group number.

e.g. Group II has elements with valency of 2 electrons.

2) Charges on Ions

The charges relates to the group number and number of valence electrons.

- Elements on left side periodic table lose ions to form cation.
- Elements on right side periodic table lose ions to form anion.
- Elements in Group IV can lose or gain electrons depending on reacting element.
- Transition metals may form variable cation of 2+ or 3+

Group Number	I	II	III	IV	V	VI	VII	0
Formula of ion	+1	+2	+3	varies	-3	-2	-1	stable

3) Bonding

Elements in same group form same type and number of bonds due to the same number of valence electrons.

e.g. Sodium in Group I forms NaC*l*, so other elements in Group I does the same.

(RbC*l*, KC*l*, LiC*l*, CsC*l*)

4) Metals and Non-metals

Metals	Non-metals
On the left side of periodic table	On the right side of periodic table
Have fewer (≤4) valence electrons	Have more (>4) valence electrons

From left to right, elements gradually change from metal to non-metal

Elements close to dividing line in periodic table in back part of the note (in bold) are called **metalloids** having properties of metals and non-metals.

5) Changes in Group

- Proton number increase going down the group
- On each sides of periodic table, the change of the proton number small & gradual
- In transition metals, the gradual change is larger

Using the Periodic Table

Predicting Properties

1) **Formula and Structures**

   Given chlorine, iodine and bromine of Group VII forms molecules of Cl_2, I$_2$ and Br$_2$ respectively, predict the molecular formula of Fluorine.

   $\therefore F_2$

   From example, we know elements in same group form same formula.

2) **Properties of Elements**

   Properties of element changes down the group.

   i.e. given list of Group 7 elements, predict the properties of astatine.

Element	Proton Number	Melting Point (^{o}C)	Boiling Point (^{o}C)
Fluorine	9	-220	-118
Chlorine	17	-101	-35
Bromine	35	-7	59
Iodine	53	114	184
Astatine	85	> 114	> 184

8.2 Group Properties

Group I Elements – The Alkali Metals

These are metals which react with water to form alkaline solutions. The solutions turn red litmus paper blue.

- most reactive metals in periodic table
- have one outer shell electrons
- shiny, silvery solids
- soft, easily cut with scalpel
- low densities & melting points. These increases down the group
- reacts easily in air. So they're kept in oil
- reacts vigorously (may catch fire or explode) with cold water
- they make ionic compounds of +1 charge. They have similar formulae
- they become more reactive down the group

Name	Symbol	Density (g/cm^3)	Melting point (^{o}C)
Lithium	Li	0.53	180
Sodium	Na	0.97	98
Potassium	K	0.86	64
Rubidium	Rb	1.5	39
Caesium	Cs	1.9	29

Element	Chloride	Nitrate	Sulphate	Oxide
Lithium	LiCl	LiNO$_3$	Li$_2$SO$_4$	Li$_2$O
Sodium	NaCl	NaNO$_3$	Na$_2$SO$_4$	Na$_2$O
Potassium	KCl	KNO$_3$	K$_2$SO$_4$	K$_2$O

Table 12.1: The physical properties and formulae of Group I metals

Group VII Elements – The Halogens

These are elements which reacts with most metals to form salts
- very reactive elements
- have seven outer shell electrons
- each molecule in the element is diatomic (contains two atoms, eg F$_2$)
- elements become darker and solidify down the group
- they have low melting and boiling points which increases down the group
- all halogens are poisonous

Element	Molecular formula	Melting point (^{o}C)	Boiling point (^{o}C)	State at r.t.p.	Colour
Fluorine	F$_2$	-220	-189	gas	Pale yellow
Chlorine	Cl_2	-101	-35	gas	Greenish yellow
Bromine	Br$_2$	-7	59	liquid	Reddish brown
Iodine	I$_2$	114	184	solid	Shiny black

Compounds of the Halogens

Halogens gives a charge of –1, so they give similar formulae, eg: NaBr, NaI

Reactions of the Halogens

- reacts vigorously with metals to form ionic salts for the equation:
$$2K + Br_2 \rightarrow 2KBr$$
- halogens become less reactive down the group
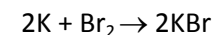
Displacement Reactions

More reactive halogen displaces less reactive halogen
eg: aqueous fluorine was added into sodium bromide solution. State the chemical equation of the reaction.
$$F_2(aq) + 2NaBr(aq) \rightarrow 2NaF + Br_2$$
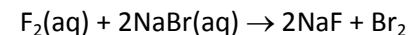
Group 0 Elements – The Noble Gases

Are least reactive elements in the state of gas. They do not form bonds
- have stable electronic configuration with full electrons on their shells
- coloured gases consisting of single atoms (*monoatomic*)
- low melting and boiling points

Uses of the Noble Gases

- argon used in light bulbs as it wouldn't react with the hot filament
- neon used in neon advertising strip lights
- helium used in small and weather balloons, and airships for less density

8.3 Transition Elements

Properties

- First transition series are all metals
- Transition elements have high melting points
- They have high density
- They have variable oxidation state, e.g. Iron (Fe) appear as Fe^{2+} or Fe^{3+}
- They form coloured compounds, e.g. $CuSO_4$ is blue, $FeSO_4$ is green
- They form complex ions, e.g. MnO_4^-, Manganate(VII) ions
- They act as catalysts

Uses of Transition Elements

Most transition elements and their compounds act as catalysts which speed up chemical reactions

- Iron is used in Haber Process for manufacture of ammonia
- Vanadium(V) oxide is used in contact process to manufacture sulphuric acid
- Nickel is used in hydrogenation of alkenes to form saturated fats (e.g. margarine)

Advantages

- Since transition elements speed up chemical processes in industries, they saves time in manufacture
- Less energy is needed for manufacture in industries, hence lower cost
- Since less energy is needed, more energy resources can be conserved, e.g. oil to generate electricity in producing iron.

END OF CHAPTER 8

CHAPTER 9 –METALS

9.1 Properties of Metals

Physical properties

- **Ductile** (can be stretched to form wires)
- **Malleable** (can be bent and beaten into different shapes)
- Good **conductors of electricity and heat**
- **Shiny**
- **High melting points** and **boiling points** (except mercury and sodium)
- **High density** (except sodium)
- **Strong**

ALLOYS

Alloy – a mixture of metallic elements or metallic with non-metallic.

Pure metals are weak as the layers of atoms slide over each other easily.

∴ in alloy of 2 metals, they have different sizes of atoms so this distrupts the orderly layer of atoms making it difficult for atoms to slide over.

Uses of Alloy:

- **Steel** (mixture of iron, little carbon and trace elements)
- **Brass** (copper and zinc) – tough and corrosive-resistant
- **Coin metals** (copper with other metals e.g. nickel) – tough, resistant and stand up to wear

Uses of Stainless Steel

is an alloy of iron containing chromium or nickel. Is the most expensive way

Applications for:

- Cutleries
- Medical instruments for hospital operations
- Kitchen sinks
- Steel objects in chemical factories and oil refineries
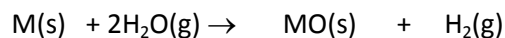
9.2 Reactivity Series

Reaction of Metals with Water

Pottasium, **Sodium**, and **Calcium** reacts with cold water to form:

$$M(s) + 2H_2O(l) \rightarrow MOH(aq) + H_2(g)$$

Metal + Water → Metal Hydroxide + Hydrogen

Magnesium, **Zinc**, **Iron** reacts with steam to form:

$$M(s) + 2H_2O(g) \rightarrow MO(s) + H_2(g)$$

Metal + Water → Metal Oxide + Hydrogen

Iron does not react with water

Copper and **Gold** have no reaction with water and steam

Reaction of Metals with Dilute Hydrochloric Acid

Pottasium, **Sodium**, **Calcium**, **Magnesium**, **Zinc** and **Iron** reacts with dilute hydrochloric acid to form:

$$M(s) + 2HCl(aq) \rightarrow MCl_2(aq) + H_2(g)$$

Metal + Acid → Metal Chloride + Hydrogen

Lead reacts with warm hydrochloric acid slowly

Copper and **Gold** have no reaction with dilute hydrochloric acid

Displacement Reactions

Displacement reaction is the displacement of ions of metal from compounds of metals lower in reactivity series by metals higher in reactivity series.

E.g. Magnesium displaces copper(II) chloride

$$Mg(s) + CuCl_2(aq) \rightarrow MgCl_2(aq) + Cu(s)$$
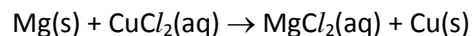
For observation, we'll see silver magnesium metal coated with brown copper metal

Displacement is due to Mg atoms transfer electrons to Cu^{2+} ions forming Cu atoms.

$$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$$
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$
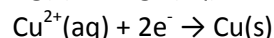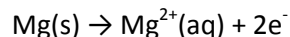
Loss of electrons is due to it's less reactive as less reactive metal has higher chance of losing electrons. That's why when Mg is placed in KCl, no reaction occurs.

$$Mg(s) + KCl_2(aq) \rightarrow \text{No reaction}$$
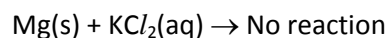
E.g. Displacement from metal oxides

Metal higher in reactivity series displaces oxides of metals lower in reactivity series.

When Ca burns with Ag_2O, Ca displaces Ag to produce CaO and Ag.

$$Ca(s) + Ag_2O(s) \rightarrow CaO(s) + 2Ag(s)$$
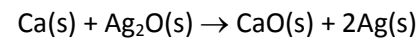
This is called **thermit reaction** large amount of heat is produced.

Reaction of Metal Oxides with Carbon

The lower the position of metal in reactivity series, the easier for carbon to remove oxygen from metal oxide by heating. At higher position, stronger heat is needed.

E.g. CuO reacts with C can be reduced by bunsen burner flame temperature

$$CuO(s) + C(s) \rightarrow Cu(s) + CO_2(g)$$
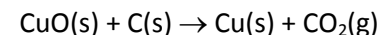
For iron oxide to be reduced, it needs very high temperature.

Reaction of Metal Oxides with Hydrogen

The lower position of metal in reactivity series, the easier hydrogen remove oxygen from metal oxide by heating. At higher position, stronger heat is needed.

E.g. PbO reacts with H_2 can be reduced by bunsen burner flame temperature

$$PbO(s) + H_2(g) \rightarrow Pb(s) + H_2O(l)$$
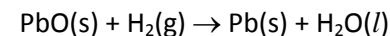
Decomposition of Metal Carbonates

The lower position of metal in reactivity series, the easier hydrogen remove oxygen from metal oxide by heating. At higher position, stronger heat is needed.

E.g. $CuCO_3$ reacts decomposes by heat of bunsen burner flame temperature

$$CuCO_3(s) \rightarrow Cu(s) + CO_2(g)$$
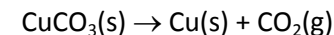
9.3 Extraction of Metals

Metals from Rocks

Minerals – elements/compounds that make up rocks

Metal ore – rock containing metal

Extracting these metals

- Metal ores are removed from ground.
- The ores contain useful and unwanted materials. Unwanted materials are separated to obtain concentrated mineral.
- Metal is extracted from the mineral.

Occurrence of Metals

Metal ores are compounds, usually as:

- **Metal oxides** – *metal + oxygen*, eg: Al_2O_3
- **Metal sulphides** – *metal + sulphur,* eg: HgS
- **Metal carbonates** – *metal + carbon + oxygen*, eg: $MgCO_3$

Some important metal ores:

Mineral Metal	Name of ore	Chemical name	Formula
Sodium	Rock salt	Sodium chloride	NaC*l*
Calcium	Limestone	Calcium carbonate	$CaCO_3$
Magnesium	Magnesite	Magnesium carbonate	$MgCO_3$
Aluminium	Bauxite	Aluminium oxide	Al_2O_3
Zinc	Zinc Blende	Zinc sulphide	ZnS
Iron	Haematite Magnetite	Iron(III) oxide Iron(II),(III) oxide	Fe_2O_3 Fe_3O_4
Tin	Cassiterite	Tin(IV) oxide	SnO_2
Lead	Galena	Lead(II) sulphide	PbS
Copper	Chalcopyrite	Copper(II) sulphide + Iron sulphide	$CuFeS_2$ (CuS + FeS)
Mercury	Cinnabar	Mercury(II) sulphide	HgS

The Extraction of Metals

Easier   Harder

Silver Gold   Mercury Copper Lead Tin Iron Zinc   Aluminium Magnesium Calcium Sodium Potassium

Least Reactive	Less Reactive	Most Reactive
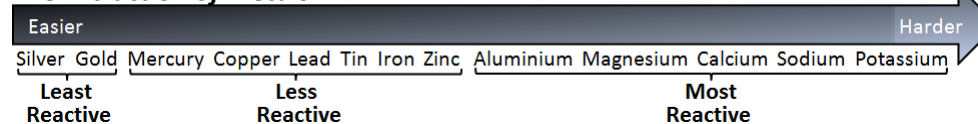
Least Reactive – easiest to extract; extracted by physical methods

Less Rective – harder to extract than least reactive; by blast furnace; usually occur as compounds of oxides or sulphides.

Most Reactive – hardest to extract – strong bonds in compounds; by electrolysis – decomposing compounds with electricity.

Uses of Metals

The choice of metals over another depends on 3 factors:

1. Physical properties (e.g. melting point, strength, density, conductivity)
2. Chemical properties (e.g. resists corrosion)
3. Cost

The Uses of Some Metals and Their Reasons		
Metal	**Uses**	**Reason for the choice**
Aluminium	- Drink cans - Window frames	- Low density, non-toxic, cheap - Resists corrosion, strong
Copper	- Electrical wires - Water pipes	- Ductile, good conductor of electricity - Strong, malleable, resists corrosion
Gold	- Jewellery - Protective coating	- Shiny and attractive, very malleable - Good reflector of heat and light
Titanium	- Supersonic aircraft - Spacecraft	- Light but strong, resists corrosion

Recycling of Metals

How Much is Left?

There are many iron on the surface but copper and tin are seriously reducing.

If you say you have only mined the surface, why don't you mine deeper for more?

High temperatures and pressures and greater depth increases hazards that prevent mining up to the lower part of crust, although there are more metals further down

Conservation of metals- Recycling

- Use alternative materials to replace the use of iron (e.g. use of plastic pipes instead of iron, use of glass bottles for soft drinks instead of aluminium)
- Recycle unused metals by melting them to produce new blocks of clean metal

How aluminium cans are recycled?

Cans are opened and the contents are drunk

Used cans are discarded and collected in recycling bins

Aluminium sheets are made into new cans.

The Recycling Circle

Used cans are crushed into bales

Aluminium blocks are rolled into sheets

Bales are melted in a furnace and cast into blocks
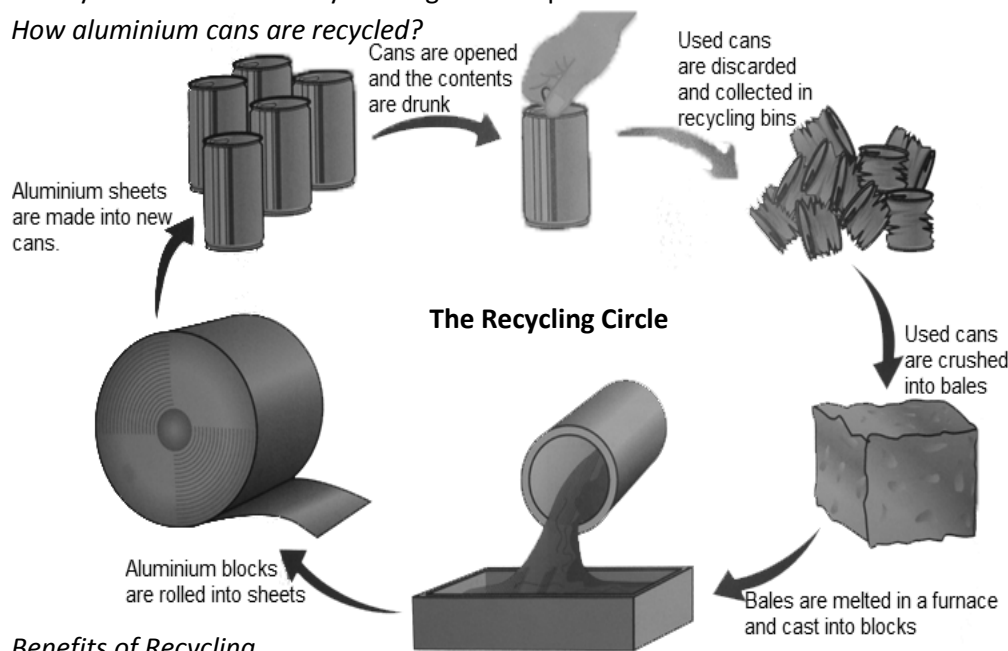
Benefits of Recycling

- Recyling helps conserving metals, especially valuables such as gold and platinum. E.g. used computer parts processed to extract gold used for electrical contacts of processors and memory chips
- Recycling saves the cost of extracting new metals
- Recycling benefits environment, e.g. if there is a car wasteland, it causes eyesore

Problems with Recycling

- Metals are recycle if the cost is cheaper than extraction. E.g. Iron, a cheap metal, is more expensive to recycle than to extract new iron

- Recycling metals can damage the environment by smelting process which sends a lot of fumes into the air
- Cost to separate metals from waste is high. E.g. separat metals in alloys is hard
- Transport costs for collecting scrap metal is high, e.g. trucks should be used
- People are not interested in depositing their used materials in recycling bins

9.4 Iron

Iron is extracted from the iron ore haematite, Fe_2O_3

Iron is extracted from the oxide in a blast furnace (next page)

The Blast Furnace

1. Oxygen in the air reacts with coke to give carbon dioxide:
$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

2. Carbon dioxide produced in 1 reacts with more coke to produce carbon monoxide
$$CO_{2(g)} + C_{(s)} \rightarrow 2CO_{(g)}$$

3. The carbon monoxide reacts with iron(III) oxide to produce molten iron, which runs down to the bottom of the furnace
$$3CO_{(g)} + Fe_2O_{3(s)} \rightarrow 2Fe_{(l)} + 3CO_{2(g)}$$

4. The limestone decomposed by heat to form calcium oxide and carbon dioxide
$$CaCO_{3(s)} \rightarrow CO_{2(g)} + CaO_{(s)}$$

5. Iron ore contains many impurities (silicon, sulphur, phosphorus, etc.) Sand, SiO_2, reacts with calcium oxide to produce slag (calcium silicate). Slag runs down to the bottom of the furnace, floating on top of molten iron

$$CaO_{(s)} + SiO_{(s)} \rightarrow CaSiO_{3(l)}$$
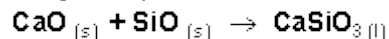
6. Molten iron & slag tapped off separately in furnace. Slag is for road construction.

7. Referring to equation, not all iron(III) oxide reacted with carbon, only small amount

$$Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(l) + 3CO_3 (g)$$
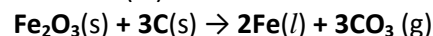
Steel

Iron made from blast furnace is not good as:

- it contains impurities which makes it brittle (can break easily)

- it cannot be bent or stretched

Most iron is converted into steel which is an alloy of iron and carbon with small amounts of other elements. Advantages of steel:

- it is strong and tough

- it can be bent and stretched without shattering

Making Steel:

- Impurities of iron is removed by blowing oxygen into molten iron to change the impurities into oxides. They are then combined with CaO and removed as slag.

- Carbon and other metals are added in certain amount to make steel.

Different Types of Steel:

- Mild steel – is a low carbon steel with 0.25% carbon

  • It is strong and quite malleable. It is used for car bodies, ships, railway lines and steel rods to reinforce concrete

- Hard steel – is a high-carbon steel with about 1% carbon

  • It is harder than mild steel and less malleable. It is used to make tools

- Stainless steel – is iron with large amounts of chromium and nickel

  • It is hard, shiny and doesn't rust. It is used to make cutleries, medical instrument and pipes in chemical industries.

Rusting

Rusting – corrosion of iron and steel

Rust – brown solid product formed during rusting

Rust is hydrated iron(III) oxide $Fe_2O_3 \cdot xH_2O$ where water molecules varies.

Conditions for Rusting
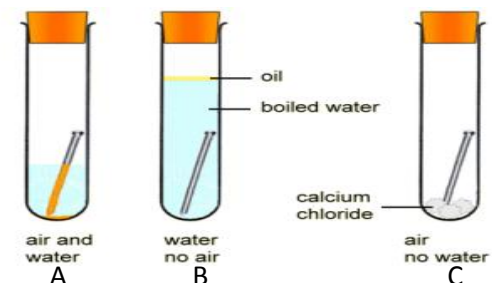
Tubes   A   B   C

After a few days, only nail in tube A rust. This shows that air and water is needed for rust. In boiled water, the nail doesn't rust in B as boiled water removes dissolved air while in C, $CaCl$ keeps air dry so there's no water.

Other factor → dissolved salt

Preventing Rusting

- Surface protection
- Sacrificial protection
- Use of stainless steel

Surface Protection – covers metal with a layer of substance

1) **Paint**

2) **Grease or oil** (also help to lubricate)

3) **Plastic**

4) **Metal Plating** – covering metal with thin layer of another metal (e.g. tin, chromium, silver)

Advantage – These methods are cheap (except metal plating)

Disadvantage – If the layer is broken, air and water an reach metal to rust

Sacrificial Protection

is to sacrifice more reactive metal to corrode with water and air by layering it over less reactive metal (e.g. iron covered by magnesium). If layer is broken, water & air reach underneath layer, overlying metal still protect it.

Applications:

1) **Galvanised Iron** – is steel coated with zinc, usually used on roofs.
2) **Protecting ships** – blocks of zinc are attached to hulls to corrode instead of steel which is the ship metal.
3) **Underground steel pipes** – these are attached to magnesium block using insulated copper cables. Magnesium corrodes first than steel.

END OF CHAPTER 9

CHAPTER 10 – ATMOSPHERE AND ENVIRONMENT

10.1 Air

The atmosphere is a layer of air containing mixture of several gases. This mixture composition varies according to time and place. The composition of water vapour varies from 0-5%, depending on the humidity of air.
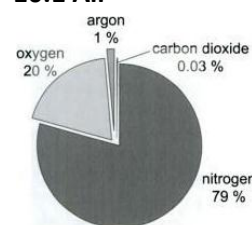
Fig. 23.1 Composition of dry air

Percentage Composition of Oxygen in Air
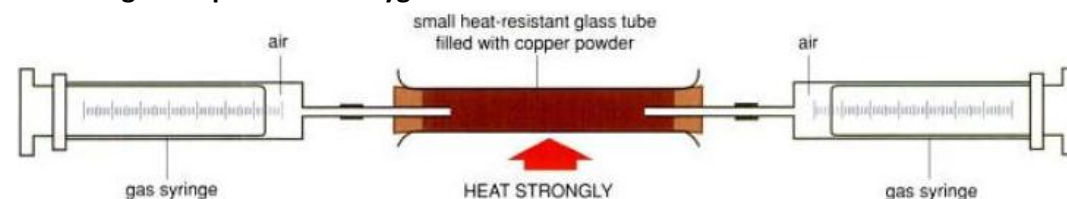
A known volume of air is passed through tube with burning copper powder and oxygen in air will react with hot copper powder to produce black copper oxide:

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$
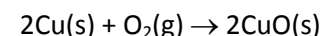
If oxygen is depleted, the readings on both syringes will be steady and the reaction has completed. Hence, to find the volume of oxygen in air collected in syringe:

$$\text{Volume of } O_2 = \text{Initial volume of air} - \text{Final volume of air}$$

For instance, the initial volume of air in one syringe is $80cm^3$ and the final volume is $64cm^3$. Hence, the percentage volume of O_2 in air is:

$$\% \text{ Volume of } O_2 = \frac{\text{Volume of } O_2 \text{ in one syringe}}{\text{Initial volume of gas in one syringe}} \times 100\%$$

$$= \frac{16cm^3}{80cm^3} \times 100\%$$

$$= 20\%$$

Liquefaction of Air

First, CO_2 is removed by passing air through NaOH. Then, the air is cooled to -25^oC to freeze water vapour to be removed and the remaining air is cooled and compressed to become liquid which is then separated into its singular constituents by fractional distillation as each constituent has different boiling point.

OXYGEN – THEY'RE AROUND US!

Oxygen Reaction – Combustion

MOST substances react with O_2 to in exothermic reaction, which is called **combustion**. If flames are produced during combustion, it's called **burning**.

ALL carbon compounds burn in O_2 to produce CO_2 while ALL hydrogen containing compounds burn in O_2 to produce H_2O.

When adequate supply of oxygen is available during burning, it will create a complete combustion. If otherwise, the combustion is **incomplete**.

E.g. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$, makes up a complete combustion

A Test for Combustion

pale blue gas burning completely)

blue (unburnt gas)

yellow flame due to hot specks of unburnt carbon

air-hole opened   air-hole closed

When air hole is closed, air cannot enter supplying oxygen, and hence soot (unburnt carbon) and CO is produced from incomplete hydrocarbon gas combustion. As a result, flame is yellow due to glowing specks of hot soot in heat and the flame is not hot. When air hole is opened, air supplies plenty of oxygen, allowing complete combustion.

Significance of Oxygen

- As rocket fuel
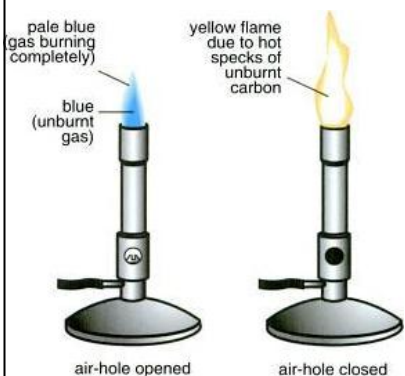- In steel making, to burn off impurities
- In oxy-acetyline cutting and welding
- In oxygen tanks for deep sea divers and mountain climbers to provide oxygen
- For respiration for most animals
- Used as oxygen tents in hospital to aid patients with respiratory problems

Air Pollution

Pollutants are substances in atmosphere which are harmful for living things and environment, for contributing to **air pollution**.

From nature, pollutant sources are volcanoes, forest fires, decay of dead matter, etc. but from humans, they're exhaust fumes, power stations, oil and gas, etc.

The main air pollutants are:

1. Carbon monoxide

Where it comes from?

Unburnt hydrocarbons; exhaust fumes; forest fires

What hazard it brings?

Combines with haemoglobin when inhaled, which produces carboxyhaemoglobin that reduces efficiency of haemoglobin to transport oxygen. Cells then die.

How to prevent this?

- Install catalytic converters in cars
- Reduce number of cars on road
- Create efficient engines in cars to ensure complete hydrocarbon combustion

2. Sulphur dioxide

Where it comes from?

Combustion of fossil fuels containing sulphur impurities; volcanic eruptions

What hazard it brings?

Lung irritant, eye irritant, acid rain

How to prevent this?

- Prevent using fuels containing sulphur impurities, e.g. coal
- Reduce the sulphur impurities inside fossil fuels
- Spray exhaust gases from factories with water/hydrated CaO/alkalis to absorb sulphur dioxide before it's released into the atmosphere
- Add CaO to soil and rivers to neutralize acid rain

3. Oxides of nitrogen (NO, NO_2, ...)

Where it comes from?

Lightning activity; forest fires; internal combustion engines (as nitrogen oxides are formed by oxygen and nitrogen under high temperature); power stations

What hazard it brings?

Eutrophication, lung damage, acid rain

How to prevent this?

- Install catalytic converters in cars
- Design car engines which run at lower temperatures

Chemistry around us: ACID RAIN

Acid rain is formed by 2 main constituents – SO_2 and NO_2

Sulphur dioxide/nitrogen dioxide, both react with oxygen and water to form sulphuric acid/nitric acid. This is called *hydrolysis*.

$$2SO_2(g) + O_2(g) + H_2O(l) \rightarrow 2H_2SO_4(aq)$$
$$4NO_2(g) + O_2(g) + 2H_2O(l) \rightarrow 4HNO_3(aq)$$
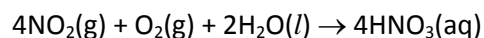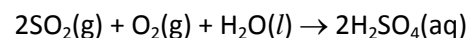
Effects of Acid Rain

- The acid corrodes buildings, $CaCO_3$ materials and metal statues.
- Acid rain damages trees
- Acid rain increases acidity of soil, making soil unsuitable for plant growth
- Fish cannot survive in acidic water
- Aggravates respiratory ailments such as bronchitis and asthma

Tackling Acid Rain

- Remove sulphur dioxide from flue gases by *desulphurization*
- Add $Ca(OH)_2$ to soil to neutralize acid from acid rain
- Burn fuels with less sulphur

Desulphurization

It is the removal of sulphur dioxide from flue (waste) gases. The product is CO_2, which is non-polluting gas, and calcium sulphite.

$$CaCO_3(s) + SO_2(g) \rightarrow CaSO_3(s) + CO_2(g)$$
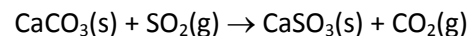
To increase profit, calcium sulphite further oxidized to produced gypsum to be sold

$$2CaSO_3(s) + O_2(g) + 4H_2O(l) \rightarrow CaSO_4.2H_2O(g)$$

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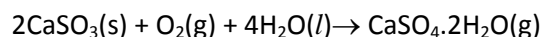
4. Methane

Where it comes from?

Decomposition of vegetable matter; rice field; cattle ranching; natural gas; mines

What hazard it brings?

It is highly flammable, greenhouse gas

How to prevent this?

- Cattle and other ruminant animals should be given improved diet
- Animal manure and rotting vegetation can be used as biomass fuel

5. Unburnt hydrocarbons

Where it comes from?

Internal combustion engunes; incomplete combustion of hydrocarbons

What hazard it brings?

Carcinogenic, forms photochemical smog

How to prevent this?

- Install catalytic converters in cars
- Reduce number of cars on road
- Create efficient engines in cars to ensure complete hydrocarbon combustion

6. Ozone

Where it comes from?

It is an allotrope (two/three different forms of a pure element) of oxygen having structural formula O_3 having characteristic odour. It's created by reaction of nitrogen oxides with volatile organic compounds in presence of UV radiation.

What hazard it brings?

- It reacts with unburnt hydrocarbons to form photochemical smog that causes headache, eye, nose and throat irritation.

- It corrodes and kills plants and trees

How to prevent this?

- Don't use CFCs/replace it with HCFCs which destroys faster.

Chemistry around us: PHOTOCHEMICAL SMOG

It is a hazy brown air, which is a mixture of fog and smog,that reduces visibility, causes eye irritation and breathing difficulties. It is produced by reaction between NO_2 and O_2 in the oresence of sunlight to form NO, O and O_2. This reaction is called *photochemical reaction.*

$$NO_2(g) + 2O_2(g) \rightarrow NO(g) + O + O_2(g)$$

The oxygen atom is reacted with the oxygen molecules formed to form ozone, O_3.

$$O_2(g) + O \rightarrow O_3(g)$$

Ozone can react with unburnt hydrocarbons to produce eye-irritating substances.

Chemistry around us: DESTRUCTION OF OZONE LAYER
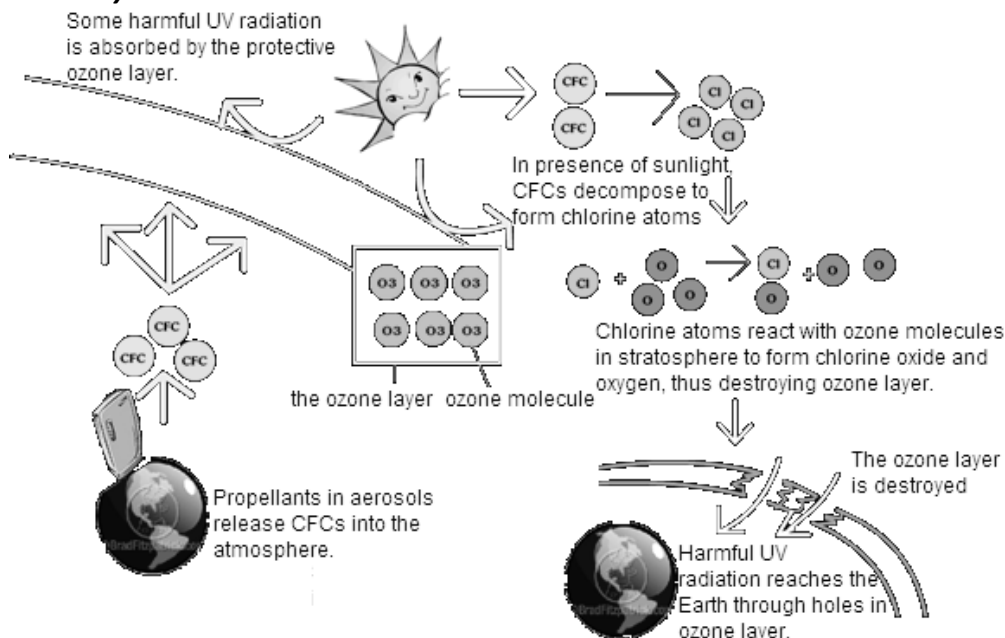
Use of ozone layer in stratosphere:
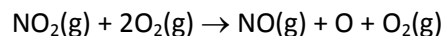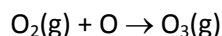
It blocks UV rays from sun which causes skin cancer; acts as blanket to block out high sun energy radiation and prevent it from penetrating into Earth's surface.

How is it depleted?

Ozone layer absorbs some UV radiation and reflects some back to space. CFCs, found in coolants in refrigerator and air conditioners, propellants in aerosols and blowing agents, are released into the atmosphere. In the presence of light, CFC decompose into Cl_2 at the stratosphere where ozone is located.

$$CFCl_3(g) \rightarrow CFCl_2(g) + Cl$$
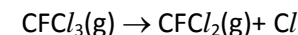
$CFCl_2$ further decomposes to produce more chlorine atoms, Cl, and $CFCl$

$$CFCl_2(g) \rightarrow CFCl(g) + Cl$$
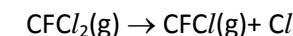
As a result, the Cl atoms produced react with O_3 molecules to form chlorine oxide, ClO, and oxygen, O_2.

$$Cl + O_3(g) \rightarrow ClO(g) + O_2(g)$$
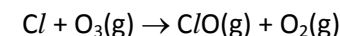
The reaction uses up ozone which covers the earth and hence creating a hole in the layer. Harmful UV rays from sun can now reach Earth through these holes.

Tackling Depletion of ozone layer:

- Don't use CFCs/replace it with HCFCs which destroys faster.

Chemistry around us: GLOBAL WARMING

Greenhouse effect is the trapping of heat from sun by greenhouse gases to regulate Earth temperature so that not all heat is reradiated back to space. However, increased industrialization releases more greenhouse gases to atmosphere, contributing to **Global Warming** (increase in temperature of Earth's atmosphere due to trapping of heat by greenhouse gases).

EXAMPLES OF GREENHOUSE GASES ARE:

1. **Carbon Dioxide** which is naturally occuring or by combustion of hydrocarbons.
2. **Methane** which occur naturally or emitted during production of fuels or from decaying vegetable matter.
3. **Nitrous Oxide** is produced by industrial and agricultural activities,and by incomplete combustion of hydrocarbons

Hazards of Global Warming

It melts polar icebergs, floods low lying areas and coastal regions, alter the climatic conditions of certain places, alter crop yield, and evaporation of water supply.

TACKLING GLOBAL WARMING

- Reduce the use of fossil fuels
- Use alternative forms of energy such as wind, tidal and hydroelectric power
- Use more electric vehicles
- Reduce number of cars on road
- Create efficient engines in cars to ensure complete hydrocarbon combustion

Chemistry around us: CARBON CYCLE

0.03% of the atmosphere is carbon dioxide and this is kept constant by the process *carbon cycle*. **Carbon cycle** is the removal of carbon dioxide by plants by photosynthesis and the replacement of these carbon molecules by combustion, respiration and natural processes. In the past the rate of absorption of carbon dioxide balances the rate of production of carbon dioxide. Man upset this balance.
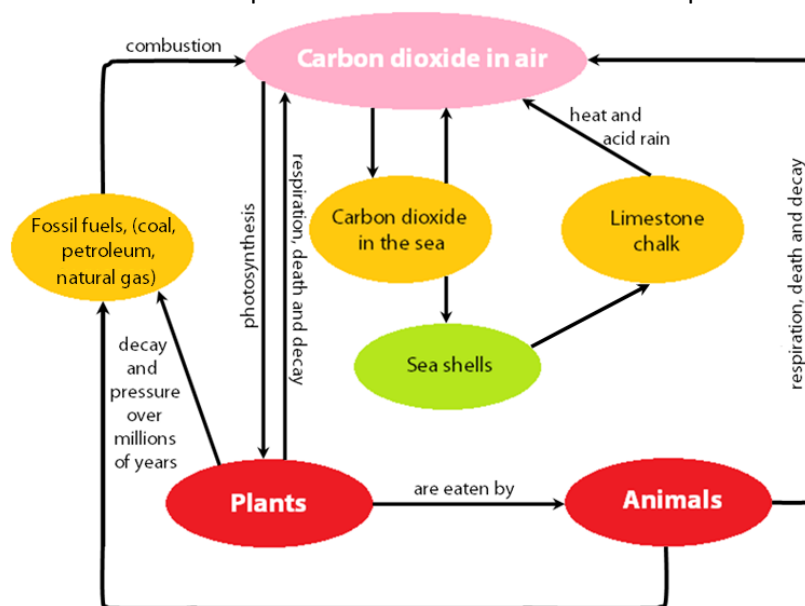
WHAT LIVING THINGS DID

Plants:

Plants use carbon dioxide from atmosphere, sunlight and chlorophyll for photosynthesis of sugars. Some carbon is used up in plants for growth and development, while some others are released to atmosphere during respiration. When plants die & decomposed by microorganisms, CO_2 released to atmosphere.

Animals:

When herbivores and omnivores eat plants, they gain carbon from them to grow and develop. Carnivores eating these animals also gain the carbon. When animals respire, they release carbon dioxide. When they die and decay due to microorganism, they release carbon dioxide which is later taken in by plants.

WHAT NON-LIVING THINGS DID

Carbon monoxide and carbon dioxide are released from electric power plants, exhaust fumes and factory emissions. Man burns fossil fuels, which needs millions of years to form, that takes in oxygen and releases carbon dioxide. This makes man depleting natural resource as they use them rapidly than the time needed to reform, damages natural environment and upsetting balance of carbon cycle.

Balancing Chemicals in Nature: CATALYTIC CONVERTERS
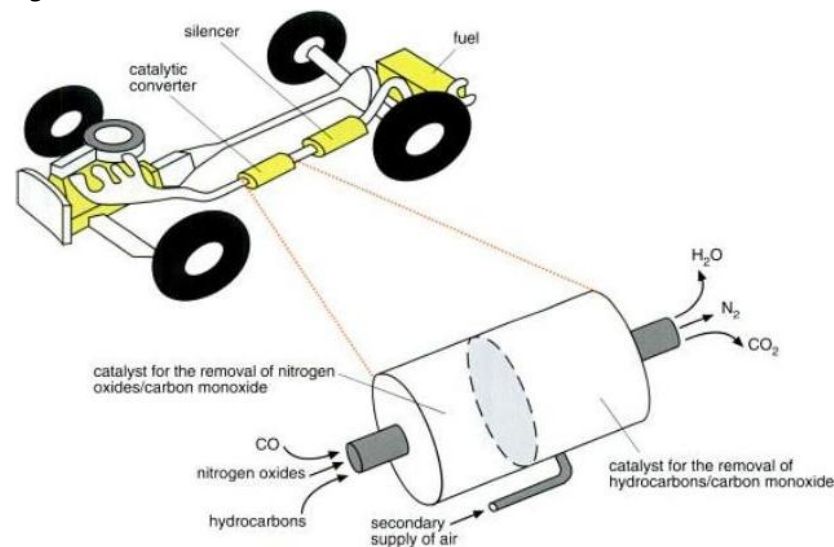
First, nitrogen oxides reacts with carbon monoxide as they pass a platinum catalyst.

$$2NO(g) + CO(g) \rightarrow N_2(g) + 2CO_2(g)$$
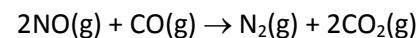
In second half of converter, unburnt hydrocarbons (e.g. octane, C_8H_{18}) reacts with more air to form CO_2 and H_2O.

$$2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$
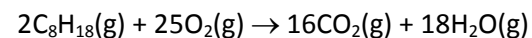
CO_2, H_2O and N_2 are all non-pollutants. These reactions are all redox.

7.Lead compounds

Where it comes from?

- Combustion of leaded petrol in car engines

What hazard it brings?

- Causes lead poisoning which leads to brain damage.

10.2 Water

Water is most available liquid on Earth covering 70% of the planet surface.

We use water at home for: drinking, cooking, washing and bathing

While in industries, we use as: heat exchanger, raw material for food and drinks, as a solvent, cleaning and purification, irrigation, dyeing and bleaching process.

Inside that Water

Naturally Occuring Substances

- Mineral salts – aluminium, calcium, potassium, etc.
- Dissolved oxygen given out by aquatic animals by photosynthesis
- Organic matter (living/dead plants, animals, microorganisms)

Pollutants

- Metal compounds such as cadmium, iron, manganese, etc. from waste discharge
- Phosphates from fertilizers, detergents or sewage treatment plants
- Nitrates from fertilizers or sewage treatment plants
- Sewage from sewage treatment plants or septic systems
- Harmful microbes from sewage treatment plants, septic systems, naturally occuring in water or growing in abundance due to pollution
- Acid from industrial discharges
- Oil spills from oil tankers

Important or Silent Killer?

Beneficial Stuff

Mineral salts

- Needed for basic functions of human body such as bone growth, fluid regulation, normalize nerve and muscle functionality, metabolism control, growth, etc.

- Needed for growth of aquatic plants to make food and produce oxygen for aquatic organisms.

Dissolved oxygen

- Needed for respiration and growth of aquatic life. Water without O_2 is stagnant.

Organic matter

- Needed for growth of aquatic organisms

Harmful Stuff

Acid

- Kills aquatic organisms and plants
- Makes water acidic and corrosive – unsafe to drink

Nitrates

- Causes eutrophication which deprives marine organisms of oxygen
- Nitrate ions may cause breathlessness or kill babies when consumed

Phosphates

- Can cause eutrophication as it encourages the growth of algae, hence killing aquatic organisms when they die and takes away oxygen

Heavy metal ions

- These are carcinogenics that can cause skin cancer, liver cancer, lung cancer, etc.

Sewage

- Contains pathogens which when consumed carries diseases such as diarrhoea.

Oil

- Traps bird's feathers and kills them eventually
- Depletes oxygen as air cannot mix with water to provide sufficient oxygen

Tackling Water Pollution

- Proper disposal of rubbish to prevent more water pollution
- Prohibit activities potentially causing water pollution near rivers/reservoirs such as camping or swimming.
- Dispatch monitor ships to prevent accidents between ships so oil will not spill

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Purification – Taking care of those harmful stuffs
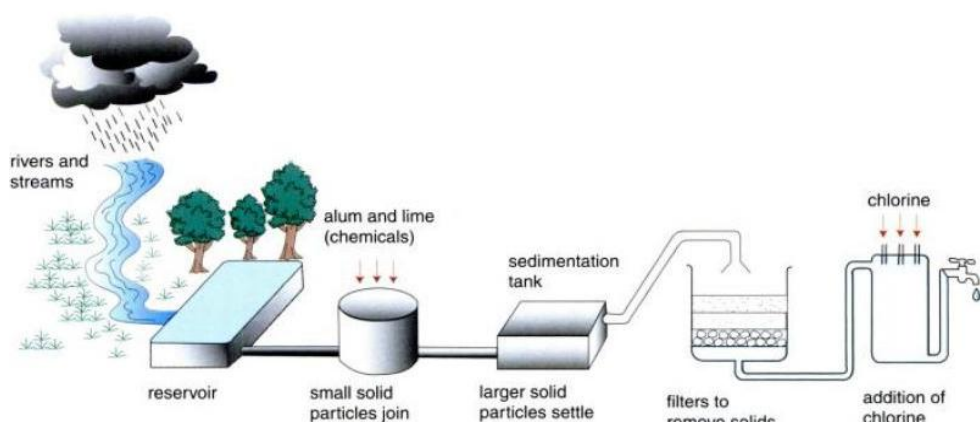
1. Water from rain and river downstream is collected in reservoir.
2. Water is transported via pipe to a flocculation tank where alum, $Al_2(SO_4)_3$, and lime, $Ca(OH)_2$ are added to water so that small solid clay particles join together into large lumps of solid (*coagulation*).
3. Water is moved to sedimentation tank where the lumps of solid settles to the bottom of the tank. This is called *sedimentation*. Carbon, in form of activated charcoal, is added to remove the taste and smell of water.
4. Water is filtered off in filtration tank, where there are sand particles filter which traps the remaining solid particles in water.
5. Chlorine and fluorine are added in chlorination tank. This is called *chlorination*. Chlorine is used to kill bacteria while fluorine is used to strengthen teeth.

Desalination

Ocean is vast source of water. Salts in seawater must be removed so it's drinkable.
Desalination is the process where seawater is distilled until it becomes steam (free of salt) which is then cooled and condensed into drinking water.

END OF CHAPTER 10

CHAPTER 11 – ORGANIC CHEMISTRY
Compounds from Living Things
Organic Compounds – compounds found in living organisms
Examples: sugar, fats, plant oils, urea

Charecteristics of Organic Compounds
- All contain carbon element
- Most come with hydrogen
- Others with oxygen, nitrogen, or a halogen

Uses of Organic Compounds
Fuels, plastic, rubber, detergents, insecticides, most medicines

Classifying Organic Compounds
Homologeous series – a family of organic compounds with a general formula and a similar chemical properties
All homologous series has:

Functional Groups
- Is the special group of atoms available in homologous series compunds which responsibles for the chemical properties of the compound

- All compounds in homologous series have functional group except alkanes.
  Examples of functional group homologous series: *alcohol*

Production of Organic Compounds

From crude oil refinery:

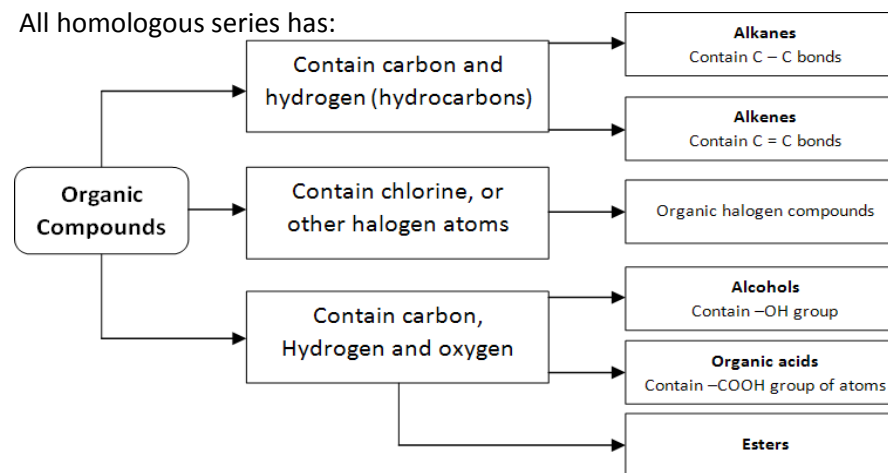Crude oil is a mixture of complex hydrocarbons with varying boiling points, depending on the number of carbon atoms and how they are arranged. Fractional distillation uses this property to separate the hydrocarbons in crude oil.

From naphtha:

Naphtha fraction is used for production of petrochemicals, such as medicines, plastics and synthetic fibres, aside from fuels. When naphtha is treated, not only it becomes a better fuel, it also contain more aromatic hydrocarbons, alkene and cyclic hydrocarbons which are important for petrochemical industry.

Crude oil is mostly used as fuel, though some allocated for chemical feedstock. As oil reserves deplete, competition between 2 main uses of oil will be more intense.

Saturated or Unsaturated?

Saturated hydrocarbons are hydrocarbons which the combining capacity of the carbon atoms is as fully used as possible in bonding with hydrogen atoms. They only have single bond (–) only.

Unsaturated hydrocarbons are hydrocarbons which the combining capacity of the carbon atoms is not fully used, e.g. only 2 or 3 hydrogen are attached to a carbon atom. This is usually indicated by double bond (=) or triple bond (≡) with another carbon atoms.

11.1 Alkanes

- Usually in fuels, **examples:** *natural gas, petrol, diesel*
- Are homologous series
- Have a formula of C_nH_{2n+2}

- **example:** propane has three carbon atom, thus *n*=3. Then the formula of propane is C_3H_8
- Ends with suffix **–ane**
- Next alkane formula differ by $-CH_2$ atoms. Eg: methane: CH_4, ethane: C_2H_6

Structure of Alkanes

Shows how all atoms in a molecule joined together by drawing lines between atoms to represent the bonds

Example: butane has a formula of C_4H_{10}, therefore the structural formula is:
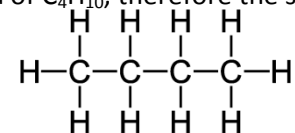
It has 4 carbon atoms bonded together with 10 hydrogen atoms

Organic compound containing only single bond is **saturated**. Eg: *methane*
All alkanes are saturated. All alkenes are unsaturated

Physical Properties of Alkanes

Name	Molecular formula	M.p. (°C)	B.p. (°C)	State at r.t.p.*	Density (g/cm³)
methane	CH_4	−182	−162	gas	0.000 68
ethane	C_2H_6	−172	−89		0.001 27
propane	C_3H_8	−187	−42		0.001 86
butane	C_4H_{10}	−138	−0.5		0.002 45
pentane	C_5H_{12}	−130	36		0.63
hexane	C_6H_{14}	−95	69	liquid	0.66
⋮	⋮	⋮	⋮		⋮
heptadecane	$C_{17}H_{36}$	22	292		0.77
octadecane	$C_{18}H_{38}$	28	308	solid	0.78
nonadecane	$C_{19}H_{40}$	32	320		0.78

From the table,

- **Melting points and boiling points increase** as the bonds become larger and heavier which increases the forces of attraction between molecules so more energy (from heat) is needed to separate them with the increase of strength of forces of attraction

- Alkanes are **insoluble** in water but **soluble** in organic solvents such as tetrachloromethane as alkanes are organic compounds

- Alkane density **increases** down the series; all alkenes are less than 1g/cm^3

- Alkanes become more **viscous** (uneasily flow) going down the series as the longer molecules tangles together when it flows

- Alkanes become **less flammable** down the series as B.P. becomes larger

- Alkanes are **unreactive** with either metals, water, acids or bases because the C – C and C – H covalent bonds are harder to break

Reaction of Alkanes

Have same chemical properties – they don't react with most chemicals

Combustion

Alkanes burn in air to ALWAYS form carbon dioxide and water. When there is insufficient oxygen, the product is ALWAYS carbon monoxide and unburnt carbon.

Example: Butane is commonly used camping gas. State the chemical equation of combustion of butane in air.

$$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O\ (l)$$
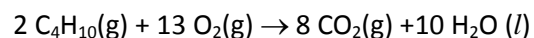
High alkanes burn less completely and gives soot (unburnt carbon) and CO

Reaction with Chlorine/Other Halogens (Alkyl Halides)

Chlorine molecule replaces alkane hydrogen atom with chlorine atom

Substitution reaction – the reaction in which one or more atoms replace other atoms in a molecule

Light is needed to break covalent bond between chlorine molecule→atoms

methane   chlorine   dichloroethane
            molecule
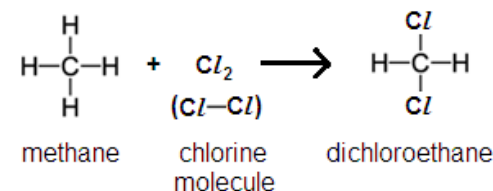
11.2 Alkenes

- have general formula **C_nH_{2n}**.
- all alkene names end with **–ene**.
- the formula of one alkene differs from the next by $-CH_2$.
- have similar properties like alkane going down the series.

No. of C atoms	Name	Molecular formula	Structural formula	Condensed structural formula
2	ethene	C_2H_4		$CH_2 = CH_2$
3	propene	C_3H_6		$CH_3CH_2 = CH_2$
4	butene	C_4H_8		$CH_3CH_2CH = CH_2$

Table 25.3 First three alkenes; they appear as gas

Structure of Alkenes

Is organic compound containing C = C double bond, said to be **unsaturated**

Reason: not all C atoms are bonded to the maximum no. of 4 other atoms

The Importance of Ethene

Ethanol – solvent & fuel   **Poly(ethene)** – PE plastic variations

Ethanoic acid – vinegar

Reactions of Alkenes

Combustion

Burns in air to form carbon dioxide and water

Example: Ethene burns in air. Write the balanced equation for the reaction

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O~(l)$$
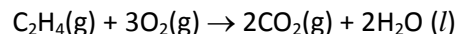
Incomplete combustion forms soot and CO. It's produced more than alkane

Addition Reaction

Is the reaction of 2 or more molecules to form a single product

- *Addition of hydrogen*

Alkenes react with hydrogen to form alkanes, called **hydrogenation**. Must use nickel as catalyst and heat.

$$C_2H_4(g) \qquad + \qquad H_2(g) \qquad \rightarrow \qquad C_2H_6(g)$$

Nomenclature

Product's an ALKANE with name according to number of carbon atoms it contain.

- *Addition of bromine*

Bromine adds to C = C double bond of alkane molecules. Phosphoric acid (H_3PO_4), high temperature of 300^oC and 60-70 atm pressure are needed as catalyst.
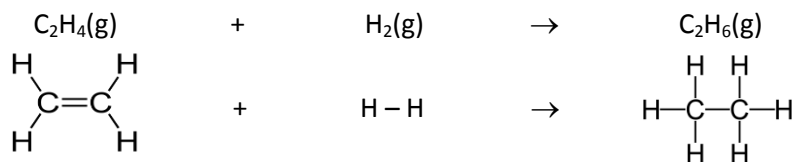
Eg: ethene to 1,2 – dibromoethene

$$C_2H_4(g) \qquad + \qquad Br_2(g) \qquad \rightarrow \qquad C_2H_4Br_2(l)$$
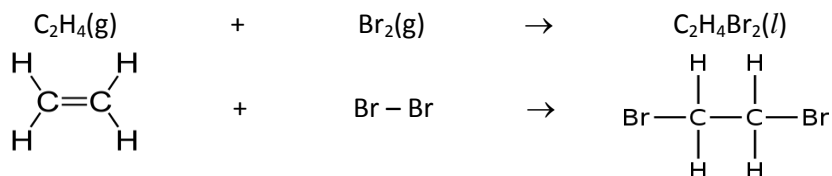
Nomenclature

(n) + (bromo) + (alkene name), where n is the number of bromine atoms.

E.g. Above, Ethene reacts with 2 bromine atoms producing DI(2)BROMO(Bromine) ETHENE(alkene name). Hence we call the product DIBROMOETHENE.

- *Addition of water*

Alkene reacts with water, in the form of steam, to produce alcohol. Alkene + steam is passed over phosphoric acid (H_3PO_4) catalyst and temperature of 300^oC. H_2O molecule adds to C = C bonds to form alcohol.

$$C_2H_4(g) \qquad + \qquad H_2O(g) \qquad \rightarrow \qquad C_2H_5OH \text{ or } CH_3CH_2OH~(l)$$
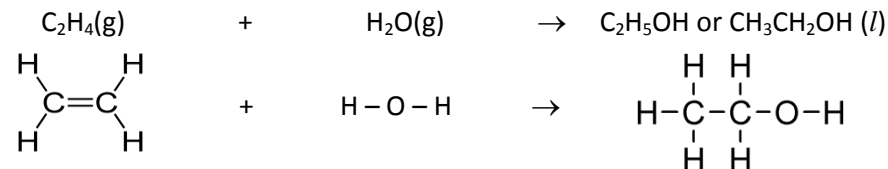
Nomenclature

(alkene name) + (-ol)

E.g. in above, the alkene ethane (C_2H_4) reacts with steam to form *ETHANOL* (alkene name – ETHAN + OL group of alcohol)

- *Polymerization*

The joining of several identical alkene molecules to form a big molecule

Eg: Ethene → poly(ethene)

Testing for Unsaturated Compounds

Mix bromine solution with alkene (for liquid alkenes – shake). Reddish-brown colour of bromine disappears. This shows that the compound is an alkene.

Characteristics of a Homologous Series

- All members of homologous series have same general formula
- Formula of each member differs by –CH_2 group
- Physical properties changes gradually in the increase of carbon atoms
- The members have similar chemical properties

Foods and Unsaturated Compounds

The Manufacture of Margarine

Polyunsaturated food – food containing C=C bond in their molecules

Eg: Vegetable oil

To produce margarine:

part of a vegetable oil molecule → part of a margarine molecule

Hydrogen is reacted with vegetable oil with presence of nickel catalyst and heat, which adds to C=C bond, increasing the molecular mass of the compound

- With increase in mass, the compound has higher boiling point. Therefore, margarine is solid at room temperature.
- Since only some C=C bonds react with hydrogen, margarine is partially hydrogenated and each has different hardness, depending on the number of C=C bonds.

The Cracking of Alkanes

Alkanes can be cracked into shorter chain hydrocarbons because of the higher value it has that it can create more variety of products in petrochemical industries.

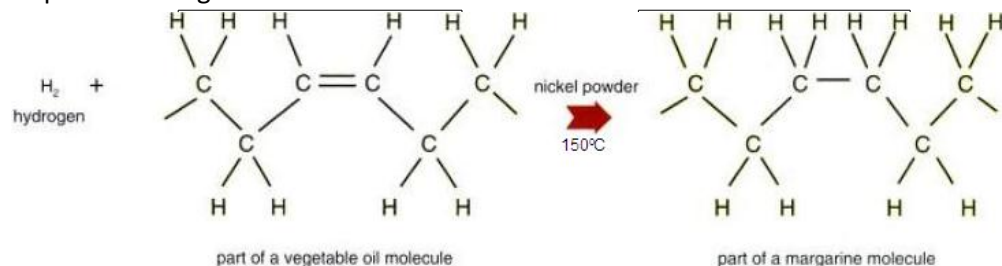We crack alkane by catalytic cracking, which is, using catalyst to break alkane into simpler hydrocarbons. We crack alkane to get more valuable hydrocarbons. The total number of carbon and hydrogen atoms from products should equal to the total number of carbon and hydrogen atoms in cracked alkane.

E.g. Octane can be cracked into simpler hydrocarbons such as the reaction below.
   Suggest the possible identity of product x.

$$C_8H_{18}(l) \rightarrow C_2H_4(g) + x + CH_4(g)$$
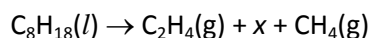
Number of C atoms in $x = 8 - 2 - 1$
$$= 5$$
Number of H atoms in $x = 18 - 4 - 4$
$$= 10$$
∴ Product x is C_5H_{10}

Isomers
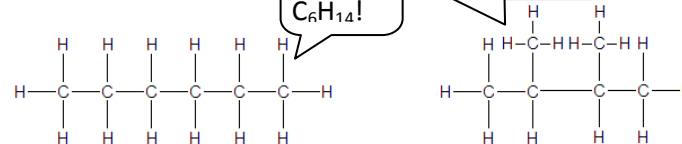
Figure 11 Isomerism

Look at the figure above and count the number of carbon and hydrogen atoms in each case. You will end up with the same C_6H_{14}. We can't deny that they have the same molecular formula. However, their structures are different. Therefore:

Isomers are compounds with same molecular formula but different structural formula. Due to different chain length, they have different physical properties (e.g. boiling point). Isomerism can occur in both alkanes and alkenes.

We therefore can't just say that C_6H_{14} is simply hexane because there are more variations of C_6H_{14} and each variation has its own name. The figure below shows the nomenclature (i.e. how to name) these isomers.

ISOMERISM INVOLVING METHYL GROUP IN ALKANE
Example: Draw and name 3 of the isomers of C_6H_{14}, with at least one methyl group in it.

ISOMER 1: hexane

CH_3—CH_2—CH_2—CH_2—CH_2—CH_3
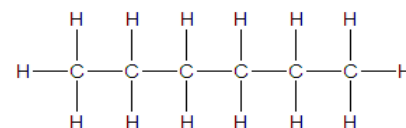
Figure 12(a)

Closest position of first methyl group relative to isomer end
ISOMER 2: 2, 2-dimethylbutane → no. of methyl groups
closest position of second methyl group relative to isomer end

2, 2-dimethylbutane

methyl group = CH_3

CAN ALSO BE WRITTEN AS

CH_3
|
CH_3C—CH_2—CH_3
|
CH_3

Figure 12(b)(i)

CH_3
|
CH_3C—CH
| |
CH_3 CH_3

Figure 12(b)(ii)

ISOMER 3: 2, 3–dimethylbutane → no. of carbon atoms in the straight chain

CAN ALSO BE WRITTEN AS

Figure 12(c)(i)

Figure 12(c)(ii)

Figure 12

Figure 12(a) is the full long chained isomer of C_6H_{14}: hexane. This is so not useful.

Figure 12(b)(i) is another isomer called 2,2-dimethylbutane.

- Note that the first number 2 indicates the position of methyl group (CH_3) attached to a carbon atom from the nearest end. There are 2 possible numbers: 2 or 3. Since 2 is closer, we put 2 in place. The second number 2 indicates the position of the second methyl group attached to carbon atom. Since it's attached to the same carbon atom as the first methyl group, we put the same number 2.

- Also note that the name is now "butane". This comes from the number of carbon atoms in the **STRAIGHT** chain only. The turns leading to methyl is ignored.
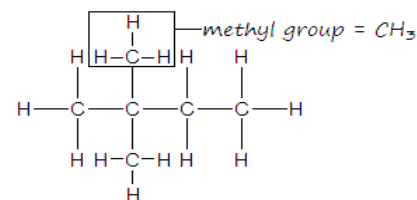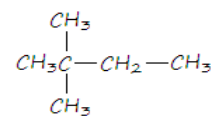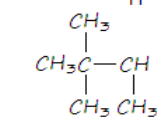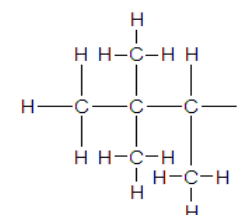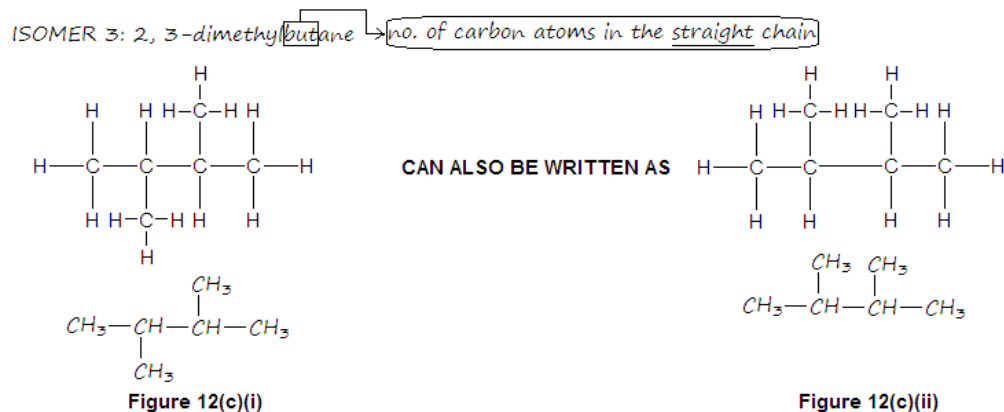
- Bear in mind that "di" in "dimethyl" indicates the number of methyl groups in the isomer ("di" means two). One methyl has no prefix, if it's three is "tri" and so on.

Figure 12(b)(ii) is another variation of the isomer 2,2-dimethylbutane

- Students often misinterpret this as 1,2,2-trimethylpropane while in fact, we don't take the last bend in the chain as another methyl group. Instead, we consider it as **PART OF THE STRAIGHT CHAIN**.

Figure 12(c)(i) and **Figure 12(b)(ii)** is another isomer called 2,3-dimethylbutane

- See that we can flip the positions of methyl group without changing formula

- Note that the second number is 3 while in fact, the position closest to the end of isomer is 2. This is to avoid confusion that the isomer would be that of **figure 12(b)(i)** and **figure 12(b)(ii)**. In this case, we put the number as the **POSITION THE SECOND NEAREST TO THE END OF THE ISOMER**, that is, 3.

ISOMERISM INVOLVING ALKENE

Involves C=C double bond:

E.g. Draw and name 3 of the isomers of C_7H_{14}

ISOMER 1: hept-1-ene

Figure 13(a)

$CH_3CH_2CH_2CH_2CH=CH_2$
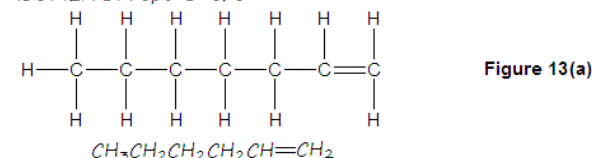
ISOMER 2: hept-2-ene → closest position no. of double bond relative to isomer end

Figure 13(b)

$CH_3CH_2CH_2CH_2CH=CHCH_3$
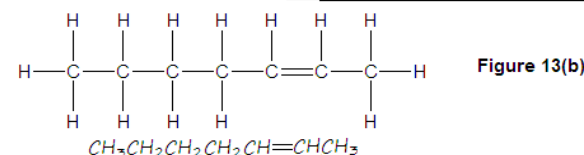
closest position of first methyl group relative to isomer end → no. of methyl groups in isomer

ISOMER 3: 2, 3-dimethylpent-1-ene

closest position of second methyl group relative to isomer end → no. of carbon atoms in the straight chain

Figure 13(c)

$CH_3CH_2CHC=CH_2$

CH_3
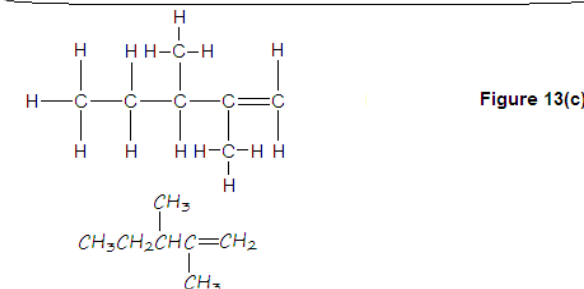
Figure 13

For isomerism in alkene in **Figure 13(c)**, we apply the same theory as isomerism in alkane, and just to only add a double bond indication.

For alkene, double bond position can be changed. In **Figure 13(a)**, it's hept-1-ene, can be called heptene, can be changed to hept-2-ene in **Figure 13(b)**, where the number in between indicates position of double bond from nearest isomer end.

11.3 Alcohols

Are homologous series with general formula $C_nH_{2n+1}OH$

- They have –OH functional group (**hydroxyl group**)
- All alcohols end with suffix **-ol**

First three members of the series (so that you'd have idea on the next)

- Methanol, CH_3OH
- Ethanol, C_2H_5OH or CH_3CH_2OH
- Propanol, C_3H_7OH or $CH_3CH_2CH_2OH$

For alcohol, the –OH is not of hydroxide ion, OH^-, but is covalent bond between oxygen and hydrogen, $O - H$

Making Ethanol

- Fermentation of sugars with yeast
- Reacting ethene with steam

Fermenting glucose

Fermentation is breakdown of sugars into smaller molecules by microorganisms.

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$
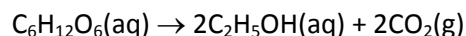
Temperature is kept constant at 37^oC to prevent destruction of yeast at higher temperatures. Oxygen is removed by limewater and carbon dioxide is produced during fermentation. Alcohol is separated from solution by fractional distillation.

Reacting Ethene with Steam

Ethene and steam are passed over phosphoric acid H_3PO_4 (as a catalyst) under high temperature of 300^oC and pressure of 65 atm.

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(aq)$$
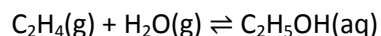
Since this is reversible reaction, both ethene and water are produced aside from ethanol. The ethanol is separated by fractional distillation.

Uses of Alcohol

As organic solvent; alcoholic drink; preservatives; vehicle fuel

REACTIONS OF ALCOHOL

Combustion

Alcohols burn in air to produce carbon dioxide and water.

E.g. combustion of ethanol

$$C_2H_5OH(aq) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$
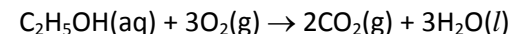
Oxidation

1. Alcohol can be prepared in laboratory by warming alcohol with oxidizing agent (e.g. acidified potassium chromate(VI)). The product is carboxylic acid and water.

E.g. oxidation of ethanol produces water and ethanoic acid

$$C_2H_5OH(aq) + 2[O]\{\text{from oxidizing agent}\} \rightarrow 2CH_3COOH(g) + 3H_2O(l)$$
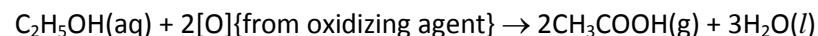
2. Alcohol can be oxidized when left in air with bacterial enzymes as catalyst. The products are carboxylic acid and water.

E.g. ethanol produces water and ethanoic acid when left in air.

$$C_2H_5OH(aq) + O_2(g) \rightarrow 2CH_3COOH(aq) + 3H_2O(l)$$
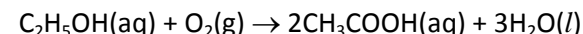
Esterification

This will be discussed in Chapter 11.4

11.4 Carboxylic Acids

homologous series with general formula $C_nH_{2n+1}COOH$ (first serie, **n** = 0, ascending)

- They have –COOH functional group (**carboxyl group**)
- All carboxylic acids end with suffix **–oic acid**

First three members of the series (so that you'd have idea on the next)

- Methanoic acid, HCOOH
- Ethanoic acid, CH_3COOH
- Propanoic acid, C_2H_5COOH

BUTANOL (ALCOHOL)   PROPANOIC ACID (CARBOXYLIC ACID)
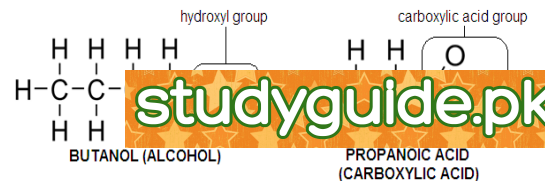
PREPARATION OF CARBOXYLIC ACIDS

1. From natural gas

Natural gas is passed over air and catalyst to form ethanoic acid and water.

E.g. production of ethanoic acid from methane

$$2CH_4(g) + 2O_2(g) \rightarrow CH_3COOH(aq) + 2H_2O(l)$$

2. Oxidation (explained in **Chapter 11.3**)

PROPERATIES OF CARBOXYLIC ACIDS

- **Carboxylic acids are weak acids** (partially ionises in water)
- **Carboxylic acids react with metals to form metal ethanoate** (salt) **and hydrogen**

E.g. Reaction between calcium and ethanoic acid forming calcium ethanoate and hydrogen

$$Ca(s) + 2CH_3COOH(aq) \rightarrow Ca(CH_3COO)_2(aq) + H_2(g)$$
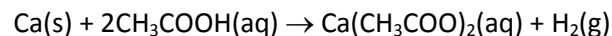
- **Carboxylic acids react with bases to form salt and water (neutralization)**

E.g. Ethanoic acid reacts with sodium hydroxide to form sodium ethanoate and water.

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(g)$$
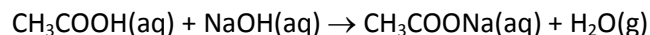
- **Carboxylic acids react with carbonates and bicarbonates to form salt, carbon dioxide and hydrogen.**

E.g. Ethanoic acid reacts with sodium carbonate to form sodium ethanoate and water.

$$2CH_3COOH(aq) + Na_2CO_3(aq) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O(g)$$
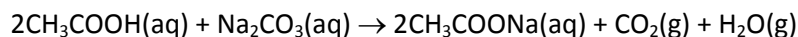
ESTERIFICATION

Ester is organic compound made from carboxylic acid and alcohol with the removal of one molecule of water. **Sulfuric acid is added as catalyst** then heat mixture.

The next page shows the reaction between an alcohol and carboxylic acid.

The reaction is **reversible**. We can **add sodium hydroxide** and heat mixture to obtain carboxylic acid and alcohol from ester. This is **HYDROLYSIS**.
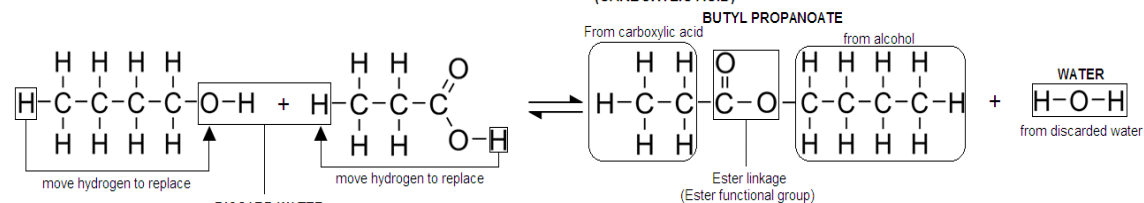
ESTER NOMENCLATURE:

Ester name is [alcohol]**yl** [carboxylic acid]**oate**. For instance, example above is *butyl propanoate*, where "butyl" is from butanol; "propanoate" is from propanoic acid.

11.5 Synthetic Macromolecules

Macromolecule is a large molecule made by joining together many small molecules

Polymer is a long-chain macromolecule made by joining together many monomers

Polymerisation is the addition of monomers to make one large polymer

ADDITION POLYMERISATION

Addition polymerisation is which small molecules (monomers) join together to form one molecule as the only product.

From monomer to polymer

Example: Formation of poly(ethene) from ethene

Ethene has double bond. Another ethene molecules add to this unsaturated compound during polymerisation to form bigger compound.
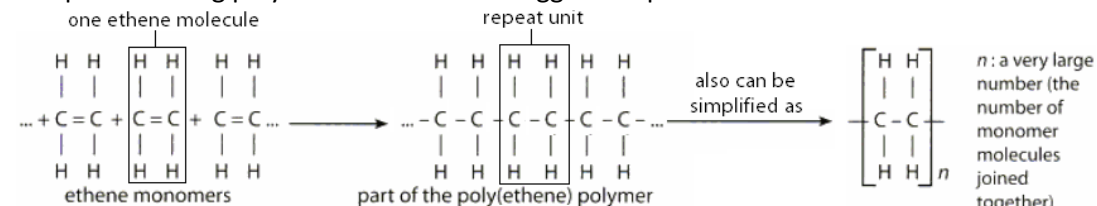
Repeat unit is the simplest part of the polymer which is repeated many times to form the polymer. We take the repeat unit to write the simplified formula of the polymer, where *n* is a large number. From this repeat unit, to find the monomer formula, we add double bond between C – C and remove the bonds on each of their sides.

Some plastic variations and their uses are shown:

Addition polymers	Uses
Poly(ethene), commonly called polythene	plastic film for wrapping fresh vegetables and meat; plastic bottles; plastic bags; buckets; pipes
Poly(chloroethene), commonly called PVC	for water pipes; water-proof plastic sheeting; insulating for electric cables and wires
Poly(phenylethene), commonly called polystyrene	for packaging and containers for electrical goods; disposable cups and food containers used in ' fast food' restaurants and hawker centres
Perspex, (this is the common name)	used for plastic windows and to make models of machines and buildings as it is transparent and does not break as easily as glass
Poly(tetrafluoroethene), commonly called PTFE	used to coat 'non-stick' pots and pans as it is chemically inert, heat-resistant and repels water

CONDENSATION POLYMERISATION

Condensation Polymerisation is the joining of monomers together to form polymers along with the elimination of water molecules.

Nylon

Dicarboxylic acid and diamine undergo condensation polymerisation to form nylon.

The linkage between monomers in nylon is called **amide linkage**. Therefore we can also call nylon as **polyamide**.

Today, we use nylon as:
- a replacement of stockings and manufacture of garments to replace silk
- make tents and parachutes due to strength
- fishing lines
- rugs and carpets

Terylene

Dicarboxylic acid (acid with 2 –COOH groups) and diol (alcohol with 2 –OH groups) undergo condensation polymerisation to form terylene

The linkage between the monomers in terylene is called **ester linkage**. Therefore we can call this polymer as **polyester**.

Today, we use terylene in fabrics as it's strong, resists stretching and sinking and doesn't crumple when washed.

PROBLEMS ASSOCIATED WITH PLASTICS

- Plastics are non-biodegradable – they cannot be decomposed by bacteria. Therefore, many plastic waste will pollute the Earth
- Plastics produce toxic gas (such as hydrogen chloride) when burnt and this contributes to acid rain.
- Plastics produce carbon dioxide when burnt – increases global warming.
- Plastics that require CFC during production may contribute to global warming when the CFC is allowed to escape.

11.6 Natural Macromolecules

CARBOHYDRATES

Carbohydrates contain **carbon**, **hydrogen** & **oxygen**. General formula is $C_n(H_2O)_n$. The simplest carbohydrate is $C_6H_{12}O_6$ (glucose).

Glucose polymerise each other to form **starch**.
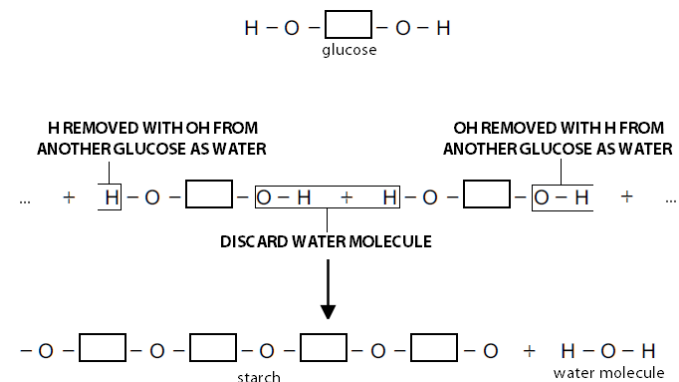
The overall reaction is:

$nC_6H_{12}O_6 \rightarrow (C_5H_{10}O_5)_n + nH_2O$

Starch can also be broken down into glucose by heating with **sulfuric acid**. This is **HYDROLYSIS**.

PROTEINS

Proteins have similar linkage to that of a nylon. Only that their monomers are only **amino acids** joined together. They are formed by condensation polymerisation.

Proteins can be called as **polyamide** as it has **amide linkage**. Proteins can also be broken down into amino acids by **boiling protein with sulfuric acid**. This adds water molecule into the polymer.

Amino acid

H REMOVED WITH OTHER OH FROM ANOTHER AMINO ACID AS WATER

OH REMOVED WITH OTHER H FROM ANOTHER AMINO ACID AS WATER

DISCARD WATER MOLECULE

water molecule

amide linkage

PROTEIN REPEAT UNIT

PROTEIN
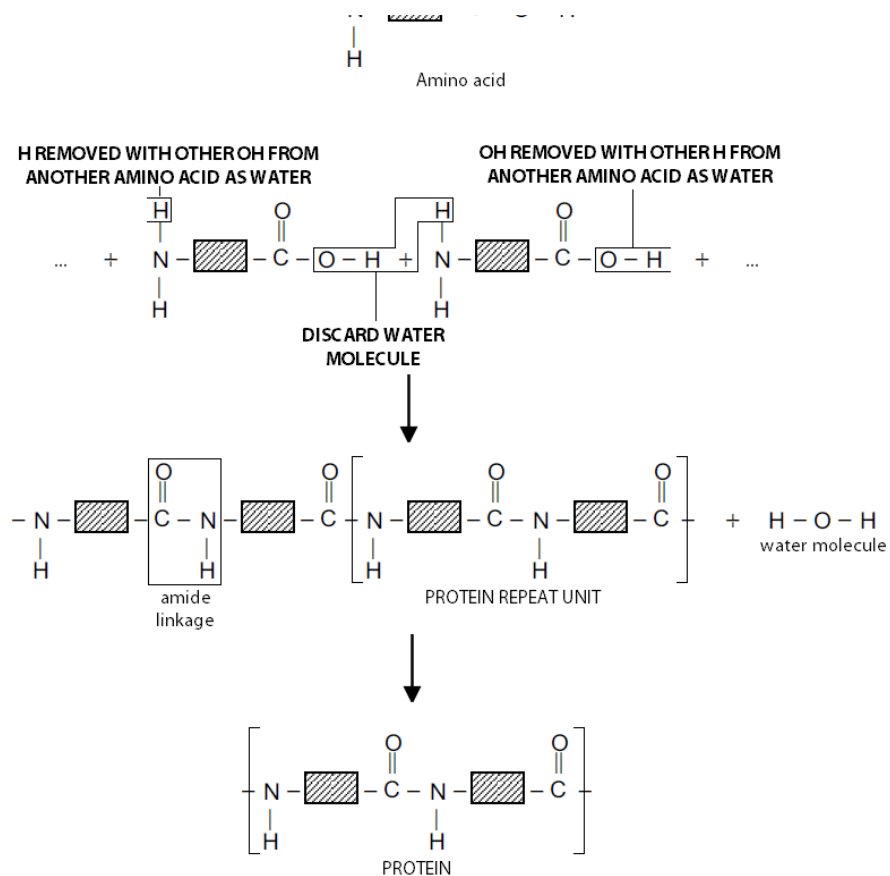
FATS

Fats have similar linkage to that of a terylene (**ester linkage**). Only that their monomers consists of glycerol and fatty acids; different from terylene.

Fats can also be broken down to sodium salts of fatty acids and glycerol by **boiling it with an acid or alkali**. This is **HYDROLYSIS**.

The reactions are given on the page on the right.

POLYMERISATION OF FATS

fatty acid

glycerol

ester linkage

DISCARD WATER

water molecule

Fatty acids + Glycerol ⟶ Fats + water

HYDROLYSIS OF FATS

+ 3 NaOH ⟶

Fats + Sodium hydroxide ⟶ Sodium salts of fatty acids + Glycerol
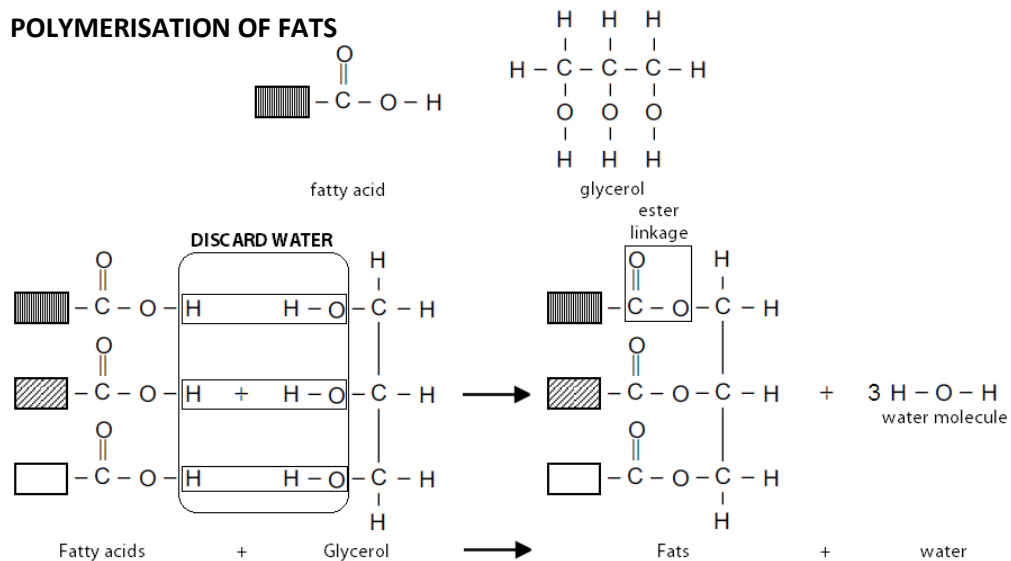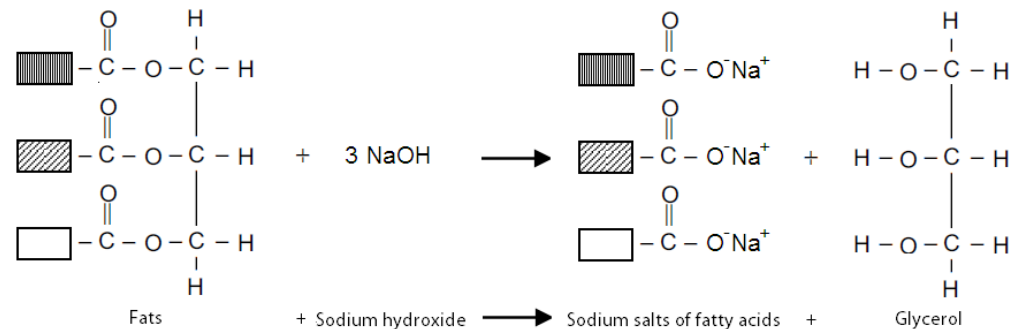
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GOOD LUCK IN YOUR EXAMINATIONS!

INSERT I

Tests for anions

anion	test	test result
carbonate (CO_3^{2-})	add dilute acid	effervescence, carbon dioxide produced
chloride (Cl^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
iodide (I^-) [in solution]	acidify with dilute nitric acid, then add aqueous lead(II) nitrate	yellow ppt.
nitrate (NO_3^-) [in solution]	add aqueous sodium hydroxide then aluminium foil; warm carefully	ammonia produced
sulfate (SO_4^{2-}) [in solution]	acidify with dilute nitric acid then add aqueous barium nitrate	white ppt.

Tests for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium (Al^{3+})	white ppt., soluble in excess giving a colourless solution	white ppt., insoluble in excess
ammonium (NH_4^+)	ammonia produced on warming	–
calcium (Ca^{2+})	white ppt., insoluble in excess	no ppt. or very slight white ppt.
copper(II) (Cu^{2+})	light blue ppt., insoluble in excess	light blue ppt., soluble in excess giving a dark blue solution
iron(II) (Fe^{2+})	green ppt., insoluble in excess	green ppt., insoluble in excess
iron(III) (Fe^{3+})	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc (Zn^{2+})	white ppt., soluble in excess giving a colourless solution	white ppt., soluble in excess giving a colourless solution

Tests for gases

gas	test and test result
ammonia (NH_3)	turns damp red litmus paper blue
carbon dioxide (CO_2)	turns limewater milky
chlorine (Cl_2)	bleaches damp litmus paper
hydrogen (H_2)	"pops" with a lighted splint
oxygen (O_2)	relights a glowing splint
sulfur dioxide (SO_2)	turns acidified aqueous potassium dichromate(VI) from orange to green
water vapour (H_2O)	turns blue cobalt(II) chloride paper pink

DATA SHEET
The Periodic Table of the Elements

INSERT II

Group																	
I	II												III	IV	V	VI	VII	0

1 **H** Hydrogen 1		4 **He** Helium 2

7 **Li** Lithium 3	9 **Be** Beryllium 4		11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10

| 23 **Na** Sodium 11 | 24 **Mg** Magnesium 12 | | | | | | | | | | | | 27 **Al** Aluminium 13 | 28 **Si** Silicon 14 | 31 **P** Phosphorus 15 | 32 **S** Sulfur 16 | 35.5 **Cl** Chlorine 17 | 40 **Ar** Argon 18 |

| 39 **K** Potassium 19 | 40 **Ca** Calcium 20 | 45 **Sc** Scandium 21 | 48 **Ti** Titanium 22 | 51 **V** Vanadium 23 | 52 **Cr** Chromium 24 | 55 **Mn** Manganese 25 | 56 **Fe** Iron 26 | 59 **Co** Cobalt 27 | 59 **Ni** Nickel 28 | 64 **Cu** Copper 29 | 65 **Zn** Zinc 30 | 70 **Ga** Gallium 31 | 73 **Ge** Germanium 32 | 75 **As** Arsenic 33 | 79 **Se** Selenium 34 | 80 **Br** Bromine 35 | 84 **Kr** Krypton 36 |

| 85 **Rb** Rubidium 37 | 88 **Sr** Strontium 38 | 89 **Y** Yttrium 39 | 91 **Zr** Zirconium 40 | 93 **Nb** Niobium 41 | 96 **Mo** Molybdenum 42 | **Tc** Technetium 43 | 101 **Ru** Ruthenium 44 | 103 **Rh** Rhodium 45 | 106 **Pd** Palladium 46 | 108 **Ag** Silver 47 | 112 **Cd** Cadmium 48 | 115 **In** Indium 49 | 119 **Sn** Tin 50 | 122 **Sb** Antimony 51 | 128 **Te** Tellurium 52 | 127 **I** Iodine 53 | 131 **Xe** Xenon 54 |

| 133 **Cs** Caesium 55 | 137 **Ba** Barium 56 | 139 **La** Lanthanum 57 * | 178 **Hf** Hafnium 72 | 181 **Ta** Tantalum 73 | 184 **W** Tungsten 74 | 186 **Re** Rhenium 75 | 190 **Os** Osmium 76 | 192 **Ir** Iridium 77 | 195 **Pt** Platinum 78 | 197 **Au** Gold 79 | 201 **Hg** Mercury 80 | 204 **Tl** Thallium 81 | 207 **Pb** Lead 82 | 209 **Bi** Bismuth 83 | **Po** Polonium 84 | **At** Astatine 85 | **Rn** Radon 86 |

| **Fr** Francium 87 | 226 **Ra** Radium 88 | 227 **Ac** Actinium 89 † | | | | | | | | | | | | | | | |

*58-71 Lanthanoid series
†90-103 Actinoid series

140 **Ce** Cerium 58	141 **Pr** Praseodymium 59	144 **Nd** Neodymium 60	**Pm** Promethium 61	150 **Sm** Samarium 62	152 **Eu** Europium 63	157 **Gd** Gadolinium 64	159 **Tb** Terbium 65	162 **Dy** Dysprosium 66	165 **Ho** Holmium 67	167 **Er** Erbium 68	169 **Tm** Thulium 69	173 **Yb** Ytterbium 70	175 **Lu** Lutetium 71
232 **Th** Thorium 90	**Pa** Protactinium 91	238 **U** Uranium 92	**Np** Neptunium 93	**Pu** Plutonium 94	**Am** Americium 95	**Cm** Curium 96	**Bk** Berkelium 97	**Cf** Californium 98	**Es** Einsteinium 99	**Fm** Fermium 100	**Md** Mendelevium 101	**No** Nobelium 102	**Lr** Lawrencium 103

Key

a **X** b	a = relative atomic mass X = atomic symbol b = proton (atomic) number

The volume of one mole of any gas is 24 dm³ at room temperature and pressure (r.t.p.).

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