CHEMISTRY

COMBINED SCIENCE NOTES

CHANGAMIRE

FORM 3 & 4

CHEMISTRY

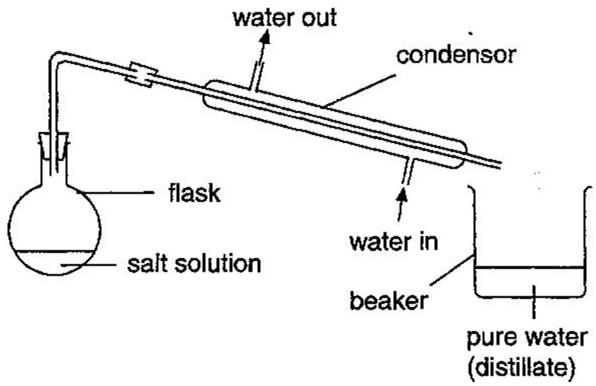
TOPICS

- **SEPARATION**
- **MATTER**
- **ACIDS, BASES AND SALTS**
- **OXIDATION AND REDUCTION**
- **INDUSTRIAL PROCESSES**
- **ORGANIC CHEMISTRY**

SEPARATION

Simple distillation

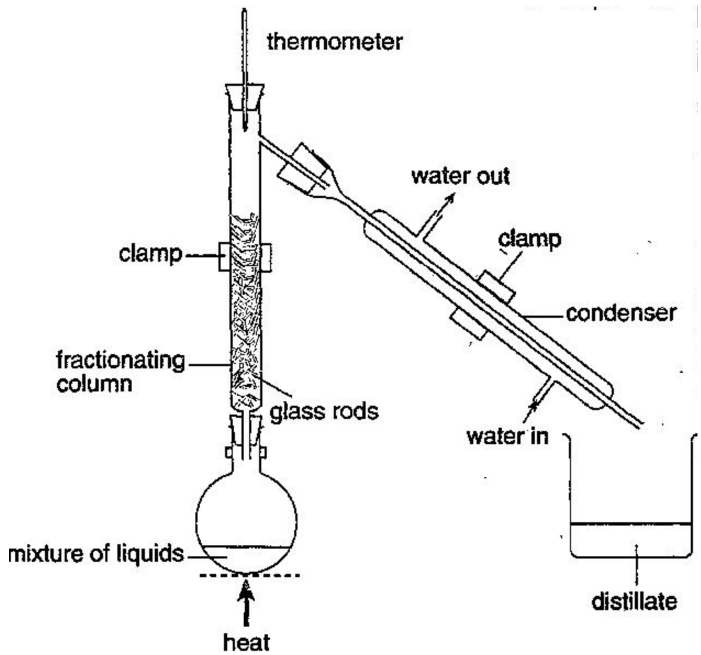
- Is a process of separating water from dissolved solids. It can be used to obtain a pure liquid from a mixture that contains dissolved impurities.
- Simple distillation can avoid the loss of water into the atmosphere, separating and keeping the two components of the original mixture.



- Impure liquid is heated, the water boils and evaporate. The vapour is passed through a condenser which is a glass tube surrounded by flowing water. This is where the vapour cools and condenses, dripping into the receiving flask as distillate.
- The solid (solute) remains behind in the flask as a residue.
- This method can be used to separate salt and water in a salt solution.
- Heat the solution in the flask. As it boils, water vapour rises into the condenser, leaving salt behind
- The condenser is cooled so that the vapour condenses to water. The water drips into the beaker as distillate. It is almost pure

Fractional distillation

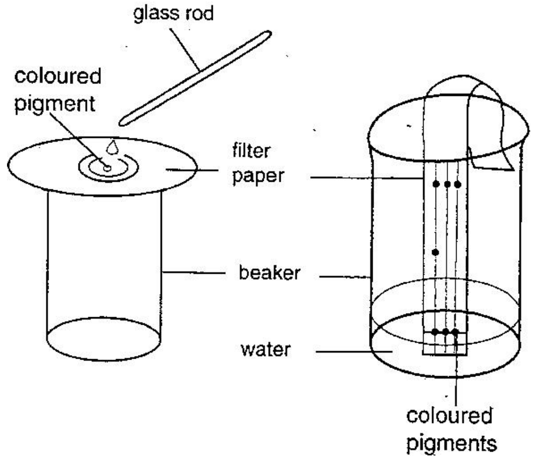
- It separates mixture of miscible (soluble) liquids with widely different boiling points e.g. alcohol and water.
- This method is ideal when one of the liquids is more volatile (evaporates more easily) than the other such as ethanol.



- The fractionating column is a glass tube that is usually packed with beads to increase surface area inside the column.
- The mixture of ethanol and water is placed in flask and heated. Ethanol with lower boiling point boils and vaporises first and reach fractionating column. It will condense on the beads in the column causing them to heat. When the beads reach a certain temperature when the ethanol won't condense anymore. It will rise while the water drip back. It is then cooled and condensed into ethanol as it passes through condenser.
- Temperature will stay constant until all ethanol is distilled.

Chromatography

- Paper chromatography is a method used to separate a mixture of substances e.g. can be used to find out how many different dyes are there in black ink.

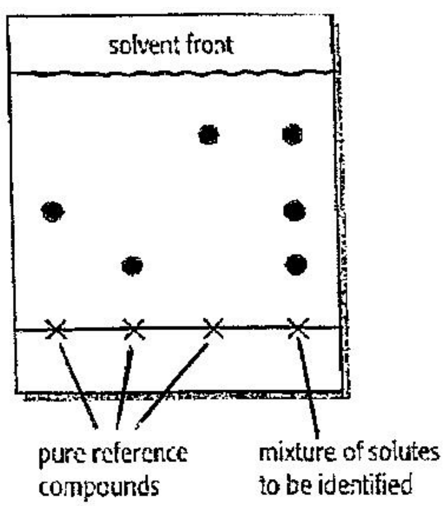


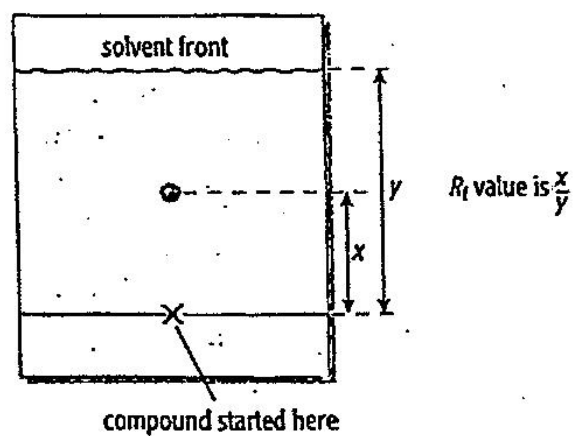
To separate substances

- Place a drop of ink onto the centre of some filter paper allow it to dry.
- Drop water on to the ink spot, one drop at a time. The ink slowly spreads out and separates into rings of different colours.
- The dyes in the ink have different solubilities in the solvent and are absorbed and travel across the paper at different rates hence they separate into rings. The filter paper with the coloured rings is called a chromatogram.
- The *stationary phase* is the material on which the separation takes place (the filter paper).
- The *mobile phase* consists of the mixture you want to separate, dissolved in a solvent.
- The solvent front is the furthest point reached by the solvent in chromatography

To identify substances

- Components of the mixture can be identified by comparison with pure reference compounds.
- Spots of substances placed onto a pencilled line which is called the **origin**. Pencil is used to draw the line because
- Paper goes in solvent, and solvent travels up paper.Cover the chromatogram so that the moving solvent does not evaporate before it has time to spread.
- To analyse the substance you could study the coloured spots on the chromatogram, as in paper chromatography
  1. Number of rings/dots = number of substances
  2. If two dots travel the same distance up the paper they are the same substance.
- **X** has separated into three spots Two are at the same height as **A** and **B**, so **X** must contain substances **A** and **B**. Does it also contain **C** and **D**?





- In a chromatogram the ratio of the distance travelled by the solute to the distance travelled by the solvent is known as the R value. You can calculate the **Rf value** to identify a substance, given by the formula:

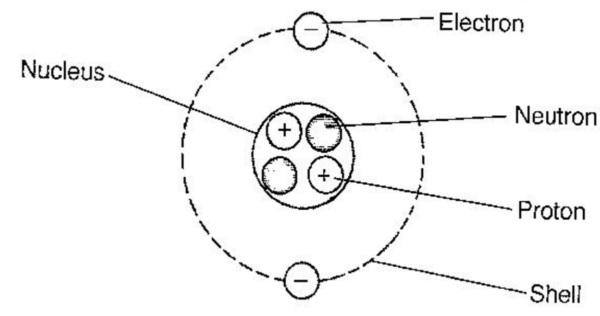
$$R_f = \frac{distance\ moved\ by\ substance}{distance\ moved\ by\ solvent}$$

Application of paper chromatography

- Separates and identify mixtures of coloured substances in dyes
- Separates substances in urine & blood for medicinal uses
- Detect traces of drugs in people's blood
- To analyse materials found at crime scenes or help in crime detection
- Used to manufacture of pharmaceutical drugs
- Purify a substance, by separating it from its impurities

MATTER

STRUCTURE OF AN ATOM



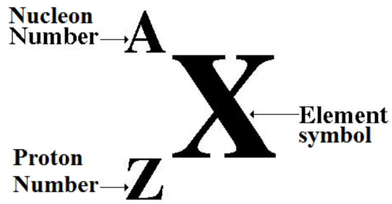
- An atom is the smallest particle of an element that can take part in a chemical reaction.
- Each atom consist of a very small and a dense nucleus, which contains protons and neutrons, surrounded by orbiting electrons.
- Protons and neutrons are both found in the nucleus of an atom and are collectively called *nucleons.*
- An atom is electrically neutral when the number of protons are equal to the number of electrons in an atom.

The sub-atomic particles

Particle	**Position**	**Charge**	**Mass**
Proton	Nucleus	+1	1
Neutron	Nucleus	0	1
Electron	Atomic shells	-1	$\frac{1}{1840}$/0.0005

THE NUCLIDE NOTATION

- If the chemical symbol for an element is represented by the letter **X** the symbol for the mass number (**A**) is written at the top left hand side of the chemical symbol. The proton number (**Z**) is written at the lower left hand side of the symbol like this;



- For example the element carbon may be represented like this; ${}^{12}_{6}C$
- This means that carbon has 12 nucleons of which 6 are protons. C is the chemical symbol for the element.

Mass number (nucleon number) – is the number of protons and neutrons in the nucleus (nucleons) of an atom.
Proton number (atomic number) – is the number of protons in the nucleus of an atom. It is equal to the number of electrons in an atom.

- We can use the mass number and atomic number to determine the number of neutrons in an element. The difference between the proton number and the mass number gives the number of neutrons for each element.

$$\text{number of neutrons} = \text{mass number} - \text{proton number}$$

Examples For each of the following, state the number of neutrons it have.

oxygen $= 16 - 8 = 8\ neutrons$

chlorine$= 35 - 17 = 18\ neutrons$

calcium$= 40 - 20 = 20\ neutrons$

Isotopes

- These are atoms of the same element that have different number of neutrons but same number of protons.
- Isotopes of an element have the same proton number but different nucleon numbers. They will have the same properties because they have the same number of electrons.
- The mass number of an element with isotopes becomes the weighted average of the masses of each of the isotopes. Weighted averages take into account the proportion of each isotope present on earth.

Chlorine isotopes

- Chlorine has two isotopes i.e. Cl – 35 and Cl – 37. Chlorine – 35 has a mass number of 35 (17 protons, 18 neutrons). Chlorine – 37 has a mass number of 37 (17 protons, 20 neutrons). Using the abundances of chlorine (75% of all chlorine on earth is chlorine – 35; 25% is chlorine – 37), then;

$$mass\ of\ chlorine = \left(\frac{75}{100} \times 35\right) + \left(\frac{25}{100} \times 37\right) = 35.5$$

Oxygen isotopes

- Oxygen has two isotopes i.e. O – 16 and O – 18. Oxygen – 16 has a mass number of 16 (8 protons; 8 neutrons). Oxygen – 18 has a mass number of 18 (8 protons; 10 neutrons). Using the abundances of oxygen (99.8% of all oxygen on earth is oxygen – 16 and the remaining 0.2% is oxygen – 18), then;

$$mass\ of\ oxygen = \left(\frac{99.8}{100} \times 16\right) + \left(\frac{0.2}{100} \times 18\right) = 15.84$$

Carbon isotopes

- Carbon has two isotopes i.e. C – 12 and C – 14. Carbon – 12 is far more abundant and is present in organisms in large proportions. Carbon – 14 is present in living organisms, but in samller amounts than carbon – 12. Carbon 14 is radioactive

The first 20 elements

Group / Period	I	II	III	IV	V	VI	VII	VIII
1	$_{1}H$							$_{2}He$
2	$_{3}Li$	$_{4}Be$	$_{5}B$	$_{6}C$	$_{7}N$	$_{8}O$	$_{9}F$	$_{10}Ne$
3	$_{11}Na$	$_{12}Mg$	$_{13}Al$	$_{14}Si$	$_{15}P$	$_{16}S$	$_{17}Cl$	$_{18}Ar$
4	$_{19}K$	$_{20}Ca$						

The electronic configuration of the first 20 elements

- Electronic configuration is the arrangement of electrons in an atom. The electronic configuration of any element can be deduced by finding its position in the periodic table.
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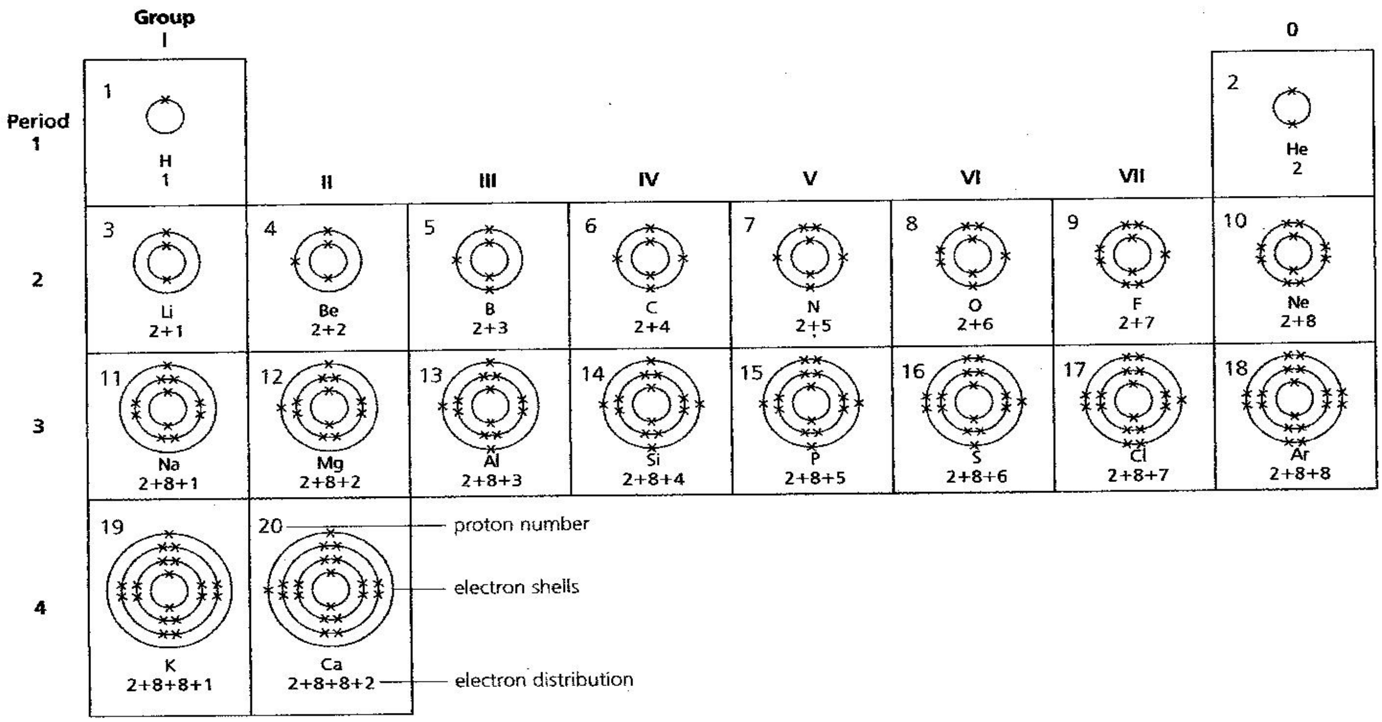
- For example: find the position of aluminium on the periodic table. It is in group III and period 3. The means that it has three electrons in the outer shell which is the third shell. The first two shells are full. Its electronic configuration is 2.8.3

- Each energy shell can only have a certain number of electrons. The energy shells become larger, the further they are from the nucleus. The larger a shell, the more electrons it can hold. The shells fill in order, from lowest energy level to highest energy level
    - First atomic shell is small and can only accommodate two electrons.
    - Second, third and fourth shells can accommodate 8 electron
    - The first shell is filled before an electron can go into the second shell



GROUP PROPERTIES

- The element in each number group show trends in their properties.
- The outer shell electrons are also called valency electrons and their number shows how elements behave. All elements in a group have similar properties.
- Some groups have special names;
    - ✓ Group I – alkali metals
    - ✓ Group II – alkaline earth metals
    - ✓ Group VII – halogens
    - ✓ Group VIII – noble gases

Group I Elements – The Alkali Metals

Physical properties

- They are soft solid and can be easily cut. Softness increase down the group.

- They are grey solids with shiny silvery surfaces when freshly cut and turn dull when exposed to air because they are very reactive
- Melting and boiling points decrease down the group
- Have low densities which increase down the group.
- They are good conductors of heat and electricity

Chemical properties

- Form ionic compounds in which the metal ion has a charge of +1.
- React violently with chlorine, water and oxygen
- Produce soluble white compounds which dissolve in water to give colourless solutions.
- They become reactive down the group because the valency electron so it is lost more easily due to less strong force of attraction.

Group II Elements – The Alkali Earth Metals

Physical properties

- They are all shiny, silvery white metals that are less metallic in character than group 1 metals.
- Melting and boiling points decrease down the group
- Have very low densities that decrease down the group

Chemical properties

- Form ionic bonds
- Are reactive but less than group 1 metals. Reactivity increase down the group
- React with water to form alkaline solutions

Group VII Elements – The Halogens

Physical properties

- Form diatomic molecules (containing two atoms) e.g. F_2, Cl_2
- Boiling points increases down the group
- Form coloured gases. Fluorine is pale yellow and chlorine is a green gas, bromine forms a red vapour and iodine a purple vapour. Colour gets darker down the group
- Do not conduct electricity and heat
- Are poisonous/toxic non metallic elements
- Density increases down the group

Chemical properties

- React with metals to form halides. They are very reactive which decrease down the group because the smaller the atom, the easier it is to attract the electron – so the more reactive the element will be.

- Form acidic solutions
- A halogen will displace a less reactive halogen from a solution of its halide

Uses of halogens

- Water purification e.g. chlorine
- Making chlorofluorocarbons (aerosols) e.g. fluorine
- Making pesticides e.g. fluorine
- Making hydrochloric acid e.g. Chlorine
- Bleaching agents e.g. chlorine in wood pulps
- Flame retardants e.g. bromine
- Refrigerants and lubricants e.g. F
- Disinfectants e.g. Cl, Br and I

Group 0 Elements – The Noble Gases

- A non-metal group
- They exist as colourless which occur naturally in air and are insoluble in water
- Low melting and boiling points which increases down the group
- Have very low densities which increase down the group
- Exist as mono-atomic molecules
- They are unreactive because they have full outer shell, making them safe to use.

CHEMICAL BONDING

- A chemical bond is a force that holds atoms together to form a compound or molecule
- The electrons involved in bonding are ones found in the outermost shell of the atom (valence electrons).
- Valency is the number of electrons that an atom must gain or lose to obtain an outer electron shell configuration that is the same as that of the nearest noble gas.
- Atoms combine with other atoms so that they have a stable configuration (noble gas configuration) by donating, accepting and sharing electrons.

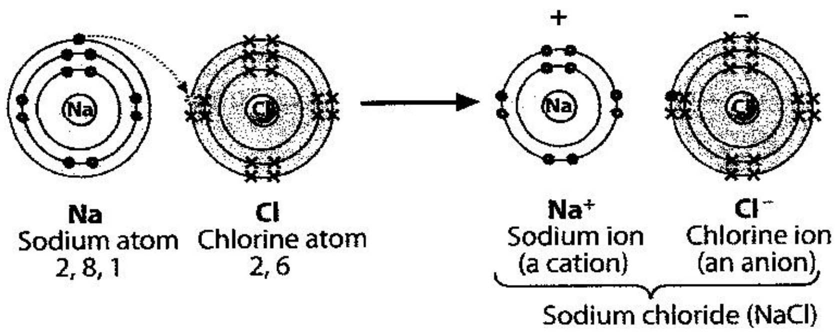
Ionic bonding

- Occurs between a metal and a non-metal
- Involves the transfer of electrons from metals to non-metals. Atoms of metal donate electrons to form a positive ion. Atoms of non-metals accept electrons to form a negative ion.
- The positive and negative ions are attracted to each other through electrostatic forces, forming an ionic bond.

- When elements combine to form ionic compounds, the positive charge on the metal ions must be balanced by the negative charge on the non-metals.
- The charge on an ion is equal to its valency i.e. the charge on the metal ion is equal to its group number and the charge on a non-metal ion is equal to 8 minus the group number.

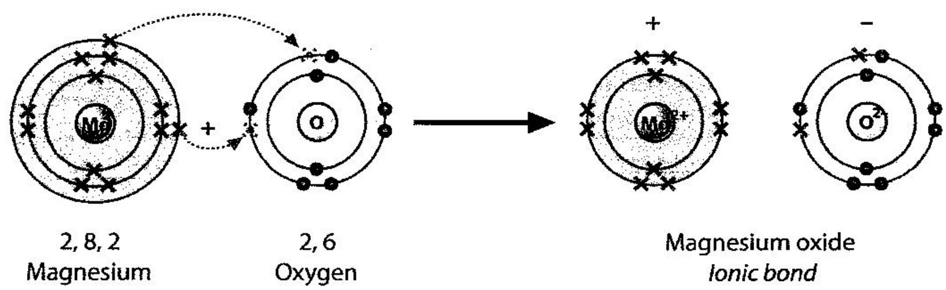
Sodium and Chlorine

- Sodium atoms have one electron in their outer shell. They will lose this electron during bond formation and have an ionic charge of +1. Chlorine atoms have seven outer electrons. They gain one electron during bond formation and have an ionic charge of – 1. Hence one sodium atoms and one chlorine atom combine to form sodium chloride, formula NaCl.



Magnesium and oxygen

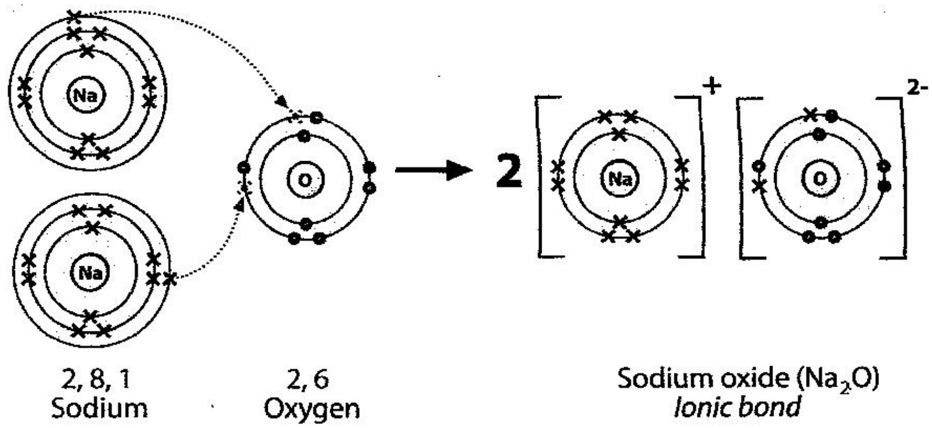
- Magnesium atoms have two electrons in their outer shell. They will lose these electrons during bond formation and have an ionic charge of +2. Oxygen atoms have six outer electrons. They gain two electrons during bond formation and have an ionic charge of – 2. Hence one magnesium atoms and one oxygen atom combine to form magnesium oxide, formula MgO.



Sodium and oxygen

- Sodium atoms have one electron in their outer shell. They will lose this electron during bond formation and have an ionic charge of +1. Oxygen atoms have six outer electrons.

They gain two electrons during bond formation and have an ionic charge of – 2. Hence two sodium atoms and one oxygen atom combine to form sodium oxide, formula Na_2O.



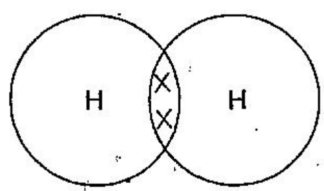
Properties of Ionic bonds

- ✓ Hard solids at room temperature
- ✓ High boiling point because of strong attraction forces
- ✓ Conduct electricity in molten state because ions are free to move
- ✓ Water soluble

Covalent bonding

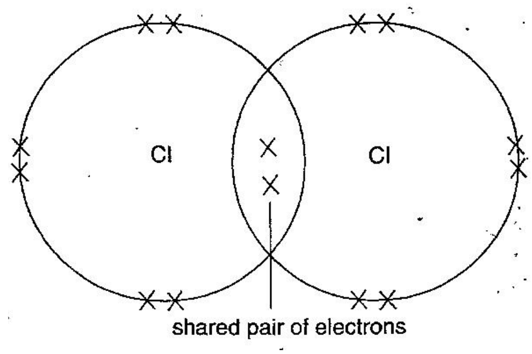
- Involves the sharing of electrons between non-metalsso that each atom attains a noble gas configuration.
- When elements combine to form covalent compounds the valency of each element determines how many of each atom combine.

Hydrogen molecule



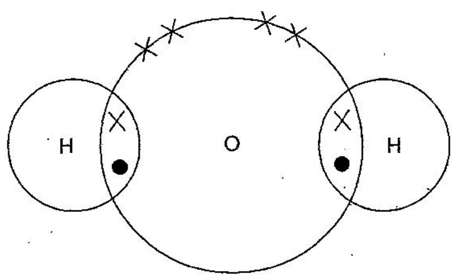
- Hydrogen atoms has a valency of 1. Two hydrogen atoms combine by sharing their electrons to form molecules of hydrogen, H_2. The two shared electrons revolve around both atoms, which in effect have the stable helium duplet.

Chlorine molecule



- Two chlorine atoms combine by sharing two electrons, one from each atom. Each atom then has 6 electrons of its own in the outer orbit plus the shared pair, which revolves around both atoms i.e. they each have a complete octet and the stable argon structure 2.8.8

Water molecule



- Oxygen needs 2 electrons and when bonded with two hydrogen atoms, which need an atom each, they combine to provide 2 electrons on both sides of oxygen.

Properties covalent bonds

- ✓ Most of them do not conduct electricity because molecules are not charged.
- ✓ Most are insoluble in water.
- ✓ Low melting and boiling point because of weak forces of attraction between molecules

THE MOLE CONCEPT

Relative atomic mass (A_r)

- Reative atomic mass is the mass of one atom of an element relative to the mass of one atom of carbon-12.
- The mass of any atom is compared to the mass of a nucleon in a carbon 12 atom.
    - A_r (Ca) = 40

- A_r (Cl) = 35.5
- A_r (O) = 16
- A_r (Zn) = 52

Relative Formula or Molecular Mass (Mr)

- Relative molecular mass is the average mass of a molecule relative to the mass of a carbon-12 atom.
- Relative formula mass is given by the sum of atomic masses of the elements.
- The formula mass of sodium chloride, sodium oxide and sodium carbonate are calculated as follows:
    - Mr (NaCl) = 23 + 35.5 = 58.5
    - Mr (Na_2O) = 2(23) + 16 = 62
    - M_r(Na_2CO_3) = 2(23) + 12 + 3(16) = 106
- The relative molecular mass of a covalent compound is found in the same way as the formula mass of an ionic compound.
    - M_r(HCl) = 1 + 35.5 = 36.5
    - $M_r(CO_2)$
    - $M_r(NH_3)$

The Mole

- The mole is the amount of substance which contains the same number of molecules or atoms as they are in 12 grams of carbon.
- One mole is the
    - Amount of substance which contains an Avogadro number of particles
    - Atomic mass of an element in grams
    - Molecular mass or formula mass of a compound in grams

✓ 1 mole of MgO = 24 + 16 = 40g
✓ 2 mole of HCl = 2[1 + 35.5] = 73g
✓ 1 mole of O_2 = 2[16] = 32g
✓ 1 mole of Mg = 12g

$$number\ of\ moles(n) = \frac{mass\ (m)}{molecular\ mass(Mr)}$$

Examples 1 Express the following in moles
(a) 5g Na
(b) 22g CO_2
(c) 35g $CuSO_4$

Examples 2 Calculate the mass

(a) 0.25 moles of carbon monoxide

(b) 3 moles of water

(c) 2 moles carbon dioxide

Chemical equations

- A chemical equation represents a chemical change by means of symbols and formulae
- Equations tell us how much of each reactant is involved and how much product is formed.
- A balanced chemical equation shows the proportions of substances reacting and of products formed.
- When writing chemical equations
    - Write a word equation with reactants on the left hand side of the equations and products on the right hand side.
    - Write correct formula for each reactant and each products. (symbols for elements, formulae for compounds)
    - Write state symbols next to each substance i.e. (s) for solid, (aq) for aqueous solutions, (l) for liquid and (g) for gas
    - Balance the equation so that there are the same number of each type of atom on both sides of the equations. Number to balance can only be written in front of each substance.

Radicals and their valency

Hydroxide	OH^-	valency – 1
Nitrate	NO_3^-	valency – 1
Carbonate	CO_3^{2-}	valency – 2
Ammonium	NH_4^+	valency +1
Sulphate	SO_4^{2-}	valency – 2

Example 1 For the burning of carbon to give carbon dioxide:

$$carbon + oxygen \rightarrow carbon\ dioxide$$
$$C(s) + O_2(g) \rightarrow CO_2(g)$$

This equation is balanced because there are the same number of atoms on each side of the equation.

Example 2 For the reaction between sodium and water:

$$\text{sodium } + \text{water} \rightarrow \text{hydrogen} + \text{sodium hydroxide}$$
$$Na(s) + H_2O(l) \rightarrow H_2(g) + NaOH(aq)$$
$$\mathbf{2}Na(s) + \mathbf{2}H_2O(l) \rightarrow H_2(g) + \mathbf{2}NaOH(aq)$$

Calculating Quantities

- If the actual amounts of two substances that react is known, the other amounts that willl react and products formed will be predicted.
- Elements always react in the same ratio to form a compound. The total mass does not change, during a chemical reaction, so total mass of reactants is equal to the total mass of products.

Calculations of Mass of Reactant or Mass of Product

Example1 What mass of magnesium oxide is obtained from the complete combustion of 12g of magnesium?

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

1 mole of Mg weighs 24g
1 mole of MgO weighs 40g
Therefore 24g of magnesium form 40g of magnesium oxide
12g of magnesium forms X gof magnesium oxide

$$\therefore x = \frac{40}{24} \times 12$$

= 20g of Magnesium oxide

Example 2 If 4.20g of sodium hydrogencarbonate are heated, what mass of anhydrous sodium carbonate will be formed?

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

2 moles $NaHCO_3$ = 168g
1 mole Na_2CO_3= 106g
Therefore 168g sodium hydrogencarbonate give 106g sodium carbonate
4.20g sodium hydrogencarbonate give Xg sodium carbonate

$$\therefore x = \frac{106}{168} \times 4.2$$

= 2.65g sodium carbonate

Example 3 Calculate the mass of ammonium chloride that will just react completely with 14.8g of calcium hydroxide.

$$2NH_4Cl + Ca(OH)_2 \rightarrow 2NH_3 + CaCl_2 + 2H_2O$$

2mole NH_4Cl = 2[14 + (4 x 1) + 35.5] = 107g
1mole $Ca(OH)_2$ = 40 + 2(16 + 1) = 74g
Therefore 74g calcium hydroxide reacts with 107g of ammonium chloride
14.8 g of calcium hydroxide reacts with Xg of ammonium chloride

$$\therefore x = \frac{107}{74} \times 14.8$$

= 21.4g ammonium chloride

Avogadro number (L)/ Constant

- Is obtained from an estimate of the number of carbon atoms in 12g of carbon 12
- **Avogadro number** is the number of particles (ions, molecules or atoms) in one mole of any substance (Mr).
- One mole (mol) is equal to **6.02 x 10^{23}** particles that make up the substance.
- One mole of any substance contains the same number of particles. This number is the avogadro number or constant.
  - ✓ 12g of carbon contains 6 x 10^{23} carbon atoms
  - ✓ 2g of hydrogen contains 6 x 10^{23} hydrogen molecules
  - ✓ 44g of CO_2 contains 6 x 10^{23} carbon dioxide molecules

$$\boldsymbol{number\ of\ particles = number\ of\ moles \times avogadro\ constant}$$

Example 1 Given Avogadro's constant is 6 x 10^{23}/ mol, calculate the number of atoms in

(a) 35.5g of chlorine
(b) 3.1g phosphorus

Example 2 How many particles are there in

(a) 0.2 moles of carbon
(b) 0.5 moles of ammonia

Example 3 What mass do the following have?

(a) 5 x 10^{23} atoms of sodium
(b) 3 x 10^{23} molecules of carbon dioxide

Empirical Formula

- An empirical formula is the **simplest** ratio in which atoms combine to form up a given substances e.g. the formula of compound may be N_2O_4 but the simplest ratio of N to O is NO_2 therefore the empirical formula of that compound is NO_2
- If the percentage of a compound is known by mass, number of moles of each element present in the compound is calculated and then the empirical formula of the compound.

General method

- Convert the given % composition or mass of each element into moles
- Divide each mole value by the smallest number of moles obtained
- Write each mole in its simplest whole number ratio. This is the mole ratio which is represented by subscripts in the empirical formula

Example 1 Calculate the empirical formulae of substances which have the following compositions by mass; 43.4% sodium, 11.3% carbon, 45.3% oxygen

Symbols	Na	C	O
Number of moles	$\frac{43.4}{23}$ = 1.89	$\frac{11.3}{12}$ = 0.942	$\frac{45.3}{16}$ = 2.83
Simplest ratio (divide by the smaller number)	$\frac{1.89}{0.942}$ = 2.01	$\frac{0.942}{0.942}$ = 1	$\frac{2.83}{0.942}$ = 3
Formula	Na_2CO_3		

Example 2 6g of magnesium is heated in oxygen, after cooling and reweighing it is found that there is 10g of magnesium oxide. Find the empirical formula for magnesium oxide

Example 3 Calculate the empirical formula for a compound that contains 1.82g of potassium, 5.93g of Iodine and 2.24g of oxygen

Molecular Formula

- A molecular formula gives the exact number of atoms of the different elements in a molecule of the compound.

$$\textbf{molecular formular} = \frac{\textbf{Mr}}{\textbf{emperical formula mass}} \times \textbf{emperical formula}$$

Example What is the molecular formula of the compound, A, which has an empirical formula CH_2O and a relative molecular mass of 60?

Molecular formula $= \frac{60}{30} \times CH_2O$

$= 2 \times CH_2O \qquad = C_2H_4O_2$

CONCENTRATION

- The concentration of a solution is the amount of solute, in grams or moles that is dissolved in 1 $dm^3$of solution.
- Concentration shows the number of moles present in the solution.

$$\text{Concentration} = \frac{\text{amount of solute (mol)}}{\text{volume of solution } (dm^3)}$$

$$c = \frac{n}{V}$$

- Units are mol/dm^3 or g/dm^3
- Molar concentration is symbolised by M. it is the number of moles of the solute dissolved in $1dm^3$, expressed in mol/dm^3.
  - ✓ 1 mole of solute in $1dm^3$ of solution has a concentration of $1mol/dm^3$ (1M)
  - ✓ 1 mole of solute in $2dm^3$ of solution has a concentration of $0.5mol/dm^3$ (0.5M)
  - ✓ 2 moles of solute in $1dm^3$ of solution has a concentration of $2mol/dm^3$ (2M)
  - ✓ 1 mole of solute in $500cm^3$ of solution has a concentration of $2mol/dm^3$ (2M)
- Mass concentration is the mass of solute in grams in $1dm^3$ of solution expressed in g/dm^3 and is calculated as follows;

$$\text{concentration} = \frac{\text{mass of solute}}{\text{volume of solution}}$$

  - ✓ 2.5g of copper sulphae in 1 dm^3 of water has a concentration of $2.5g/dm^3$
  - ✓ 125g of copper sulphate in $0.5dm^3$ of water has a concentratin of $250g/dm^3$
- Moles can be converted to grams by multiplying the number of moles by M_r.
- Grams can be converted to moles by dividing the mass by the M_r.

Example 1 Calculate the concentration in mol/dm^3 of a solution containing 36.5g of hydrogen chloride in 4.00dm^3 of solution.

Method:

M_r of HCl = 35.5 + 1 = 36.5

Amount in moles present in 36.5g = 1.00mol

Volume = 4.00dm^3

Concentration of solution $= \frac{\text{amount of solute in moles}}{\text{volume solution in dm}^3}$

$= \frac{1.00\text{mol}}{4.00\text{dm}^3}$

= 0.25mol/dm^3 or 0.25M solution

Example 2 Calculate the amount of solute in moles in cm^3 of a solution of sodium hydroxide which has a concentration of 2.00mol/dm^3

Method:

Concentration of solution = 2.00mol/dm^3

Volume of solution = 250cm^3 = 0.250 dm^3

Amount of solute = volume x concentration

= 2.00 x 0.250 = 0.500 mol

Example 3 What is the concentration of solution made by dissolving 5.00g of Na_2CO_3 in 250cm^3water?

Moles = mass/Mr

= 5.00 / (23 x2 + 12 +16 x3)

= 0.0472 mol

Concentration = moles/Volume

= 0.0472 / **0.25**

= 0.189 mol dm^{-3}

Example 4 Find the number of moles of Sodium hydroxide in 25cm^3 of solution of concentration 0.1mol/dm^3

Example 5 Find volume of solution of concentration 2 mol/dm^3 that contains 0.005 moles of hydrochloric acid.

METALS AND NON METALS

General properties of metals

- ✓ good conductors of electricity and heat
- ✓ high melting and boiling points – which means they are solid at room temperature
- ✓ can be hammered into different shapes (they are **malleable**)

- ✓ can be drawn out to make wires (they are **ductile**)
- ✓ look shiny when they are polished – have a **lustre**
- ✓ make a ringing noise when struck – they are **sonorous**
- ✓ have high density
- ✓ react with oxygen to form oxides that are **bases**.

General properties of non-metals

- ✓ non-metals have low density
- ✓ look dull, in the solid state
- ✓ non-metals are not malleable or ductile – they are brittle
- ✓ non-metals break up easily – they are **brittle**
- ✓ do not conduct electricity or heat
- ✓ lower melting and boiling points – many are gasesat room temperature
- ✓ react with oxygen to form oxides that are **acidic**.

REACTIVITY OF METALS

Reaction of Metals with Water

- Potassium and Sodium reacts vigorously with cold water to alkali solutions (hydroxides) and hydrogen gas. Calcium reacts slowly with cold water.

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

- Magnesium reacts vigorously with steam to form a white oxide and hydrogen gas. A lot of heat and light energy is emitted. No reaction with cold water.

$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$

- Zinc do not react with cold water but reacts more slowly with steam to produce zinc oxide which is yellow when hot and white when cold.

$$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$$

- Iron reacts with steam but does not react with cold water. Rusting occurs very slowly in the presence of oxygen.

$$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + H_2(g)$$

- Copper and Lead does not react with water and steam

Reaction of Metals with dilute Acids

- Potassium, Sodium and Calcium reacts explosively to with dilute acids to form salts and hydrogen.

$$2K(s) + 2HCl(aq) \rightarrow 2KCl_2(s) + H_2(g)$$
$$Na(s) + 2HCl(aq) \rightarrow NaCl_2(s) + H_2(g)$$

- Mg, Al, zinc and iron reacts mildly with dilute acids to form a salt and hydrogen.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(s) + H_2(g)$$
$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(s) + H_2(g)$$

- Copper and Lead have no reaction with dilute acid except with concentrated acids

NB To test for hydrogen use a burning splint, it will explode with a pop sound

Reaction of metals with oxygen

- Sodium burns with a bright yellow flame to produce a white oxide

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

- Calcium burn with a red flame to produce a white oxide

$$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$$

- Magnesium burn with a bright white flame to produce a white solid oxide

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

- Aluminium burns when strongly heated to form a white oxide

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

- Zinc burns with a to produce a yellow oxide when hot and a white when cold

$$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$$

- Iron powder burns with bright yellow flame to produce a blue-black oxid

$$3Fe(s) + 2O_2(g) \rightarrow Fe_3O_2(s)$$

- Copper powder burns with dull red glow to produce a black oxide

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

Reactivity Series

- Reactivity series is the arrangement of metals in order of their reactivity with different substances such as the reaction of metals with water, dilute acids and oxygen.

Potassium **most reactive**
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Lead
Copper **least reactive**

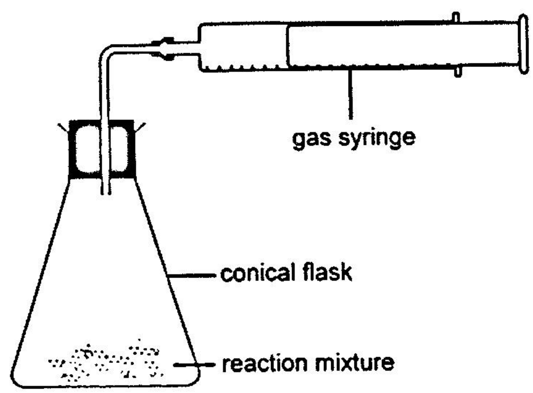
SPEED OF REACTION

Measuring Speed of Reaction

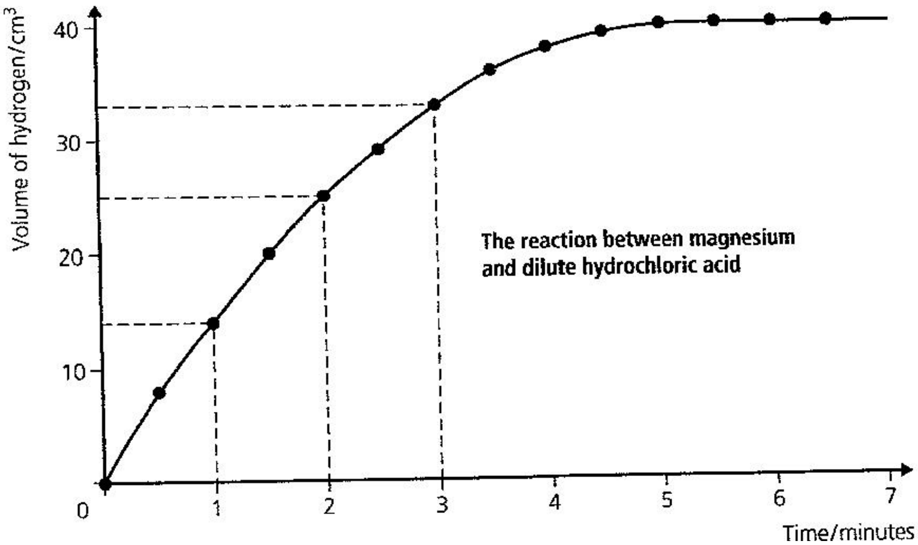
- Rate is a measure of the change that happens in a single unit of time.
- In general, to find the rate of a reaction, you should measure: the amount of a reactant used up per unit of time *or* the amount of a product produced per unit of time.
- Speed of reaction is *inversely proportional* to time taken; the shorter the time needed for reaction to complete, the faster the speed of reaction is.
- The speed of a reaction in which a gas is produced can be followed by measuring the change in volume which occurs as the reaction proceeds.

Measuring the amount of gas evolved

- When zinc is added to dilute sulfuric acid, they react together. The zinc disappears slowly, and a gas bubbles off.
- As time goes by, the gas bubbles off more and more slowly.This is a sign that the reactionis slowing down.
- Finally, no more bubbles appear. The reaction is over, because all the acid has been used up. Some zinc remains behind.
- For this reaction, it is easiest to measure the amount of hydrogen produced per minute, since it is the only gas that forms. It can be collected as it bubbles off, and its volume can be measured. 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.



Effect of surface area

- The smaller the particles are, the faster the rate of reaction.This is because many small particles will have more surface area exposed for collision is larger, frequency of collision between the reactant particles increases and hence the frequency of effective collision also increases. More products are formed per unit time and hence the rate of reaction increases.

Effect of concentration

- The higher the concentration of the reactant particles, the higher the rate ofreaction. This is because in a more concentrated solution, there are more reactant particles per unit volume. The frequency of collision between thereactant particles increases and hence the frequency of effective collisionalso increases. More products are formed per unit time and hence the rate ofreaction is higher..
- Since rate of reaction is inversely proportional to time, the shorter the time, thehigher the rate of reaction.

Effect of temperature

- The higher the temperature of the reacting system, the higher the rate ofreaction.This is because at higher temperature, the average kinetic energy of the particles increases. 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. More products areformed per unit time and hence, the rate of reaction is higher.

Effect of catalyst

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

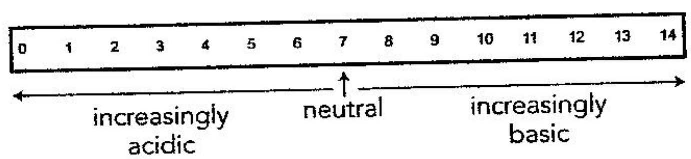
- Speed of catalysed reactions can be increased by:
  - ✓ increasing temperature
  - ✓ increasing concentration of solutions
  - ✓ increasing pressure of gas reactions
  - ✓ using catalyst provide "alternative path" which results in lower activation energy.

ACIDS, BASES AND SALTS

ACIDS AND BASES

pH

- pH is the **p**ower of the concentration of **H**ydrogen ions (H^+) present in a solution.
- The more acidic a solution is the larger the concentration of H^+ ions present.
- Alkalis release hydroxide (OH^-) ions when they react with water. The more alkaline a solution is, the higher the concentration of OH^- ions.
- The pH scale is a numerical measure of the relative acidity or alkalinity of a substance using a numerical scale.
- The lower the pH value the more acidic a substance is. The higher the pH value, the more basic a substance is.
- Using the pH scale, we are able to tell the relative acidity or alkalinity of a substance instead of simply identifying them as being acidic or alkaline.
- The pH scale ranges from 0-14 to express acidity and alkalinity.
- Any neutral solution have a pH of 7; a solution with a pH greater than 7 is alkaline and one with a pH less than 7 is acidic.



Neutralization

- During neutralization H^+ ions combine with OH^- ions to form water. As more and more OH^- ions are added the H^+ ions are used up to form water, the acidity therefore decrease. At the neutral point, just enough OH^- ions have been added to remove all the H^+ ions.
- There are not only hydrogen ions and hydroxide ions present when an acid reacts e.g. in the neutralization of sodium hydroxide with hydrochloric acid; there are also sodium ions from the alkali and chloride ions from the acid. These two ions remain in the solution and if the solution is allowed to evaporate, a solid called sodium chloride remains a salt.

Indicators

- An indicator is a substance that changes colour according to the acidity or alkalinity of the solution it is in e.g. litmus paper, methyl orange, phenolphthalein

Indicator	Colour in acid	Colour in alkali
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Red

- Universal indicator is a mixture of dyes that changes color to correspond with each value of the pH scale and the colours vary slightly depending on the pH of the solution.

pH	1	2 - 4	5 - 6	7	8 - 10	11 - 13	14
Colour changes	red	orange	yellow	green	blue	purple	violet

To identify pH values for different substances

- ✓ Half fill five test tubes with each of the following substances i.e. sodium hydroxide, vinegar, lemon juice, dilute hydrochloric acid and distilled water.
- ✓ Add a few drops of universal indicator solution in each beaker
- ✓ Note the colour changes and compare them against universal indicator chart.
- ✓ Record the results in the table below and draw conclusions about the colour change in each substance.

Substance	Universal indicator colour	pH	Conclusions
HCl			
NaOH			
Lemon juice			
Vinegar			
Distilled water			

PREPERATION OF SALTS

Reaction of acids with metals

- When an acid reacts with a metal, hydrogen is displaced, leaving a salt in solution
- The salt produced when an acid reacts with a metal depends on the acid that is used in the reaction i.e.
    - Sulphuric acids reacts with metals to form sulphate salts
    - Hydrochloric acid reacts with metals to form chloride salts
    - Nitric acid reacts with metals to form nitrate salts

$$Mg(s) + H_2SO_4\,(aq) \rightarrow MgSO_4(aq) + H_2(g)$$
$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Reaction of acids with carbonates

- Carbonates react with acids to give salt, water and carbon dioxide. There is fizzing or effervescence.
  $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2\,(aq) + H_2O(l) + CO_2(g)$
  $MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4\,(aq) + H_2O(l) + CO_2(g)$
  $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

NB to test for the gas, bubble it into lime water or bicarbonate indicator. If it is carbon dioxide, it turns lime water milky or bicarbonate indicator red.

Reaction of acids with bases

- Bases reacts with acids and neutralize them, giving salt and water. Metal oxides and hydroxides are bases

$$HCl(aq) + NaOH(s) \rightarrow NaCl(s) + H_2O\,(l)$$

$$H_2SO_4(aq) + CuO(s) \rightarrow CuSO_4(aq) + H_2O(l)$$

$$HNO_3(aq) + Na_2O(s) \rightarrow 2NaNO_3(aq) + H_2O(l)$$

TITRATION

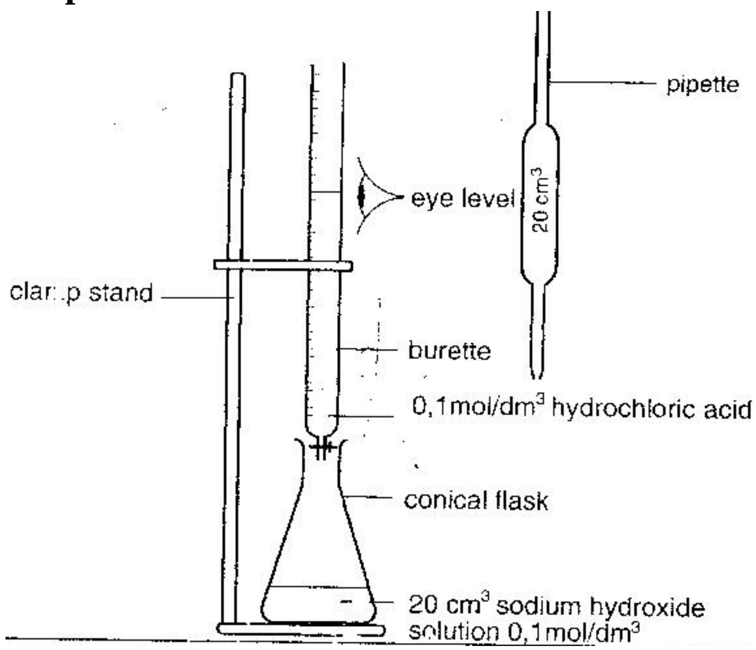
- Titration is a method of analysis in which the solution of known concentration is used to determine the concentration of another unknown solution
- Titration pipettes an exact volume of one solution and reacts this with another solution delivered from a burette. In this way the number of moles that have reacted in each solution can be calculated
- In titration one reactant is slowly added to the other in the presence of an indicator. The indicator changes colour when the reaction is complete.

Apparatus used in a titration

- *Pipette* – used to transfer an exact or fixed volume of a solution into the flask
- *Indicator* – to show when the end point has been reached in an acid – base reaction titration e.g. methyl orange, phenolphthalein or universal indicator solution.

- *Burette* **-** is a narrow graduated glass tube with a tap at the bottom which is used during titrations to deliver exact volumes of solution.
- *Conical flask* – allows the mixture to be safely swirled without spilling
- *Stand & clamp* – to hold the burette in a perfectly vertical position and make sure that measurements are accurate.

Acid-base titration procedure



- Add $25cm^3$ of the sodium hydroxide to a conical flask using a pipette.
- Add a few drops of indicator e.g. phenolphthalein to determine the end point of the reaction.
- Add the dilute acid from a burette slowly while stirring, until indicator goes colourless (end point). This shows that all the alkali has been neutralised. The solution is neutral and no more acid added.
- At the end point, the volume of dilute HCl that neutralise $25cm^3$ of dilute NaOH solution was accurately determined.
- The whole experiment should be repeated to obtain an average value for the volume of HCl needed — called the **titre**.
- Repeat without the indicator. Put $25cm^3$ of alkali in the flask and add the correct amount of acid to neutralise it.
- Finally heat the solution from the flask to evaporate the water. White crystals of sodium chloride will be left behind

Calculations involving titrations

Example 1 30cm^3 of 0.1mol/dm^3 NaOH reacted completely with 25cm^3 of H_2SO_4 in a titration. Calculate the concentration of H_2SO_4 in mol/dm^3 given that:

$$\mathrm{2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)}$$

Step 1 Find the reacting mole of NaOH

$$n(NaOH) = concentration \times volume$$
$$= 0.1 \times 0.03 =$$

Step 2 find the ratio of number of moles of H_2SO_4 to number of moles of NaOH

$$\frac{n(H_2SO_4)}{n(NaOH)} = \frac{1}{2} = 0.5$$

Step 3 use the ratio to find number of moles of H_2SO_4 that reacted

$$n(H_2SO_4) = 0.5 \times = 0.0015mol$$

Step 4 find the concentration of H_2SO_4 in mol/dm^3

$$concentration\ = \frac{n(H_2SO_4)}{volume} = \frac{0.0015}{0.025} =$$

Example 2 In a titration experiment, a learner found out that 100cm^3 of 1.0mol/dm^3 NaOH solution neutralized 40cm^3 of HCl with unknown concentration. Given that;

$$\mathrm{NaOH + HCl \rightarrow NaCl + H_2O}$$

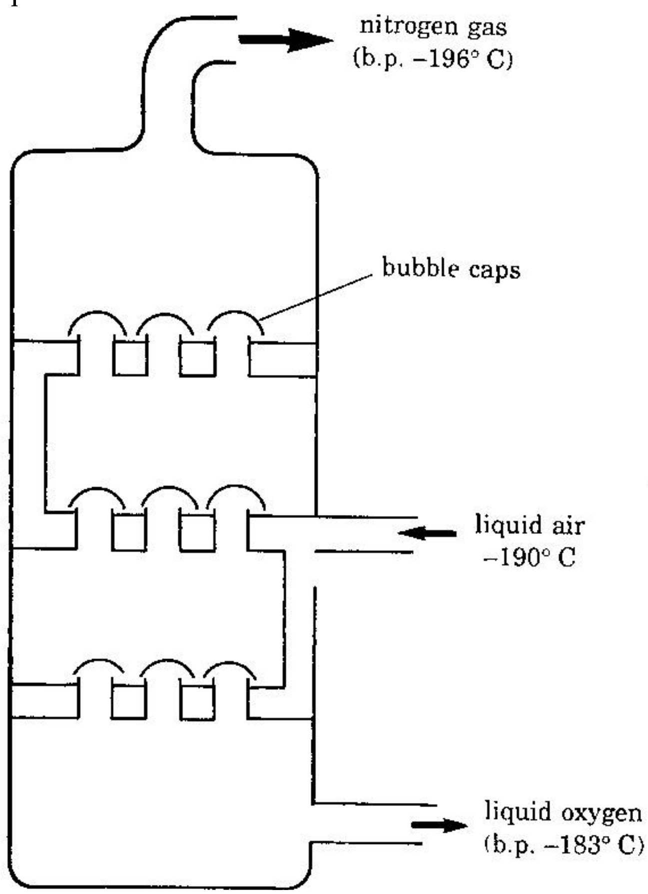
(a) Determine the number of moles of NaOH solution that were needed to neutralise the volume of acid.

(b) Determine the concentration of the acid.

INDUSTRIAL PROCESSES

Production of nitrogen and oxygen

- Air is cooled to -78°C, carbon dioxide and water are removed as solids.
- Remaining gases are compressed, cooled and allowed to expand rapidly causing further cooling.
- The process is repeated and temperature drops to – 200°C; oxygen and nitrogen are liquefied.
- Liquid air is pumped into a fractionating column and it is fractionally distilled.
- Nitrogen has a lower boiling point. The liquid nitrogen boils at -196°C and rises to the top where it is piped off and collected as a gas.
- Oxygen has a higher boiling point of – 183°C. It is collected at the bottom of the column and is removed as liquid



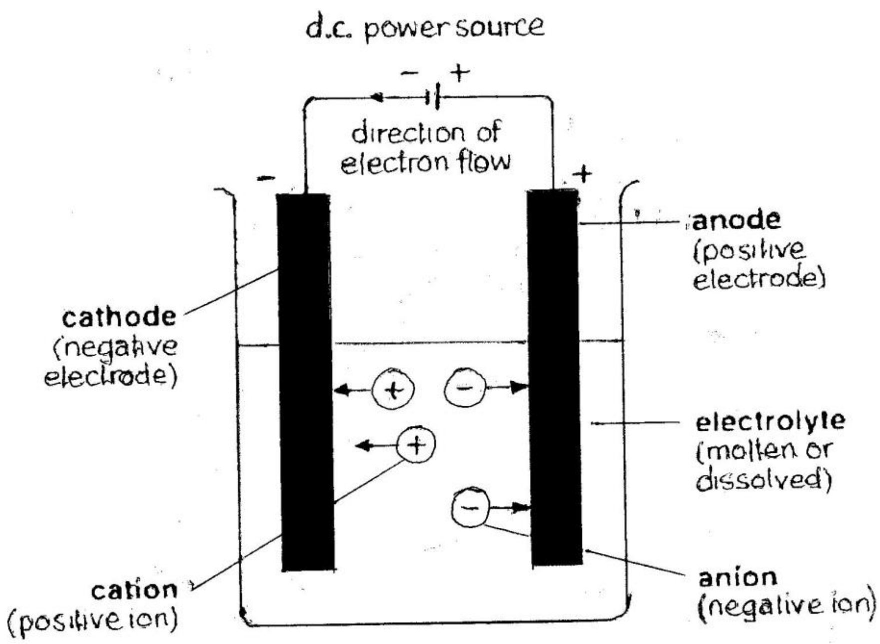
ELECTROLYSIS

- **Electrolysis** is a process through which compounds are decomposed into their elements by use of electricity.

- Ions must be free to move i.e. when an ionic substance is dissolved in water or when melted through heat.

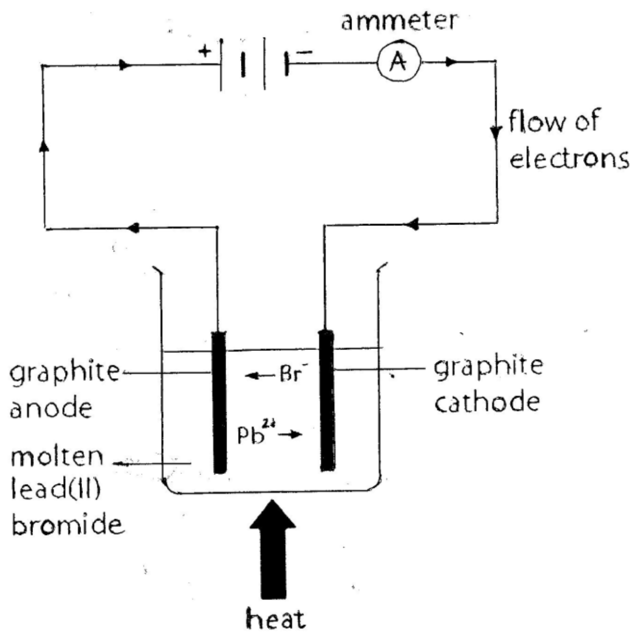
Electrolytic cell

- Converts electrical energy into chemical energy.
- It consists of;
  - An electrolyte i.e. is an ionic compound which conducts electric current in molten or aquous solution. The ions formed carry the current through the solution.
  - Electrodes i.e. conductors through which current passes into the electrolyte. Reactions take place at the electrodes. Electrode connected to the negative terminal is the **cathode.** Reduction takes place there**.** The **anode** is connected to the positive terminal of the battery. Oxidation occurs at the anode**.** Inert electrodes which do not take part in the reactions are used e.g. carbon.
  - Battery i.e. d.c. voltage supply or supply of electricity.



Electrolysis of molten lead bromide

- Lead bromide ($PbBr_2$) is an ionic compound in solid form (white powder)
- When $PbBr_2$ is heated, it melts and dissociates, leaving(lead ions) Pb^{2+} and (bromide) Br^- ions free to move.



- At the cathode the lead ions (Pb^{2+}) gains two electrons and are reduced to lead atoms which will form a silvery liquid metal at the bottom of the reaction vessel.

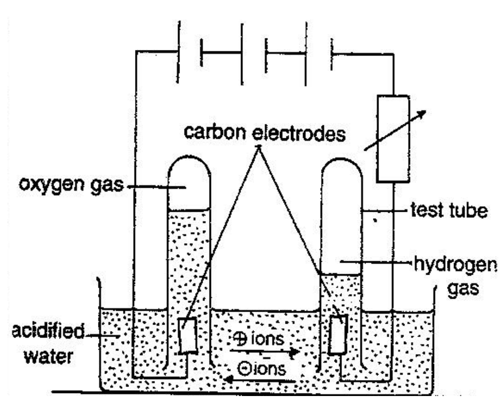
$$Pb^{2+} + 2e^{-} \rightarrow Pb$$

- At the anode bromide ions (Br^{-}) lose electrons to form bromine atoms which combine to form molecules of bromine gas/vapour. Bubbles of brown gas (Br_2) are seen at the anode.

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

Electrolysis of water

- Inert platinum or carbon electrodes are used to carry current into and out of the electrolyte.
- Charged ions carry electric current through an electrolyte and since pure water does not contain many ions some potassium hydroxide or sulphuric acid is added to ionize the water and increase conductivity.



- Water is broken into the elements hydrogen and oxygen in the ratios 2 : 1 respectively
- While current is flowing positive hydrogen (H^+) ions migrate to the cathode. They accept electrons to form hydrogen atoms, which then combine to form molecules of hydrogen gas. They are discharged as hydrogen gas

$$\mathbf{2H^+(aq) + 2e^- \rightarrow H_2\,(g)}$$

- The negative hydroxide ions (OH^-) migrate to the anode and give up electrons forming water and oxygen molecules.

$$\mathbf{4OH^- \rightarrow 2H_2O\,(l) + O_2\,(g) + 4e^-}$$

Products formed during electrolysis of water

- Oxygen
- Hydrogen

Uses of oxygen

- Manufacture of steel
- Cutting and welding metals
- Medical purposes
- Bleaching agents
- Used by sea divers and astronauts
- Making acids e.g. sulphuric acid and nitric acid

Test for Oxygen

- Insert a glowing splint into oxygen, glowing splint rekindles/relights

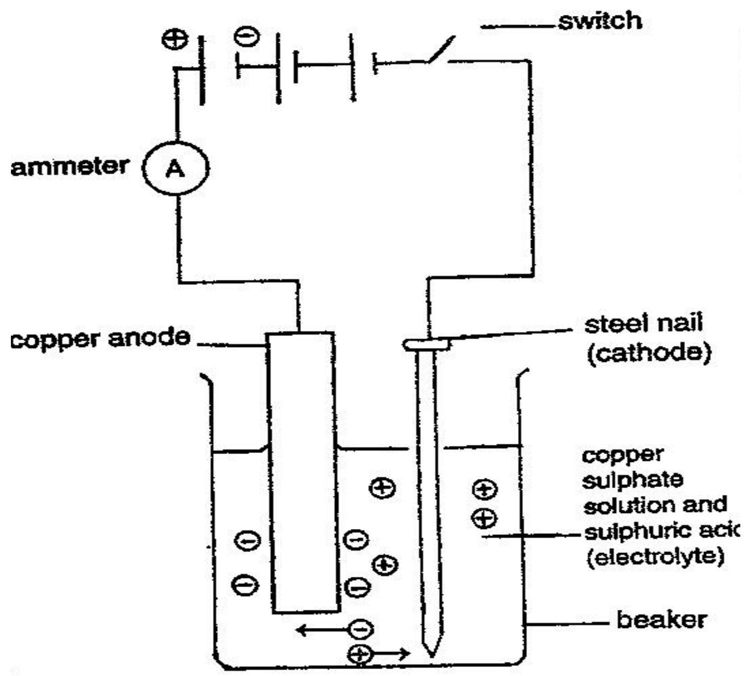
Uses of hydrogen

- Manufacture of ammonia
- Manufacture of margarine
- Welding
- Manufacture fertilisers
- Fuel

ELECTROPLATING

- Electroplating is the process during which an object of one type of metal is covered with a thin layer of another metal using electricity.
- It is a process of coating one metal with another through electrolysis

Copper plating



- Anode is copper(metal used to plate)
- Cathode is the iron nail (object to be plated).
- Acidified copper sulphate solution is the electrolyte (contains copper ions).
- The nail is cleaned well with sand paper and dipped in sulphuric acid to remove impurities and ensure that when atoms are deposited, they stick well to the nail
- When current is passed through the electrolyte, the electrolyte breaks down into ions.
- At the anode copper atoms lose electrons to become copper ions which move to the cathode where they gain electrons to become copper atoms and slowly deposits on the nail. Overtime the iron nail is covered with a thin coating of copper.

$$Fe\ (s) + CuSO_4\ (aq) \rightarrow Cu\ (s) + FeSO_4\ (aq)$$

Reasons for electroplating materials

1. Prevent corrosion or rusting
2. Decoration i.e. making an object more attractive

PRODUCTION OF AMMONIA

Conditions needed for the production of ammonia

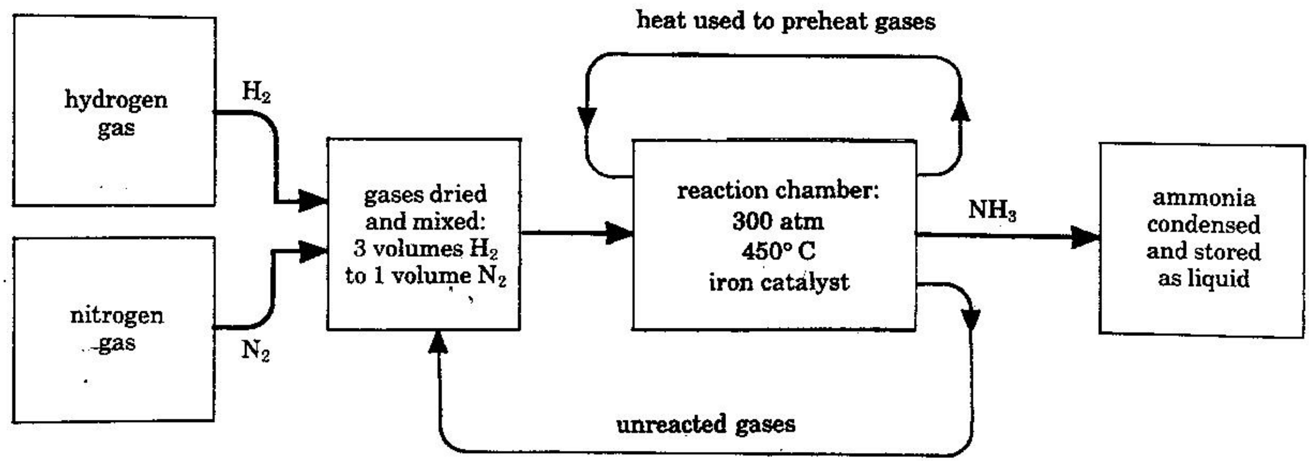
- 450 – 500^{o}C temperature

- High pressure of 200 - 300 atmosphere
- Powdered iron catalyst

Raw materials used

- Hydrogen from electrolysis of water
- Nitrogen from fractional distillation of liquid air

The Haber Process



- Nitrogen from air and hydrogen from water are reacted in ratio 1 : 3 by volume and passed over an iron catalyst. The gas pressure is raised to 200 atmospheres and the temperature to 450 – 500^{o}C. 12-17% of the mixture is converted to ammonia. The ammonia produced is separated by condensation in a cooler. Since the reaction is reversible so hydrogen and nitrogen (unconverted gases) reproduced from the decomposition of ammonia are mixed with more nitrogen and hydrogen and passed over the catalyst again.

$$N_2\,(g) + 3H_2(g) \rightarrow 2HN_3(g)$$

- Since this is a reversible reaction low temperature and high pressure is needed to enhance forward reaction. The reaction is slowed down by low temperature thus iron catalyst is used to speed up production of ammonia.

Industrial uses of ammonia

- For producing fertilisers, nitric acid, nylon, dyes, cleaners, detergents and dry cells
- Refrigerant
- Household cleaners/cleaning agents

SULPHURIC ACID

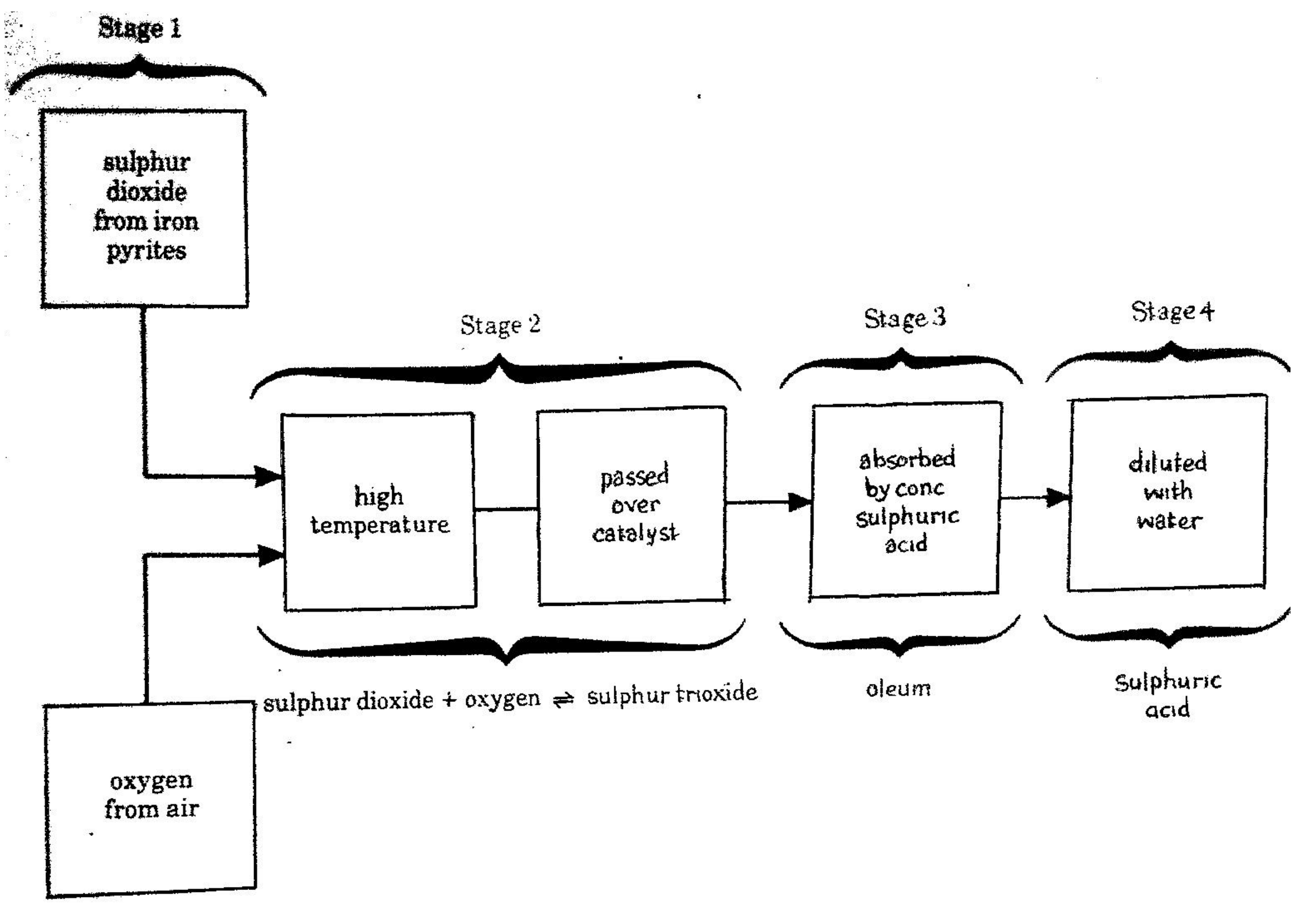
Raw materials used

- Oxygen from air or water
- Sulphur dioxide from roasting iron pyrites or sulphur in air

Conditions needed

- 450^oC temperature
- Vanadium (v) oxide
- 1 atmosphere pressure.

The Contact Process



- Sulphur or iron pyrites are roasted in air to form sulphur dioxide

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

$$2FeS\ (s) + 3O_2(g) \rightarrow Fe_2O_2(s) + SO_2(g)$$

- The sulphur dioxide is converted to sulphuric acid by the contact process. It is oxidized first by air to sulphur trioxide.

$$2SO_2\ (s) + O_2(g) \leftrightarrow 2SO_3(g) + \text{heat}$$

- Because the reaction gives out heat (exothermic) a good yield could be obtained at low temperature. Although lower tempearture is required for high yields, 450°C is instead used as the reaction will be slow at low temperature. Vanadium (v) oxide catalyst is used to increase rate of reaction. Pressure of 1-3 atmospheres is used in practise
- Sulphur trioxide is cooled and absorbed in concentrated sulphuric acid to produce oleum

$$SO_3\ (g) + H_2SO_4(aq) \rightarrow H_2S_2O_7(g)$$

    - Sulphur trioxide is not reacted with water because heat of reaction would produce a mist of sulphuric acid which difficult to condense.
- The oleum is diluted with water to give concentrated sulphuric acid

$$H_2S_2O_7\ (g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Uses of sulphuric acid

- Making of fertilizers such as superphosphate and ammonium sulphate
- 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, paint, pesticides and making detergents etc.
- Drying agent or flocculant for paper making

Production of ammonium nitrate

- Ammonium nitrate is made by direct neutralization of nitric acid by ammonia. Ammonia is bubbled through nitric acid, an acid base reaction takes place and ammonium nitrate is produced in solution.

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

- The solution is evaporated to produce a saturated solution. The saturated solution is dried in a prill tower to give solid prills (granules).
- The ammonium nitrate crystallizes out of the solution.

OXIDATION AND REDUCTION

- Oxidation is gain of oxygen or loss of hydrogen or loss of electrons
    - $\mathbf{S\ (s) + O_2(g) \rightarrow SO_2}$ Sulphur is oxidized to sulphur dioxide by gaining oxygen
    - $\mathbf{H_2S(g) + Cl_2(g) \rightarrow 2HCl(g) + S(s)}$ Hydrogen sulphide is oxidized to sulphur by losing hydrogen
    - $\mathbf{Fe \rightarrow Fe^{2+} + 2e^-}$ iron atom loses 2 electrons to form the iron (II) ion
    - $\mathbf{2Cl^- \rightarrow Cl_2 + 2e^-}$ loss of electrons by chlorine ions to form chlorine molecule
- Reduction is Loss of oxygen or gain of hydrogen or gain of electrons
    - $\mathbf{N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)}$ Nitrogen is reduced to ammonia by gaining hydrogen
    - $\mathbf{PbO(s) + H_2(g) \rightarrow H_2O(l) + Pb(s)}$ Lead oxide is reduced by losing oxygen
    - $\mathbf{Cu^{2+} + 2e^- \rightarrow Cu}$ the copper (II) ion gains 2 electrons to form neutral copper atoms
    - $\mathbf{2H^{2+} + 2e^- \rightarrow H_2}$ hydrogen ions gain electrons to form neutral hydrogen molecules
- A reducing agent is substances that removes oxygen, adds hydrogen and or donates electrons. An oxidizing agent oxidizes another substance – and is itself reduced Hydrogen and metals are reducing agents
- An oxidizing agent is a substance that adds oxygen, removes hydrogen and or accepts electrons. A reducing agent reduces another substance – and is itself oxidized. Chlorine, oxygen and metal ions are oxidizing agents

Redox reactions

- Redox reactions are simultaneous oxidation and reduction reactions In a redox reaction one substance is oxidized and the other is reduced e.g.
    - Copper oxide is reduced to copper, hydrogen is oxidised to water. Hydrogen is the reducing agent (removes O from CuO). Copper oxide is the oxidising agent (donates O to hydrogen)

$$\mathrm{CuO + H_2 \rightarrow Cu + H_2O}$$

    - The magnesium atom loses 2 electrons (oxidation) to form the magnesium ion, the iron (II) ion gains two electrons (reduced) to form iron atoms.

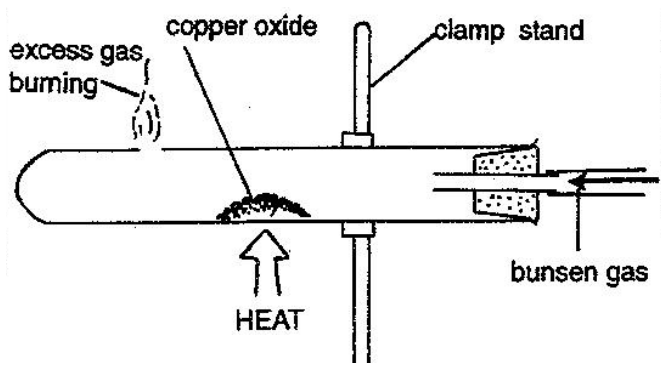
$$\mathrm{Fe^{2+} + Mg \rightarrow Fe + Mg^{2+}}$$

    - Chlorine oxidizing iron (II) chloride

$$\mathrm{2Fe^{2+} + Cl_2 \rightarrow 2Fe^{3+} + 2Cl^-}$$

- A charge on an ion is called its oxidation number. An element has a zero oxidation number because there is no charge on its atoms.
- During redox reactions there is a change in oxidation state e.g. when magnesium burns in oxygen, its oxidation state increases from 0 to +2 while oxidation state for oxygen decreases from 0 to – 2 . Magnesium is oxidized and oxygen is reduced.
- The number of electrons lost by the reducing agent is equal to the number of electrons gained by the oxidizing agent.

Reaction of copper oxide with hydrogen



- Place 3g of black CuO in a test tube and keep the apparatus horizontal and steady with the retort stand and clamps.
- Connect a rubber tube from a gas cylinder to the glass tube and allow a very slow stream of gas to flow through the apparatus.
- Heat the CuO until there is a change in colour and then remove the heat source. After heating, the Copper (II) Oxide turn pink as it was reduced to the element copper.
- Hydrogen removes oxygen from Copper (II) Oxide to form water.
- The copper oxide is reduced while hydrogen is oxidized. Hydrogen is the reducing agent while Copper Oxide is the oxidizing agent

$$CuO\ (s) + H_2(g) \rightarrow Cu\ (s) + H_2O(l)$$

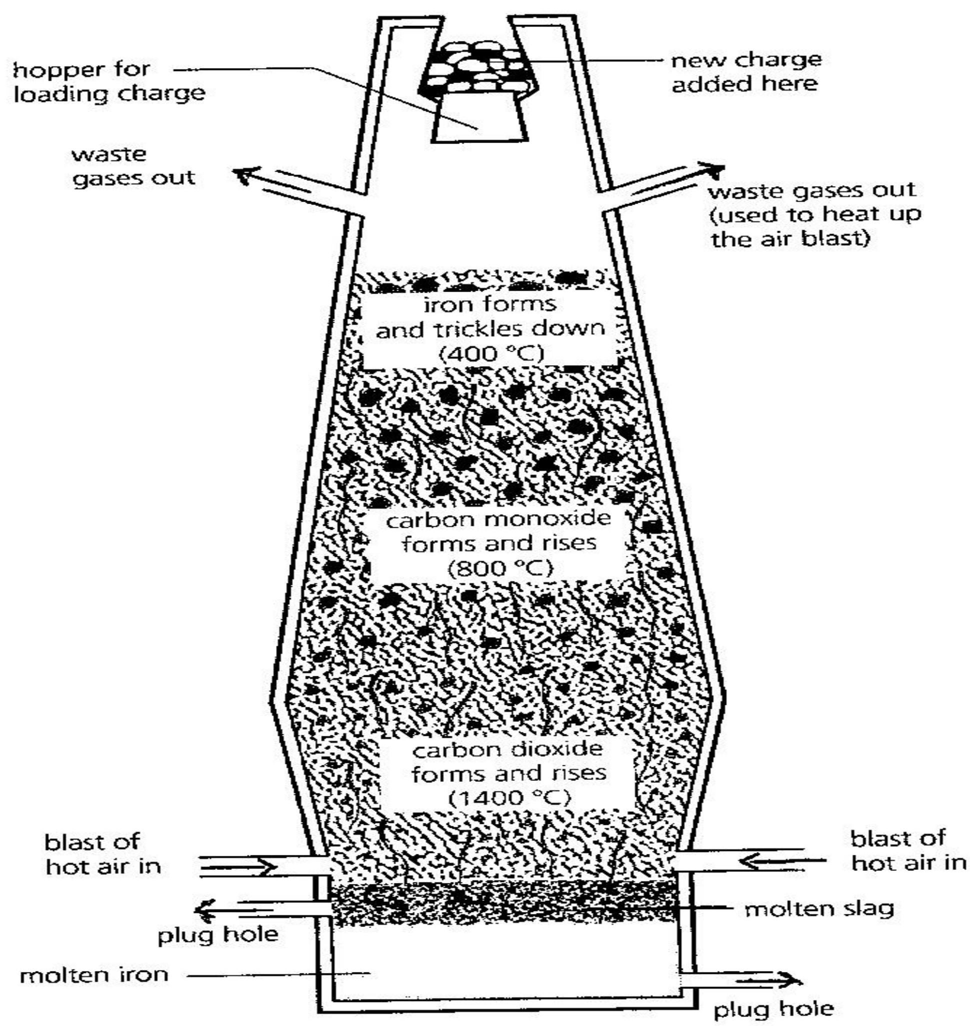
EXTRACTION OF IRON AT ZISCO

Raw materials used

Raw materials	Source	Uses
Limestone/$CaCO_3$	Buchwa/ Redcliffe	Acts as a flux i.e. it reacts with impurities to form slag
Haematite/Iron ore/Fe_2O_3	Redcliffe/Buchwa/Ripple creek	Source of iron

Coke/ Carbon	Hwange	Burns in air to produce heat and reacts to form CO which reduces Fe_2O_3
Hot air/Oxygen	Sable Chemicals	Source of oxygen to burn coke and produces heat and CO

The blast furnace



- The iron is extracted from its ore in the blast furnace by reduction of iron (II) oxide.

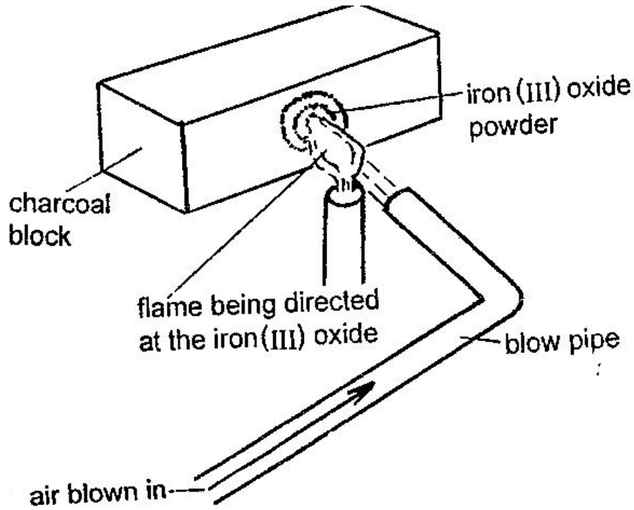
Reactions in the blast furnace

1. Near the bottom of the furnace, the carbon in coke burns in hot air, forming carbon dioxide and producing temperature up to 1700°C. Further up the furnace, the carbon dioxide is reduced to carbon monoxide
$$C\ (s) + O_2(g) \rightarrow CO_2(g) + heat$$
$$CO_2(g) + C(s) \rightarrow 2CO(g)$$
2. In the middle of the furnace the carbon monoxide reduces the iron oxides to produce molten iron, which runs down to the bottom of the furnace. The waste gases leaving the furnace are burnt to heat the air blast entering the furnace.
   Iron (II) oxide $FeO\ (s) + CO(g) \rightarrow Fe\ (l) + CO_2(g)$
   Iron (III) oxide $Fe_2O_3\ (s) + 3CO(g) \rightarrow 2Fe\ (l) + 3CO_2(g)$
3. The limestone decomposes to calcium oxide. This combines with sand/silica (SiO_2) present as an impurity in the ore and forms a slag of calcium silicate
$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2(g)$$
$$CaO\ (s) + SiO_2(s) \rightarrow CaSiO_3(l)\ (slag)$$

- The iron absorbs carbon as it moves down the furnace and this lowers its melting points. Both iron and slag are molten and drop to the bottom of the furnace.
- The less dense slag floats on top of the iron and prevents oxidation of the iron by the air blast.
- Molten iron & slag run out of separate holes from time to time

Heating Iron Oxide on a charcoal block

- A hole was made on the side of the charcoal block and a spatula full of iron oxide was placed in the hole.
- A burner was lit and the burner flame directed at the iron oxide in the hole by blowing into the blowpipe for some time.
- A small bead of greyish metal was seen at the end of the experiment, when the charcoal block had cooled.
- The iron (III) oxide was reduced to iron by carbon in the charcoal block.
$$\text{Iron (III)oxide} + \text{Carbon} \rightarrow \text{Iron} + \text{Carbon dioxide}$$



ALLOYS OF IRON

- An alloy is a mixture of metals or metals and non-metals to improve its properties.
- They are harder, more resistant to corrosion and have a more attractive appearance than the metals they are formed from.

Alloy	Composition	Properties	Uses
Cast Iron	3-5% carbon 95-97% iron	- Hard and brittle - Low strength - Cannot be welded - Rust rapidly	Used to make cookers, water pipes, hot water radiators, drain pipes, stoves, railing, steel, engines
Mild steel	0.3-0.5% carbon 99.5-99.7% iron	- Can be welded - Easily machined - Strong under tension and compression	Used to make railway lines, steel rods, beams, girders, car bodies, bridges, pipes, nails, bolts
Stainless steel	20% chromium 10% nickel 70% iron	- Resist corrosion by heat acids and rust. - Strong, tough and hard - Resist staining	- Used to make cutlery, sink units and surgical equipments

ORGANIC CHEMISTRY

ORGANIC COMPOUNDS

- Organic compounds are classified according to families of compounds that all contain the same functional group.
- A homologous series is a family of organic compound with a general formula and a similar chemical properties. Only the length of the carbon chain differs.
- A functional group of a compound is a group of atoms that give the organic molecule its physical and chemical properties.
- The name of the organic compound is determined by the number of carbon atoms it contains in the main carbon chain.
- All compounds in homologous series have functional group except alkanes
- The number of carbon atoms determines the prefix of the name and the functional group the suffix e.g.
  - 1 carbon meth-
  - 2 carbons eth-
  - 3 carbons prop-

Homologous series	General formula	Functional group
Alkanes	C_nH_{2n+2}	
Alkenes	C_nH_{2n}	C = C
Alcohols	$C_nH_{2n+1}OH$	C - OH

HYDROCARBONS

- Hydrocarbons are compounds of hydrogen and carbon only e.g. Alkanes and alkenes.

Alkanes

- They are a homologous series of molecules where the molecules are characterized by single bonds between their carbon atoms hence they are said to be saturated hydrocarbons.
- A saturated hydrocarbon have only single carbon to carbon bond because the combining capacity of the carbon atom is fully used. Each carbon is bonded to four other atoms
- They have a general formula $\mathbf{C_nH_{2n+2}}$ where n is a whole number.
- Names will always end with suffix –ane e.g. Methane (CH_4), Ethane (C_2H_6) and Propane (C_3H_8)

Alkenes

- Belongs to a homologous series that is characterized by at least on double bond between the carbon atom in the chain hence they are unsaturated hydrocarbon carbon
- An unsaturated hydrocarbon has double or tripple carbon to carbon atoms. Combining capacity of carbon atoms is not fully used e.g 2 or 3 hydrogen atoms are attached to a cabon atom.
- They have a general formula $\mathbf{C_nH_{2n}}$
- Functional group is C = C
- Name of alkenes compound always end with the suffix –ene e.g. ethene (C_2H_4) and Propene (C_3H_6)

Uses of alkanes and alkenes

Methane

- Main constituent of natural gas
- Mainly used in heating as fuel for stoves, automobile turbines and water heater

Ethane

- Used in heating
- Used for production of ethene which is used in many chemical processes

Propane

- Used as a fuel for engines and oxy-gas torches that uses fuel gases and oxygen to weld and cut metals.
- Used for portable (camping) stoves and residential central heating.

Ethene

- Used as the building block for a vast range of chemicals from plastics to anti-freeze solutions, solvent and coating on electrical wires.

Propene

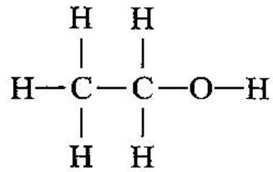
- Mainly used for the production of polypropylene, which in turn is used for textiles, furniture, packaging of food and toys.

Alcohols

- Alcohols are organic compounds which contain a hydroxyl group –OH (functional group) covalently bonded to a grouping of carbon and hydrogen atoms only.
- Their general formula is $\mathbf{C_nH_{2n+1}OH}$
- Alcohols are derived from alkanes by replacing one –H atom with an –OH group e.g.
    - Methane – methanol
    - Ethane – ethanol
    - Propane - propanol

Displayed structural formula of ethanol

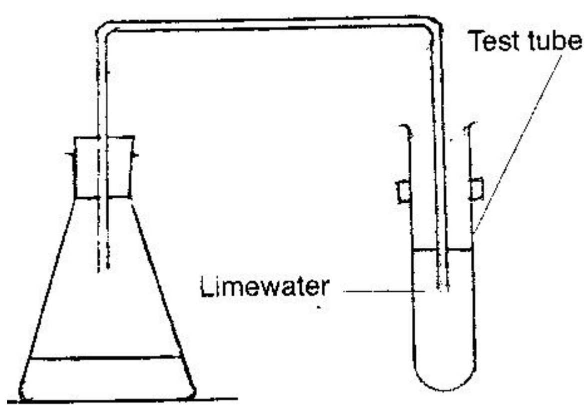
- Ethanol has the molecular formula $\mathbf{C_2H_5OH}$ or $\mathbf{CH_3CH_2OH}$



Production of Ethanol

Fermentation of sugar or maize meal solution

- Fermentation is a chemical reaction in which sugars are broken down into smaller molecules such as ethanol by yeast in the absence of oxygen e.g. fermentation of glucose to form ethanol
- Yeast is added to a sugar solution and left at room temperature for a few days in the absence of air.
- Enzymes in yeast break down glucose to ethanol and carbon dioxide, giving out heat.
- Temperature must be kept between 25 and 35^{o}C, above and below this temperature range, the enzymes become inactive.
- Sugar or sucrose is converted to glucose. Glucose is converted into ethanol and carbon dioxide
- Yeast is killed when the mixture contains more than 15% alcohol, therefore fermentation stops.



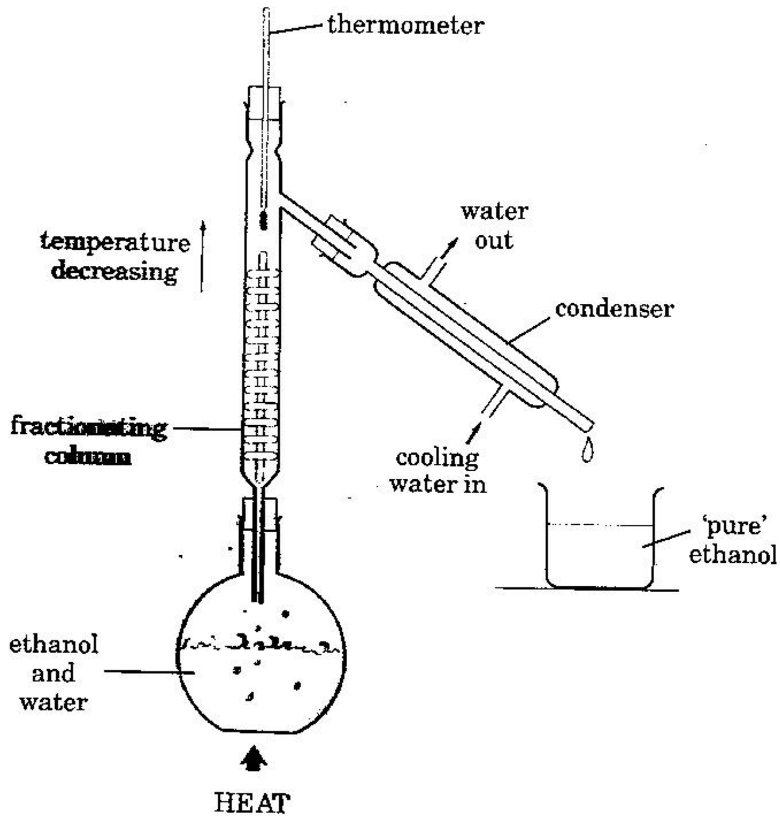
$$C_6H_{12}O_6 \rightarrow 2CH_3CH_2OH + 2CO_2$$
$$glucose \rightarrow ethanol + carbon\ dioxide$$

- Carbon dioxide is released during fermentation hence frothing is observed in the flask. Lime water changes from clear to milky.
- A dilute solution of ethanol is formed i.e. it is only about 15% concentration

Conditions for ethanol production

- pH of 6-8
- temperature of 25 – 37^oC, higher temperature will denature enzymes
- anaerobic conditions
- enzymes in yeast

Production of concentrated ethanol



- Heat the mixture in the flask. At about 78^oC the ethanol begins to boil. Some water evaporates too, so a mixture of ethanol and water vapour rises up the column.
- The vapour condenses on the glass beads in the column, making them hot.
- When the beads reach about 78^oC, ethanol vapour no longer condenses on them. Only the water vapour does. So water drips back into the flask. The ethanol vapour goes into the condenser where it condenses into pure ethanol.
- Eventually, the thermometer reading rises above 78^oC – a sign that all the ethanol has evaporated.

Uses of ethanol

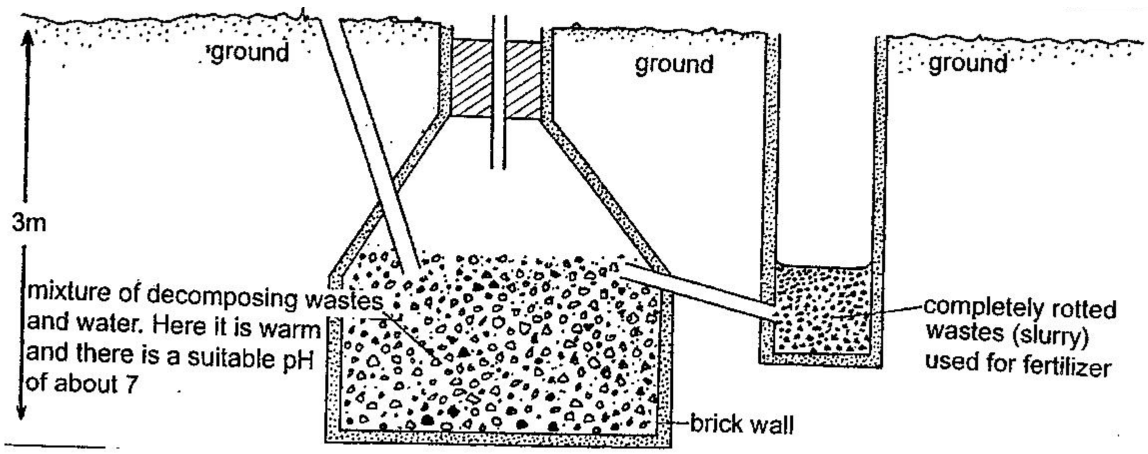
- Beverages or alcoholic drinks e.g. beer or wine

- Solvent
- Medical purposes
- To produce cosmetics, detergents, plastics and lubricants
- Fuel

BIOGAS

- Is a renewable source of energy that is made mostly of methane.
- The gas is formed when biological matter like cow manure or kitchen waste is decomposed in special biogas digesters.

Production of biogas



- Fresh cow dung is collected and then mixed with water to form sludge.
- Organic wastes are added to the digester and mixed with water.
- The anaerobic bacteria start to break down the organic wastes. Solid wastes collect at the bottom .
- The biogas rises to the top of the digester and the gas can be piped off and used in homes for cooking, lighting
- The digested sludge (slurry) is collected and used as a fertiliser for farms or gardens.
- The digester is built underground so that the heat produced during decomposition is retained and facilitates further decomposition of the cow dung

Factors affecting the production of biogas

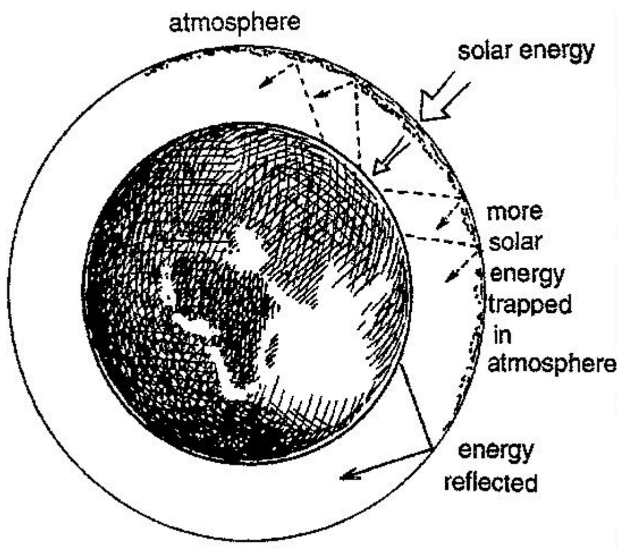
- Nature of waste
- Time
- Anaerobic bacteria
- pH between 6 & 8

- Suitable temperature of about 25-60^oC
- Air

Uses of biogas

- Cooking
- Lighting
- Operating a refrigerator

GLOBAL WARMING



- Global warming is the gradual heating of the earth that is caused by elevated levels of green house gases in the atmosphere. Green house gases absorb heat in the atmosphere and prevent it from escaping into space.
- Excessive burning of fuels such as coal and emissions from vehicles, veld fires and burning of wastes have created a buildup of carbon dioxide in the atmosphere which forms an insulating (greenhouse) effect on the earth. This prevents the sun's heat from escaping or from being reflected back into space. This effect keeps the earth warm thus increasing the atmospheric temperature in a process referred to as global warming
- Deforestation causes an increase in the carbon dioxide level in the atmosphere because there is now less vegetation to remove atmospheric carbon dioxide in photosynthesis.

Effects of global warming

- Temperature of oceans and atmosphere rise causing polar ice caps to melt. This results in higher sea level causing coastal areas to be flooded.

- Global warming is also associated with change in climatic and wind patterns bringing about change in precipitation and rainfall patterns.
- Increased air temperature (heat waves)
- Droughts

Ways of reducing global warming

- Reducing and controlling deforestation and more trees are planted
- Reducing emission of carbon dioxide into the atmosphere from burning fuels by using alternative sources of energy such as solar, water and wind power