

AQA  
GCSE  
Chemistry  
New Unit 2  
Summary  
Notes

## Atoms, Molecules and Ions.

AN ATOM is the smallest particle of an element. They cannot be split into smaller particles in chemical reactions. Iron is made of iron atoms (Fe). Sulphur is made of sulphur atoms (S)  
A MOLECULE is a small group of atoms joined together.

The atoms may be the same (e.g. O<sub>2</sub>) or different (e.g. H<sub>2</sub>O). The chemical formula shows the number and type of atoms present. Non-metal compounds are made of molecules:

Carbon dioxide contains CO<sub>2</sub> molecules

Methane (natural gas) contains CH<sub>4</sub> molecules

AN ION is an atom or group of atoms with an electrical charge (+ or -). Metal compounds such as sodium chloride or copper sulphate contain ions.

Sodium chloride is made of Na<sup>+</sup> and Cl<sup>-</sup> ions

Copper Sulphate is made of Cu<sup>2+</sup> and SO<sub>4</sub><sup>2-</sup> ions

Note that metals form positive ions while non-metals form negative ions.

A solid is represented by (s). e.g. H<sub>2</sub>O<sub>(s)</sub> is ice.

A liquid is represented by (l) e.g. Fe<sub>(l)</sub> is molten iron.

A gas is represented by (g) e.g. H<sub>2</sub>O<sub>(g)</sub> is steam.

A solution in water is represented by (aq). Salt dissolved in water is NaCl<sub>(aq)</sub>.

You should remember that the common gases are diatomic (have 2 atoms in each molecule). These are Oxygen O<sub>2</sub>; Hydrogen H<sub>2</sub>; Nitrogen N<sub>2</sub>; and Chlorine Cl<sub>2</sub>.

### Elementary Particles

Atoms are made up of smaller particles called protons, neutrons and electrons.

The protons and neutrons cluster together in a small nucleus at the centre of the atom while the electrons orbit the nucleus.

The main properties of the particles are:

Particle	Mass	Charge
PROTON	1	+1
NEUTRON	1	0
ELECTRON	Very Small	-1

Every element has an atomic number, which is the number of protons in the nucleus.

Atoms are neutral, so that the number of electrons is the same as the number of protons.

The mass of an atom is almost entirely made up of protons and neutrons, which have the same mass as each other (the mass of electrons can be ignored).

These relations can be summarised as:

ATOMIC NUMBER = NUMBER OF PROTONS

NUMBER OF PROTONS = NUMBER OF ELECTRONS

MASS OF ATOM = PROTONS + NEUTRONS

Element	Atomic Number	Protons	Electrons	Neutrons	Mass Number
Na	11	11	11	12	23
C	6	6	6	6	12
U	92	92	92	146	238

## Electronic Structure

The electrons orbit the nucleus in 'shells'. These can hold the following numbers of electrons:

the innermost shell can contain up to 2 electrons

the next shell can contain up to 8 electrons

the next shell can contain up to 8 electrons (although this can be expanded up to 18)

Fluorine has atomic number 9. Its electronic structure is 2.7

Sodium has atomic number 11. Its electronic structure is 2.8.1

Calcium has atomic number 20. Its electronic structure is 2.8.8.2

## Ionic and Covalent Bonding

Noble gases have 2 things in common:

- a) They are very unreactive (stable)
- b) They have complete outer shells of electrons.

Other elements have a tendency to change their electronic structures to become more like noble gases and therefore more stable. They can do this by losing, gaining or sharing electrons.

### Ionic Bonding

1. Ionic bonding takes place between a metal and a non-metal.
2. The metal atoms lose their outer shell electrons to form positive ions.
3. The non-metal atoms gain these electrons to fill their outer shells and form negative ions
4. The positive and negative ions attract each other in a lattice

### Covalent Bonding

1. Covalent bonds are formed between 2 non-metallic elements.
2. The atoms share electrons in order to complete their outer shells.
3. The atoms all attain noble gas structure (complete outer shells).
4. The new particles formed are neutral molecules.

## Structures of Substances

There are 4 main structures which substances can have. These are known as:

SIMPLE MOLECULAR., GIANT COVALENT, GIANT IONIC and GIANT METALLIC

### Simple Molecular.

Simple molecular substances have small molecules, such as  $\text{H}_2\text{O}$  or  $\text{CO}_2$  (see diagrams on p 77 & 78). The atoms in these molecules are held together by strong forces called covalent bonds, but there are only very weak forces between the molecules. This means:

- 1) They have low melting and boiling points (many are liquids or gases).
- 2) They tend to be soft and/or have little strength.
- 3) They do not conduct electricity

Simple molecular substances are nearly always non-metallic elements or their compounds.

e.g. Water ( $\text{H}_2\text{O}$ ), Iodine ( $\text{I}_2$ ), Carbon dioxide ( $\text{CO}_2$ ), Sugar ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ ).

### Giant Covalent Substances

In these substances, strong covalent bonds join atoms together in large numbers to make giant structures. Sand (silicon dioxide) is one example, diamond and graphite (both forms of carbon) are others. Because the bonds between all the atoms are very strong:

- 1) They have very high melting points.
- 2) They are very hard (graphite is an exception)
- 3) They do not conduct electricity (graphite is an exception)
- 4) They are insoluble in all solvents.

The important examples are diamond (C), graphite (C) and silicon dioxide (SiO<sub>2</sub>). The last-mentioned is the base for the structure of many rocks, such as flint or granite.

### Giant Metallic Structures

The atoms in metals are in layers which can slide over each other, this makes it possible to bend them or beat them into shape.

The atoms in metals share their outer electrons with all the other metal atoms, so that a metal consists of positive ions held together by free electrons which can move throughout the structure. Like other giant structures, the forces (called metallic bonds) holding the atoms together are strong. The main properties of metals are:

- 1) Metals are strong.
- 2) Most metals have high melting points.
- 3) Metals are malleable (they can be bent or beaten into different shapes)
- 4) Metals are good conductors of electricity & heat
- 5) Metals are lustrous (shiny)

### Giant Ionic Structures

Compounds of metals, such as sodium chloride (NaCl) or copper sulphate (CuSO<sub>4</sub>) are made up of positive metal ions (e.g. Na<sup>+</sup>) and negative non-metal ions (e.g. Cl<sup>-</sup>).

These are arranged in a regular lattice so that each positive ion is surrounded by negative ions and each negative ion is surrounded by positive ions. The particles are held together by the strong attraction of + for -. The main properties of ionic substances are:

- 1) They have high melting and boiling points.
- 2) They are hard but brittle (shatter easily)
- 3) They conduct electricity *only when melted or dissolved in water.*
- 4) Many ionic substances are soluble in water.

### Calculations Summary

Formula masses (Ar and Mr)

The Ar is the relative atomic mass in the periodic table. The Ar values of the atoms in a formula are added to get the formula mass or Mr

$$\text{H}_2\text{SO}_4 = 1 \times 2 + 32 + 16 \times 4 = 98$$

$$\text{Mg}(\text{NO}_3)_2 = 24 + (14 + 3 \times 16) \times 2 = 148$$

### Percentage Calculations

If you need to find the percentage of an element E in a compound, you use formula

$$\text{Percentage} = \frac{\text{Ar of E} \times \text{No of atoms of A}}{\text{Mr of compound}} \times 100$$

e.g. Find the percentage of nitrogen in ammonium nitrate ( $\text{NH}_4\text{NO}_3$ )

a) Mr of N = 15

b) No of N atoms = 2

c) Mr of ammonium nitrate =  $14 + 4 \times 1 + 14 + 3 \times 16 = 80$

Percentage of nitrogen =  $\frac{2 \times 14}{80} \times 100 = 35\%$

80

#### Empirical Formulae

This is the formula which shows the lowest whole number ratio of the atoms

e.g. Molecular formula =  $\text{C}_2\text{H}_4$ , Empirical Formula =  $\text{CH}_2$

To calculate an empirical formula

- 1) Find the mass (or %) of each element present
- 2) Divide each of these masses by the relative mass of that element
- 3) Divide each number obtained in stage 2 by the smallest of those numbers.
- 4) This should give whole numbers which can be used in the empirical formula.

20g of a compound of Silicon with hydrogen contains 17.5g of silicon. Find the empirical formula.

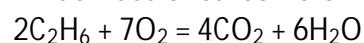
Element	Si	H
Mass of element present	17.5	$20 - 17.5 = 2.5$
Relative Mass (Ar)	28	1
Mass $\div$ Ar	$17.5 \div 28 = 0.625$	$2.5 \div 1 = 2.5$
Divide through by smallest.	$0.625 \div 0.625 = 1$	$2.5 \div 0.625 = 4$
Whole Number Ratio	1	4
Empirical Formula	$\text{SiH}_4$	

#### Equation Calculations

In these, you will always be given the mass of one substance (call it A) and be asked to find the mass of another substance (B)

You must divide the mass of A by the formula mass of all the molecules of A and multiply by the formula mass of all the molecules of B. (the mass of all the molecules of a substance in the equation is the formula mass time the big number)

What mass of carbon dioxide is obtained by burning 3g of  $\text{C}_2\text{H}_6$ ?



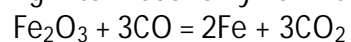
Mr:  $\text{C}_2\text{H}_6 = 30$ ,  $\text{CO}_2 = 44$

Answer =  $\frac{3}{2 \times 30} \times (4 \times 44) = 0.05 \times 176 = 8.8\text{g}$

#### Atom Economy

This is given as  $\frac{\text{Formula Mass of all molecules of Useful Product}}{\text{Formula Mass of all molecules of Reactants}} \times 100$

Eg. Atom economy for making iron from the equation



(Mr :  $\text{Fe}_2\text{O}_3 = 160$ ,  $\text{CO} = 28$ ,  $\text{Fe} = 56$ ,  $\text{CO}_2 = 44$ )

Mass of useful product (iron atoms) =  $2 \times 56 = 112$

Mass of all reactants =  $160 + (3 \times 28) = 244$

Atom economy =  $\frac{112}{244} \times 100 = 45.9\%$

### Percentage Yield.

To find the percentage yield, divide the mass obtained in an experiment by the maximum mass which could have been obtained and multiply by 100.

(the maximum mass is the amount calculated in an equation type calculation – see above)  
e.g. the calculated mass of alcohol which could be obtained by fermenting some sugar is 4.6g of alcohol. The actual mass obtained in an experiment was 3.5g. Find the % yield.

Answer =  $\frac{3.5}{4.6} \times 100 = 76.1\%$

### The Mole

The mole is an important quantity used in all calculations you will meet later in chemistry.

What you need to know now is

- 1) A mole of any substance contains approximately  $6 \times 10^{23}$  particles (this number is called the Avogadro constant)
- 2) The mass of 1 mole of any substance is its formula mass (Mr) in grams.
- 3) 1 mole of any gas occupies  $24 \text{ dm}^3$  (24 litres) at room temperature and pressure.

e.g. 1 mole of  $\text{CO}_2$  (Mr =44) contains  $6 \times 10^{23}$  molecules and weighs 44g. As a gas its volume at RTP will be  $24 \text{ dm}^3$

Being able to relate the number of molecules, the mass and the volume is very powerful.

### Energy Changes in Reactions and Equilibrium

Reactions which give out energy are called Exothermic. They cause the temperature to rise. Many reactions are exothermic, including all burning (combustion) reactions

e.g.  $2\text{Mg} + \text{O}_2 = 2\text{MgO}$

Reactions which take in energy are called Endothermic. They may cause the temperature to fall. Thermal decompositions are always endothermic

e.g.  $\text{CaCO}_3 = \text{CaO} + \text{CO}_2$

Some reactions require an initial input of energy to start the reaction (called the activation energy) but are exothermic. E.g. the match to light the fire.

### Reversible reactions

If a reaction is exothermic in one direction, it will be endothermic in the other direction and the amount of heat given out by the exothermic reaction will be equal to the amount of heat taken in by the endothermic reaction.

e.g. HYDRATED COPPER SULPHATE ? ANHYDROUS COPPER SULPHATE + WATER  
(Blue) (White)

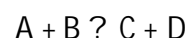
If the hydrated copper sulphate is heated, it turns white as the water is given off. This requires energy so it is endothermic

When water is added to anhydrous copper sulphate, it turns blue and heat is given out as the reaction is exothermic. This reaction is sometimes used as a test for water.

In any reversible reaction, a rise in temperature will encourage the endothermic reaction, while reducing the temperature will encourage the exothermic reaction.

### Equilibrium

Many reactions are reversible. e.g. for the reaction



A and B can react to form C + D, but C and D can also react to form A + B. Generally both forward and reverse reactions will occur at the same time.

At the start of the reaction, concentrations of both A & B will be high, so the forward reaction will be fast. As the reaction proceeds, the concentrations of A & B both fall, so that the forward reaction gets slower.

Similarly, at the start of the reaction, concentrations of C & D will be low, but as the reaction proceeds their concentrations will increase and so the reverse reaction becomes faster.

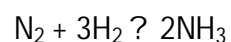
Eventually, as the forward reaction gets slower and the reverse reaction gets faster, the rates will eventually become equal. This is known as equilibrium.

When this state is achieved, the amounts of both reactant and product remain constant.

The conditions of the reaction, such as pressure, temperature and concentration will determine where and how equilibrium occurs. Manufacturers of chemicals need knowledge about this so that they can make the maximum amount of their chemicals and discourage reverse reactions.

In particular:

a) Pressure. Increasing the pressure will encourage the reaction to go in the direction of fewest molecules (the fewer the molecules, the less space is taken up), e.g. in the reaction to form ammonia:



High pressure encourages the forward reaction, since the number of molecules goes from 4 to 2

b) Concentration. Adding more reactant or removing the product will encourage the forward reaction

c) Controlling temperature is also very important. Raising temperature will cause a reversible reaction to go in the endothermic direction, while cooling will cause it to go in the exothermic direction.

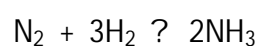
## Ammonia

### Manufacture of Ammonia

1) Ammonia is made from nitrogen and hydrogen. Nitrogen is separated from air, while hydrogen is obtained by the cracking of oil.

2) The equation for the reaction is:

Nitrogen + Hydrogen  $\rightleftharpoons$  Ammonia (exothermic)



3) The conditions needed for the reaction are 200 atmospheres pressure, 450°C and an iron catalyst.

- a) The pressure brings the molecules closer together, so that they react faster and also give a greater yield of ammonia since the forward reaction is encouraged. However, pumps to provide high pressure are expensive.
- b) High temperature makes molecules move faster, so that they collide more often and more energetically, and the reaction goes faster. However, because the forward reaction is exothermic, a high temperature also tends to drive the reaction in the reverse direction, so that less ammonia is produced. 450C is a sort of compromise to produce the best results.
- c) A catalyst speeds up the reaction, but is not changed itself.
- d) Removal of the ammonia is also favourable to the forward reaction.

4) Under these conditions, about 28% of the gases are converted to ammonia. This can be separated from the unreacted nitrogen and hydrogen by cooling. Ammonia turns to a liquid at -33C, while the other gases do not turn to liquids until much lower temperatures. The unreacted gases can then be recycled to make more ammonia.

#### Properties of Ammonia (NH<sub>3</sub>)

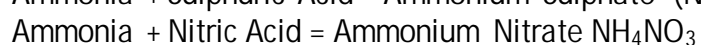
It is a colourless gas with a pungent smell.

It is alkaline (turns pH paper blue).

It is very soluble in water.

#### Uses of Ammonia

Most ammonia is converted to fertilisers, which provide nitrogen for plants. This is achieved by reacting with acids.



### Rates of Reaction

There are 4 main ways of controlling the rate of a reaction:

#### CATALYSTS

Catalysts are substances which increase the rate of chemical reactions without being chemically changed. They do this by providing a different route for the reaction with a lower activation energy.

Enzymes are biological catalysts.

Other examples of catalysts are:

NAME	PURPOSE
Aluminium Oxide	Cracking of Oil
Iron	Manufacture of ammonia
Vanadium Pentoxide	Manufacture of sulphuric acid.

Most of these catalysts are solid substances, used to catalyse reactions between gases. They do this by allowing the gas molecules to collect on their surface where they are close enough to react quickly.



Catalysts are very valuable in chemical industry, since they can be reused and they provide a much cheaper way of speeding up a reaction than heating, so providing cheaper goods.

#### TEMPERATURE

Another way of controlling the rate of a reaction is by temperature. A rise of  $10^{\circ}\text{C}$  will roughly double the rate of a reaction. This is because at higher temperature, the particles move faster, they therefore collide more often and more energetically, making the reaction faster.

Cooling down a reaction will make it slower, this means that food will deteriorate more slowly at low temperatures (e.g. in a fridge or freezer)

Remember that raising temperature does not make any difference to the amount produced, only to the speed of the reaction.

#### CONCENTRATION

The concentration of a solution is the amount of solute present in a solution. The concentration is usually measured in moles per  $\text{dm}^3$  ( $1\text{dm}^3 = 1\text{ litre}$ )

This means that equal volumes of solution of the same concentration contain equal number of particles.

e.g.  $25\text{cm}^3$  of a solution of HCl contains the same number of particles as  $25\text{cm}^3$  of a solution of NaOH if they both have a concentration of  $0.1\text{ moles}/\text{dm}^3$

Reactions happen more quickly when concentrated solutions are used because the particles of reactants are closer and so are more likely to collide and react.

#### PRESSURE

When gases are at high pressure, their particles are closer, and so are more likely to collide and react.

Note that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

#### SURFACE AREA

Solids in powder form will react much more quickly than those in lump form. This is because the powder particles have much more surface area exposed to the other reactant. If a lump is used, only the outside of that lump can react.

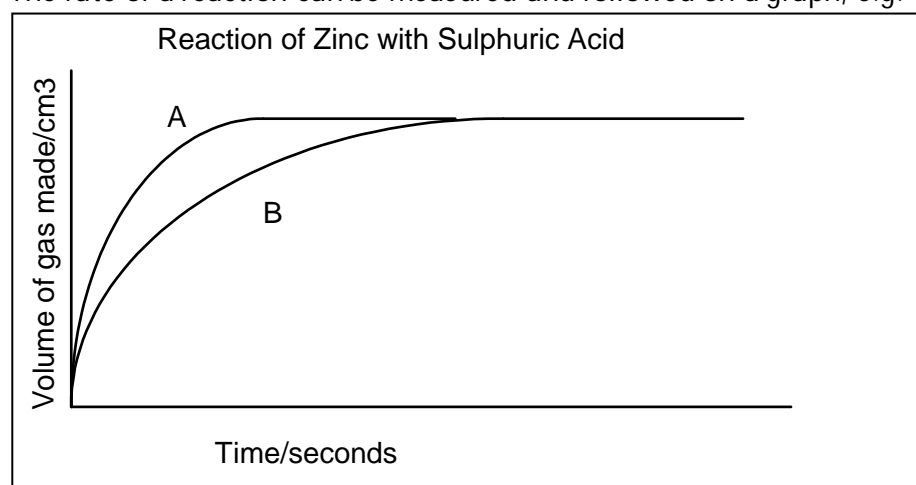
To summarise, if we take a reaction such as the reaction between zinc and sulphuric acid:



The rate of this reaction can be measured by seeing how much hydrogen gas is given off in one minute. This rate can be increased by:

- 1) Raising the temperature of the reaction.
- 2) Using more concentrated sulphuric acid.
- 3) Using powdered zinc rather than lumps.
- 4) Adding some copper, which acts as a catalyst.

The rate of a reaction can be measured and followed on a graph, e.g.



Curve A represents a reaction between powdered zinc and sulphuric acid, while curve B represents a reaction between lumps of zinc and the same amount of sulphuric acid. You should note:

- That a steeper curve (A) represents a faster reaction.
- That the curves start steeply, but then level off. This is because the concentration of the acid is high at the start of the reaction, but gets less as the acid is used up and eventually stops.
- That the amount of gas made by the end of the reaction is the same as long as the amounts of reactants are the same. The rate of a reaction does not affect the final amount of product made.

#### MEASURING THE RATE OF A REACTION

The rate of a reaction is defined as

$$\text{Rate} = \frac{\text{Amount of Reactant Used up or Amount of Product Formed}}{\text{Time}}$$

A reaction can be followed in various ways, e.g. by seeing how much mass is lost in a certain time, or how much gas is produced in a certain time.

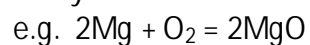
For instance, a number of similar experiments can be done under similar conditions, e.g. when reacting limestone with hydrochloric acid, different concentrations of hydrochloric acid could be used and the time taken to collect a boiling-tube of gas ( $50\text{cm}^3$ ) could be compared.

Also, the rate of a reaction changes as it proceeds. When the reaction starts, the reactants are quite concentrated, but as it goes on, the reactants become less concentrated, and the reaction slows down and eventually stops.

#### Energy changes in reactions

Reactions which give out energy are called Exothermic. They cause the temperature to rise.

Many reactions are exothermic, including all burning (combustion) reactions



Reactions which take in energy are called Endothermic. They may cause the temperature to fall. Thermal decompositions are always endothermic

e.g.  $\text{CaCO}_3 = \text{CaO} + \text{CO}_2$

Some reactions require an initial input of energy to start the reaction (called the activation energy) but are exothermic. E.g. the match to light the fire.

Reversible reactions

If a reaction is exothermic in one direction, it will be endothermic in the other direction and the amount of heat given out by the exothermic reaction will be equal to the amount of heat taken in by the endothermic reaction.

e.g. HYDRATED COPPER SULPHATE = ANHYDROUS COPPER SULPHATE + WATER  
(Blue)(White)

If the hydrated copper sulphate is heated, it turns white as the water is given off. This requires energy so it is endothermic

When water is added to anhydrous copper sulphate, it turns blue and heat is given out as the reaction is exothermic. This reaction is sometimes used as a test for water.

In any reversible reaction, a rise in temperature will encourage the endothermic reaction, while reducing the temperature will encourage the exothermic reaction.

In the Haber process, the reaction  $\text{N}_2 + 3\text{H}_2 = 2\text{NH}_3$  is exothermic in the forward direction, and endothermic in the reverse direction

Therefore raising the temperature favours the reverse reaction and gives a poor yield of ammonia (but speeds the process up).

## ELECTROCHEMISTRY

Conductors

Metals and graphite are the only solids which conduct electricity, but no chemical change is involved. Liquid (melted) metals also conduct, but again there is no chemical change.

Electrolytes

These are liquids which conduct electricity, and are decomposed by it. They are ionic substances which are dissolved in water or have been melted. This includes all acids and metal compounds.

Examples: Copper sulphate solution, iron chloride solution, molten sodium chloride, dilute sulphuric acid.

Non-electrolytes are covalent substances, e.g. pure water, sugar solution, alcohol, petrol.

Electrolysis

This is when an electric current passes through an electrolyte.

Electrons enter the solution through the negative electrode (cathode), cause a chemical change and leave by the positive electrode (anode).

Molten electrolytes are split into their elements by electrolysis.

The metal is produced at the cathode (-), while the non-metal is produced at the anode (+)  
e.g. Lead Bromide (molten) = Lead (at the cathode) + Bromine (at the anode)

With aqueous electrolytes, the electrolyte is also split up, but if the metal is reactive, then hydrogen from the water is produced at the cathode in place of the metal.

e.g. Copper chloride (aq) = Copper (at the cathode) + Chlorine (at the anode)

e.g. Sodium chloride (aq) = Hydrogen (at the cathode) + Chlorine (at the anode)

The lower in the reactivity table a metal is (including hydrogen), the more likely it is to be released at the cathode.

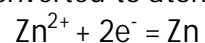
Because positive ions go to the cathode, they are called Cations

Because negative ions go to the anode, they are called Anions

Electrolysis of molten Zinc Chloride

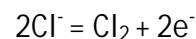
The main ions present in zinc chloride are  $Zn^{2+}$  and  $Cl^-$ .

The  $Zn^{2+}$  ions are attracted to the negative cathode. Here they pick up electrons, and the ions are converted to atoms.



The zinc metal is then deposited on the cathode.

The  $Cl^-$  ions are attracted to the positive anode. Here they lose electrons to form chlorine atoms. These atoms pair up to form chlorine molecules. Chlorine gas is given off at the anode.



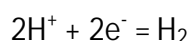
The products are zinc & chlorine.

Electrolysis of Sodium Chloride Solution

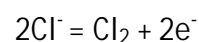
The main ions present in sodium chloride solution are  $Na^+$  and  $Cl^-$ , but there are also a few  $H^+$  and  $OH^-$  ion present because water is very slightly ionised.

The  $Na^+$  ions and  $H^+$  ions are attracted to the negative cathode. Here the  $H^+$  ions pick up electrons, since hydrogen is less reactive than sodium.

The hydrogen atoms take electrons to form hydrogen atoms, which then pair up to form hydrogen molecules.



The  $Cl^-$  ions are attracted to the positive anode. Here they lose electrons to form chlorine atoms. These atoms pair up to form chlorine molecules. Chlorine gas is given off at the anode.



The products are hydrogen & chlorine, but  $Na^+$  and  $OH^-$  ions are left in solution to make sodium hydroxide.

Uses of products:

Chlorine: Purifying water, making PVC plastic.

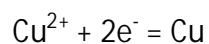
Hydrogen: Making margarine or ammonia.

Sodium Hydroxide: Making soap.

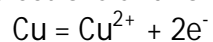
Electrolysis of copper sulphate solution. (copper electrodes).

The following ions are present in copper sulphate solution:  $Cu^{2+}$ ,  $SO_4^{2-}$  (from  $CuSO_4$ )  $H^+$ ,  $OH^-$  (from water).

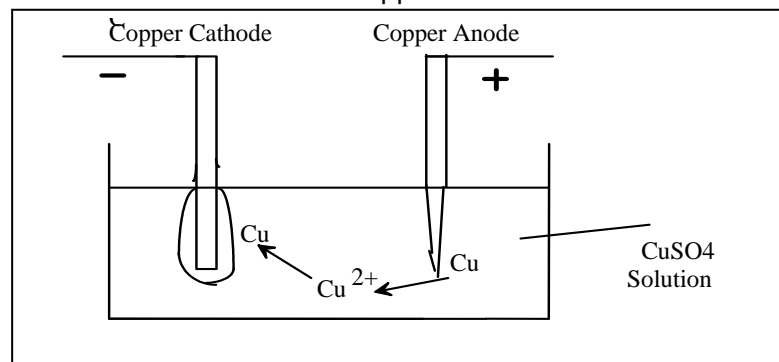
As above, the  $H^+$  ions and  $Cu^{2+}$  are attracted to the cathode but Cu gains electrons more easily so that copper is deposited



At the anode, rather than the sulphate or hydroxide ions releasing their electrons, the atoms in the copper anode release their electrons and form ions.



Therefore at the cathode, copper ions from solution are being deposited, and the cathode increases in mass. At the anode the copper atoms from the anode are going into solution as copper ions. The concentration of copper ions in solution remains unchanged.



This is used industrially to purify copper. A lump of impure copper is used as the anode and a rod of pure copper as the cathode. Copper is dissolved away from the anode, as above, leaving behind any impurities. Pure copper is deposited on the cathode.

#### Other Uses of Electrolysis

Many of the more reactive metals (e.g. sodium, magnesium, calcium) can only be extracted from their ores by electrolysis.

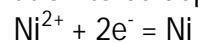
Electroplating allows a thin layer of one metal to be deposited on another.

e.g. Iron can be protected from rust by coating with chromium or nickel, or a cheap metal can be coated with silver or gold to make it look expensive.

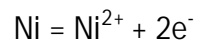
The object to be plated is made the cathode, while the metal to coat it is made the anode. The coating metal is also present in the solution.

e.g. to nickel plate a piece of iron, the iron would be the cathode, the nickel would be the anode and the solution would be nickel sulphate.

Electrolysis would cause  $\text{Ni}^{2+}$  ions in solution to be deposited on the iron cathode.



These ions would then be replaced by Nickel metal slowly dissolving from the anode.



## Acids Bases and Alkalis

### Acids

- 1) They are corrosive.
- 2) They have a pH of less than 7
- 3) They react with and neutralise bases.
- 4) They produce hydrogen ( $H^+$ ) ions in water.

### Examples

Sulphuric acid  $H_2SO_4$

Hydrochloric acid  $HCl$

Nitric acid  $HNO_3$

Ethanoic acid  $CH_3CO_2H$

Acids can also be dilute or concentrated. This is not the same as weak/strong and depends simply on the amount of water present.

### Bases

- 1) Bases react with and neutralise acids.
- 2) They are generally solids, but may be used in solution.
- 3) Metal oxides and metal hydroxides are bases
- 4) ALKALIS are soluble bases
- 5) Alkalis produce  $OH^-$  ions in water.

### Examples

- 1) These are soluble and are therefore alkalis as well as bases

Sodium hydroxide  $NaOH$

Calcium hydroxide  $Ca(OH)_2$

Potassium hydroxide  $KOH$

Ammonia  $NH_3$

- 2) These bases are insoluble

Magnesium oxide  $MgO$

Copper oxide  $CuO$

### Salts

These are crystalline substances, generally compounds of metals. They are formed when an acid neutralises a base. The salt is formed when hydrogen in an acid is replaced by metal atoms.

e.g. Sodium chloride  $NaCl$  (A salt of hydrochloric acid,  $HCl$ )

Copper sulphate  $CuSO_4$  (A salt of sulphuric acid  $H_2SO_4$ )

Potassium nitrate  $KNO_3$  (A salt of nitric acid  $HNO_3$ )

### Reactions of Acids

- 1) Acids react with and neutralise bases

ACID + BASES = SALT + WATER

The name of the salt comes from the metal in the base and a word derive from the acid

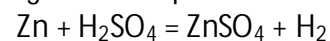
- Sulphuric acid gives sulphates
- Nitric acid gives nitrates
- Hydrochloric acid gives chlorides

e.g. Sodium hydroxide + sulphuric acid = sodium sulphate + water  
 $2\text{NaOH} + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$

2) Acids react with metals (except for the unreactive metals such as copper, silver and gold)

ACID + METAL = SALT + HYDROGEN

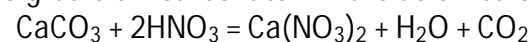
e.g. Zinc + sulphuric acid = zinc sulphate + hydrogen



3) Acids react with carbonates and hydrogen carbonates

ACID + CARBONATE = SALT + WATER + CARBON DIOXIDE

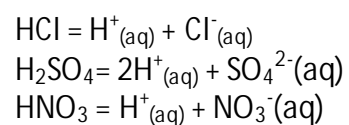
e.g. Calcium carbonate + nitric acid = calcium nitrate + water + carbon dioxide



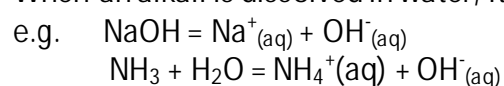
Behaviour of Acids in Water

Without water, most acids are simple molecular, covalent compounds, but when an acid is dissolved in water it produces ions. Therefore acids only behave as acids when water is present.

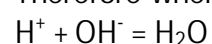
e.g.



When an alkali is dissolved in water, it produces dissolved  $\text{OH}^-$  ions.



Therefore when an acid reacts with an alkali, the ionic equation is:



This is the ionic equation for all acid - alkali reactions.

Neutralisation

To find out the neutralisation point for an acid-alkali reaction, various techniques can be used:

- 1) When Universal Indicator becomes pale green (pH7)
- 2) When a pH meter indicates that the pH is 7

Making Salts

Before deciding on the best method to make a salt, one must know the solubility of different types of compound

TYPE OF COMPOUND	SOLUBLE	INSOLUBLE
Nitrates ( $\text{NO}_3^-$ )	All	
Sulphates ( $\text{SO}_4^{2-}$ )	Most	$\text{PbSO}_4$ , $\text{BaSO}_4$
Chlorides ( $\text{Cl}^-$ ), Bromides ( $\text{Br}^-$ ), Iodides ( $\text{I}^-$ )	Most	Lead & Silver halides
Carbonates ( $\text{CO}_3^{2-}$ )	$\text{Na}_2\text{CO}_3$ , $\text{K}_2\text{CO}_3$	Most
Hydroxides ( $\text{OH}^-$ )	$\text{NaOH}$ , $\text{KOH}$	Most
Oxides		All

Methods are used for making salts

Method 1. Used to make sodium, potassium & ammonium salts. (Acid + alkali)

- 1) The alkali is measured out using a pipette, indicator is added, and enough acid is added from a burette to just change the colour of the indicator.
- 2) The volumes of acid and alkali used are noted, and the experiment is repeated using the same volumes, but no indicator.
- 3) The solution is evaporated to leave the salt

Method 2 Where an acid is reacted with an insoluble substance (base, metal or carbonate).

- 1) Some acid is measured into a beaker and warmed.
- 2) The base (or carbonate or metal) is added a little at a time until no more will dissolve.
- 3) The solution is filtered to remove any unreacted material, and the filtrate is left to evaporate, leaving the pure salt.

Method 3. Used to make an insoluble salt

- 1) We find solutions which contain the two halves of the salt.
- 2) These solutions are mixed and form a precipitate, which can then be filtered off.
- 3) The precipitate is washed and dried

e.g. to make silver, we mix solutions of silver nitrate and sodium chloride.

Silver Nitrate + Sodium Chloride = Silver Chloride + Sodium Nitrate

