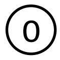
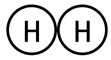
UNIT 1

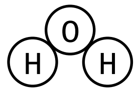
Introduction

Atom : smallest kind of particle

e.g. Oxygen atom 

Molecule : cluster of atoms bonded together

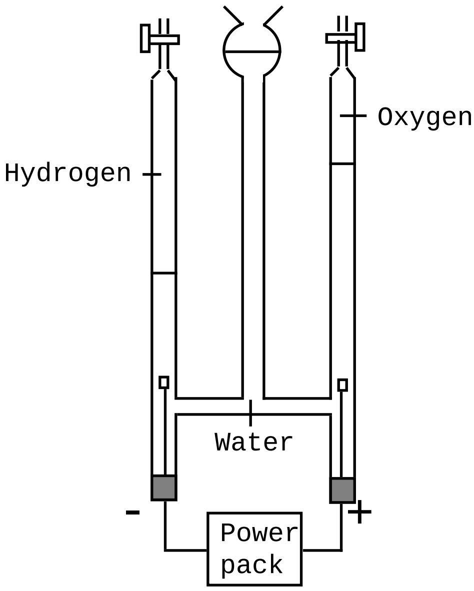
e.g. Hydrogen molecule H_2 

Water molecule H_2O 

Element : contains only **one** kind of atom e.g. Hydrogen or Oxygen

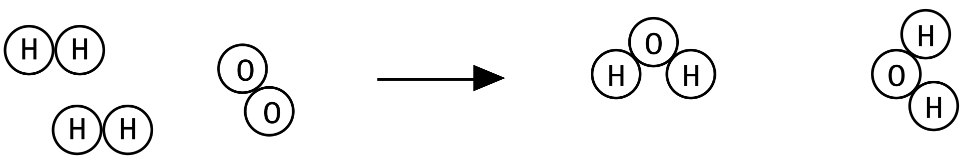
Compound : contains **different** atoms <u>bonded</u> together
e.g. Carbon dioxide

Water H_2O



Water is a compound consisting of two atoms of Hydrogen bonded to one atom of Oxygen. We can decompose Water into Hydrogen and Oxygen by putting energy into it - electrical energy. Water is a poor conductor so a little Sodium fluoride is added to make it conduct better. Sodium fluoride is a compound of Sodium and Fluorine so the Hydrogen and Oxygen must have come from the Water.

The Oxygen relights a glowing splint. The Hydrogen is lighter than air and burns with a 'pop'. When Hydrogen burns it combines with the Oxygen in the air and reforms Water (giving **out** energy) :



2 H_2 + O_2 -> **2** H_2O + energy

Three Types of Solution

Solutions can be classified according to their pH, a scale ranging from 0 to 14.

Neutral solutions have a pH of 7

Acidic solutions have a pH less than 7

Alkaline solutions have a pH greater than 7

The pH can be measured using pH paper.
A solution of the substance should be made in water.
The pH paper should be dipped in the solution and its colour matched against a chart giving pH values e.g.

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Type of solution	Strong acid				Weak acid			Neutral	Weak alkali				Strong alkali		
Colour of pH paper	Red		Orange			Yellow		Yellow green	Green		Green blue		Blue	Purple	

Substance	Colour of pH paper	pH	Type of solution
Baking soda solution	Green	8	weak alkali
Vinegar	Orange	4	weak acid
Milk of magnesia	Green	8	weak alkali
Potassium hydroxide solution	Purple	14	strong alkali
Lime water	Blue/green	10	weak alkali
Sodium carbonate solution	Green	8	weak alkali
Lemon juice	Yellow	5	weak acid
Sodium chloride solution	Yellow/green	7	neutral

Action of Metals on Water

All metals are covered in a layer of oxide due to slow reaction of the metal with the Oxygen in the air :

Metal + Oxygen -> Metal oxide e.g.

Calcium + Oxygen -> Calcium oxide
Ca O_2 CaO

Sodium (Na) must be stored under oil to prevent it reacting with the Oxygen in the air.

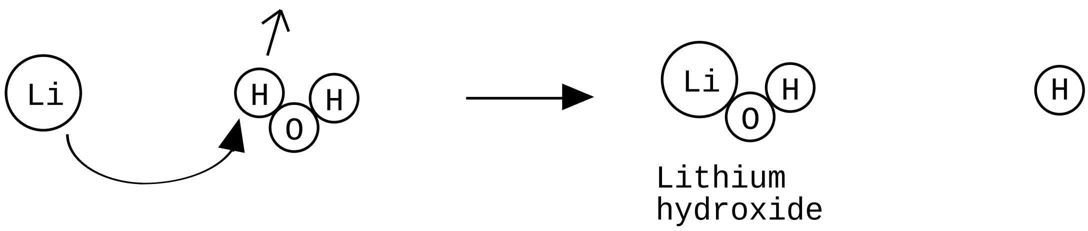
If we first scrape off the oxide to allow Water to get to the metal we can show that, on placing most metals in Water, Hydrogen gas is given off and a metal hydroxide is formed. All hydroxide solutions are alkalis so the solution turns pH paper green, blue or purple.

Metal + Water -> Metal hydroxide + Hydrogen

e.g.

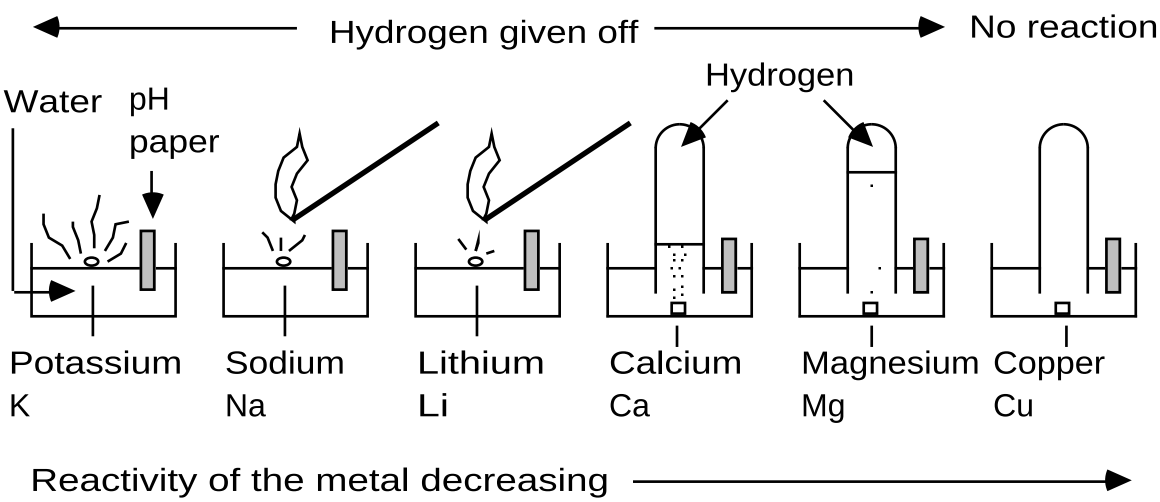
Lithium + Water -> Lithium hydroxide + Hydrogen

What happens in this reaction ? The metal atoms replace one of the Hydrogen atoms in the Water molecule H_2O and bond to the atoms left i.e. OH (the 'hydroxide group') :



The single Hydrogen atoms then join together forming a Hydrogen molecule H_2 :

Some metals are more reactive than others so the speed of reaction varies from metal to metal :



Colour of pH paper :

Purple	Purple	Blue	Blue green	Green	No change

Results :

Potassium : Explosive reaction. Hydrogen evolved burns with a lilac flame.

Sodium : Very fast reaction. Hydrogen burns (on lighting) with a yellow flame.

Lithium : Fast reaction. Hydrogen burns (on lighting) with a red flame.

N.B. Above three metals are all less dense than water and so float on the surface.

Calcium : Fairly fast reaction (test tube is full in about 30 seconds). Hydrogen burns with a 'pop' on lighting.

Magnesium : Very slow reaction (test tube contains only 1 cm^3 of Hydrogen after 2 weeks).

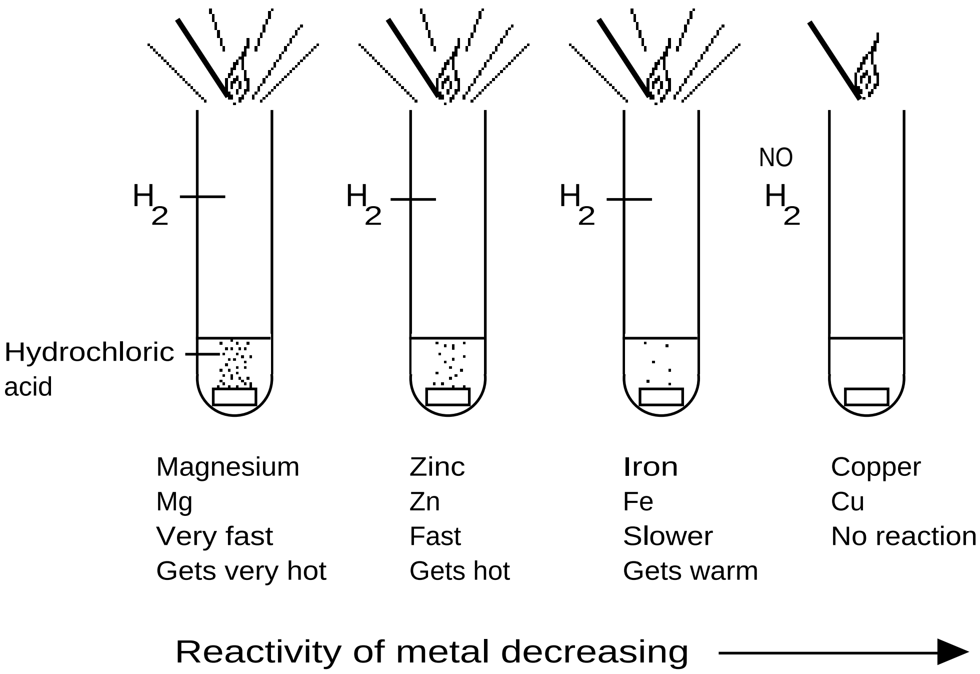
Copper : No reaction. Copper is so unreactive it does not react with Water.

Action of Metals on Acids

Like Water, all acids contain Hydrogen which can be replaced by metals e.g.

Hydrochloric acid	HCl
Sulphuric acid	H_2SO_4

The Hydrogen in acids, however, is not so firmly bonded and so is displaced much more easily by metals :



Again, the rate of reaction varies from metal to metal.
The metals dissolve forming **<u>SALTS</u>** : compounds formed by replacing the Hydrogen of an acid by a metal. Hydrochloric acid forms salts called 'chlorides' e.g.

Zinc	+	Hydrochloric acid	->	Zinc chloride	+	Hydrogen
Zn	+	HCl		$ZnCl_2$ (A salt)	+	H_2

Sulphuric acid forms salts called 'sulphates' e.g.

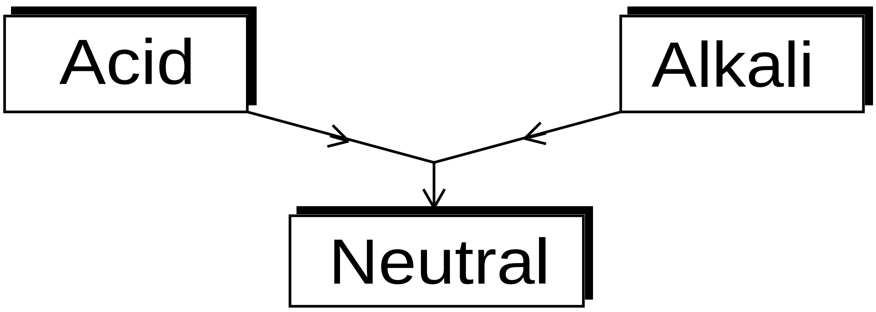
Zinc	+	Sulphuric acid	->	Zinc sulphate	+	Hydrogen
Zn	+	H_2SO_4	->	$ZnSO_4$ (A salt)	+	H_2

The salt can be recovered by evaporating slowly forming crystals.

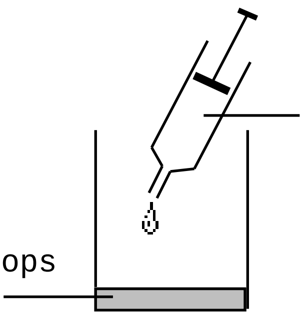
Neutralisation

When an acid is added to an alkali the acid 'cancels out' the alkali to give a neutral solution (pH=7).

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
Colour of pH paper	Red		Orange			Yellow		Yellow Green	Green		Green blue		Blue	Purple	



<u>Experiment 1</u> : Addition of Hydrochloric acid to Sodium hydroxide



Hydrochloric acid (drops at a time)

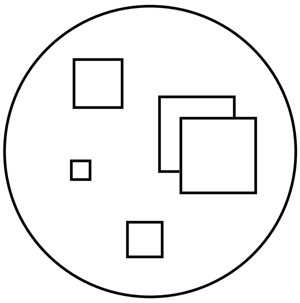
Volume required = ______ cm^3

10 cm^3 Sodium hydroxide + 5 drops pH indicator

As the acid is added to the alkali the pH drops. The indicator changes colour : purple -> blue/green -> green -> etc. We stop adding acid when the colour is yellow/green i.e. when the solution is neutral at pH=7.(We must stop at exactly 7 or the solution becomes acidic !) The solution obtained tastes salty : 'common salt' (Sodium chloride) has been formed.

Hydrochloric acid + Sodium hydroxide -> Sodium chloride + Water
$HCl + NaOH \rightarrow NaCl + H_2O$
The Sodium chloride is, however, contaminated with pH indicator !

The **same** volume (____ cm^3) of the **same** Hydrochloric acid should now be added to 10 cm^3 of the **same** Sodium hydroxide solution without the indicator. The solution is then partially evaporated and set aside to crystallise :

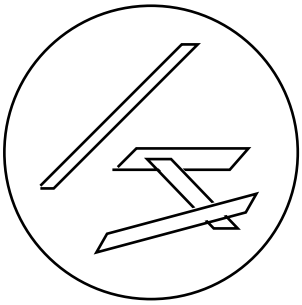


Crystals of Sodium chloride

Experiment 2 : Addition of Nitric acid to Potassium hydroxide.

The method is the same as in experiment 1.

Nitric acid + Potassium hydroxide -> Potassium nitrate + Water
$HNO_3 + KOH \rightarrow KNO_3 + H_2O$



Crystals of Potassium nitrate

What happens during neutralisation ?
It does not matter whether the acid is added to the alkali or the alkali to the acid, the H from the acid combines with the OH form the alkali forming H_2O (pH=7). The atoms remaining then combine to give a salt e.g. NaCl or KNO_3 (pH=7).

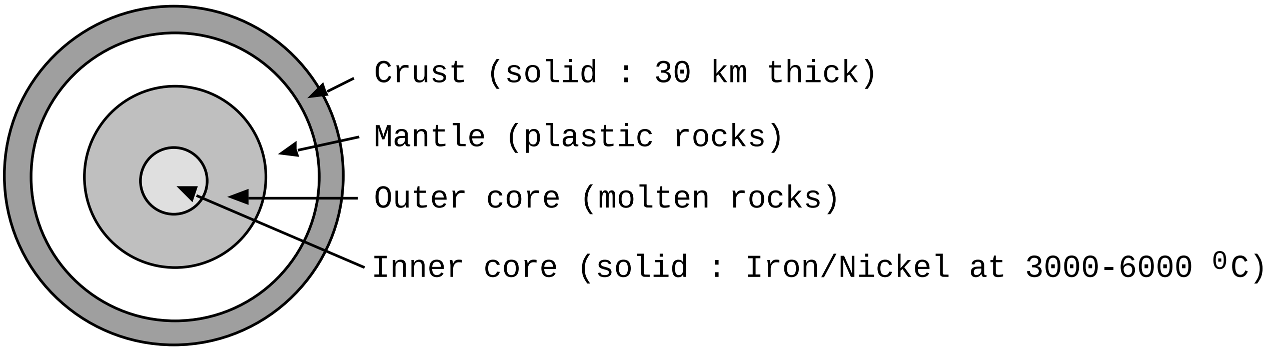
Applications of Neutralisation

1. Baking soda (alkali) neutralises acid in a bee sting.
2. Vinegar (acid) neutralises alkali in a wasp sting.
3. Milk of magnesia (alkali) neutralises excess Hydrochloric acid in the stomach.
4. Lime (alkali) neutralises acid soils.
5. Sodium carbonate (alkali) neutralises acid spills.

UNIT 2

The Earth

The Earth began as a molten mass about 4,500 million years ago. As it cooled down, rocks crystallised round the outside :



The rocks obtained when the molten rock cooled are called **igneous rocks** e.g. Granite. Subsequent erosion of the igneous rocks produced sediments which were pressed together, often trapping animal and vegetable remains (as fossils), to form **sedimentary rocks** e.g. sandstone. The action of heat and/or pressure on existing rocks can change their structure and form **metamorphic rocks** e.g. marble.

Naturally Occurring Elements

Most elements occur in nature combined with other elements in compounds. Only a few exist by themselves namely :

The 'Inert Gases'
Oxygen (in air)
Nitrogen (in air)
Carbon (in coal)
Sulphur (near volcanoes)

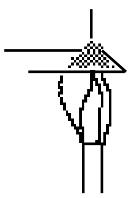
Copper (very little)
Mercury (very little)
Silver
Gold
Platinum

These five metals are <u>unreactive</u>. This is the reason they have not combined with other elements.

Reactive metals, like Magnesium and Zinc, do not occur free in nature. They have combined with other elements.
e.g.(1) with Sulphur to form sulphides :

Metal + Sulphur -> Metal sulphide

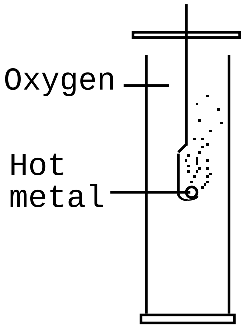
Mixture of powdered metal and Sulphur



Copper + Sulphur $\xrightarrow{\text{slow}}$ Copper sulphide
Cu + S → CuS

Zinc + Sulphur $\xrightarrow{\text{fast}}$ Zinc sulphide
Zn + S → ZnS

e.g.(2) with Oxygen to form oxides :



Metal + Oxygen $\longrightarrow$ Metal oxide

Copper + Oxygen $\xrightarrow{\text{slow}}$ Copper oxide
Cu O_2 CuO

Zinc + Oxygen $\xrightarrow{\text{fast}}$ Zinc oxide
Zn O_2 ZnO

There are many naturally occurring sulphides and oxides of the reactive metals e.g. Bauxite (Aluminium oxide). Using other metals as well, the following order of reactivity can be established in both these experiments and in others :

Most reactive	Potassium	K
	Sodium	Na
	Lithium	Li
	Calcium	Ca
	Magnesium	Mg
	Aluminium	Al
	Zinc	Zn
	Iron	Fe
	Tin	Sn
	Lead	Pb
	Hydrogen	H
	Copper	Cu
	Mercury	Hg
	Silver	Ag
	Gold	Au
Least reactive	Platinum	Pt

Another group of naturally occurring compounds is the **carbonates**, formed when the Hydrogen of Carbonic acid H_2CO_3 (formed by reaction of Carbon dioxide with Water) is replaced by a metal. If we add an acid to a carbonate, Carbon dioxide is given off.

Obtaining Metals from their Ores

This involves pulling the metal away from whatever it is combined with. This should be easy for unreactive metals like Copper which combines weakly with other elements and thus forms unstable compounds. It should be difficult for reactive metals like Calcium which combine strongly with other elements forming stable compounds.

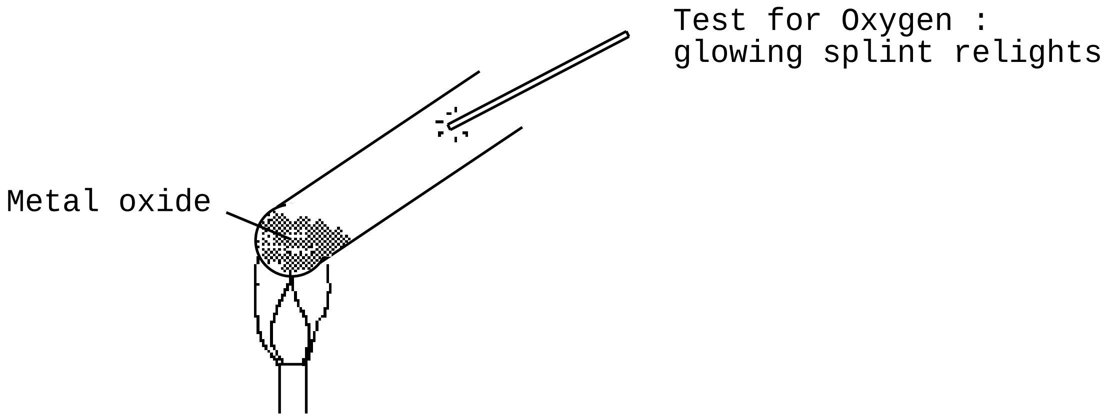
Action of Heat Alone : To try to break the bonds between the metal and the rest of the atoms in the compound.

1. Oxides

Only oxides of Mercury, Silver, Gold and Platinum decompose to give the metal and Oxygen e.g.

Mercury oxide $\xrightarrow{\text{heat}}$ Mercury + Oxygen

All other oxides do not decompose.

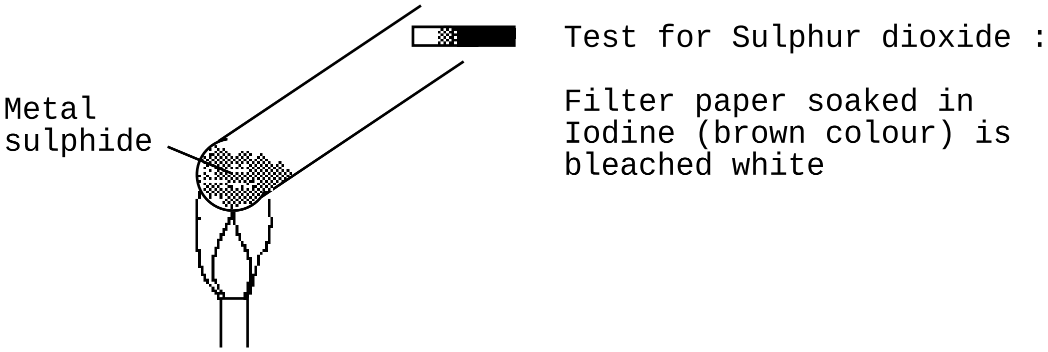


2. Sulphides

ALL sulphides react with the Oxygen in the air when heated to give oxides and the choking gas, Sulphur dioxide SO_2 e.g.

Zinc sulphide + Oxygen $\xrightarrow{heat}$ Zinc oxide + Sulphur dioxide

ZnS O_2 ZnO SO_2



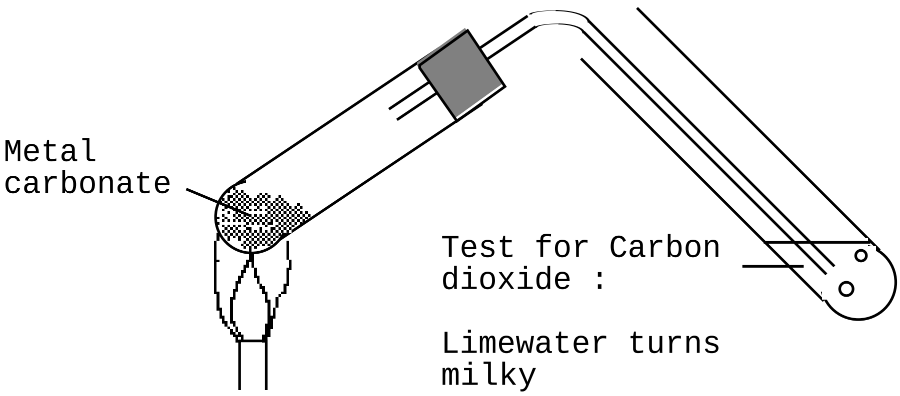
On further heating, oxides of Mercury, Silver, Gold or Platinum would break down to give the metal.

3. Carbonates

MOST carbonates decompose to give the oxide and the gas, Carbon dioxide CO_2

e.g. Lead carbonate $\xrightarrow{heat}$ Lead oxide + Carbon dioxide

$PbCO_3$ PbO CO_2



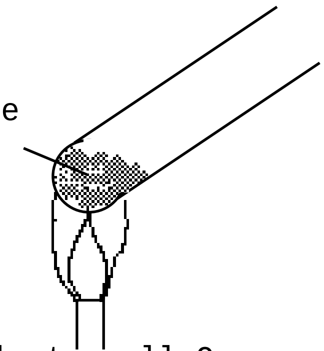
On further heating, oxides of Mercury, Silver, Gold or Platinum would break down to give the metal.

Heating with Carbon

A method has to be found to pull the Oxygen away from the metal in oxides of metals above Mercury in reactivity. If oxides of Copper, Lead, Tin and Iron are heated with Carbon, the Carbon pulls the Oxygen away from the metal and forms Carbon dioxide e.g.

Lead oxide	+	Carbon	$\xrightarrow{\text{heat}}$	Lead	+	Carbon dioxide
PbO		C		Pb		CO_2

Metal oxide
+ Carbon



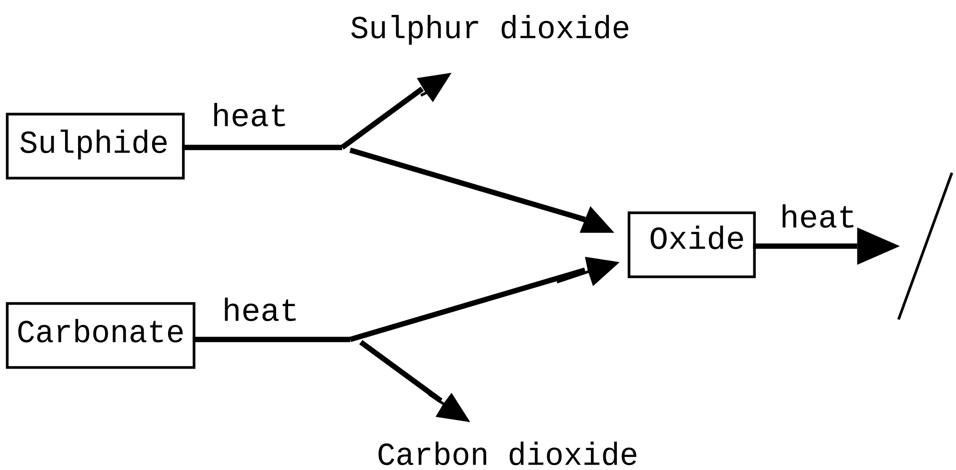
Carbon is unable to pull Oxygen off metals above Iron in reactivity - these metals hold on to the Oxygen atoms too strongly.
Since oxides are formed when sulphides and carbonates are heated, heating sulphides and carbonates of Iron, Tin, Lead and Copper with Carbon should also give the metal e.g.

Copper sulphide	+	Oxygen	->	Copper oxide	+	Sulphur dioxide
CuS		O_2		CuO		SO_2
Copper oxide	+	Carbon	->	Copper	+	Carbon dioxide
CuO		C		Cu		CO_2

Investigation of Malachite

Malachite is a green powder.

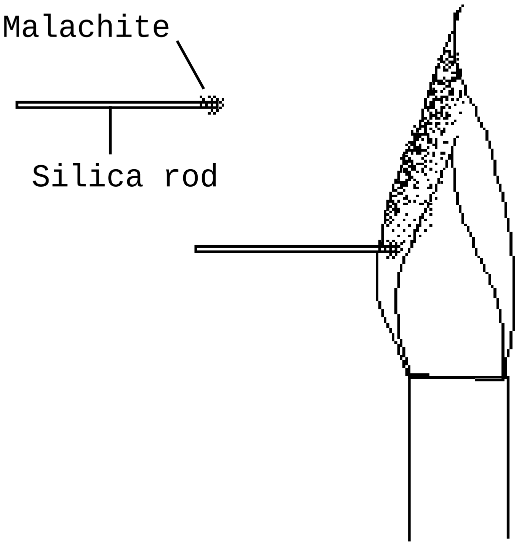
1. To find whether it is an oxide, sulphide or carbonate ... heat it.



Result : Carbon dioxide is given off (test with Limewater) ; a black powder remains - an oxide.

Conclusion : Malachite is a <u>carbonate</u>

2. Do a flame test to find out which metal is present :



Metal	Colour of flame
Sodium	Yellow
Calcium	Brick-red
Copper	Green
Potassium	Lilac

Result : green colour

Conclusion : Malachite contains <u>Copper</u>
Overall conclusion : Malachite is <u>Copper carbonate</u>

N.B. The 'black powder' formed in Experiment 1 above must be Copper oxide.

$$\text{Copper carbonate} \xrightarrow{\text{heat}} \text{Copper oxide} + \text{Carbon dioxide}$$
$$CuCO_3 \qquad CuO \qquad CO_2$$

3. Confirm that Malachite contains Copper by heating Malachite with Carbon : Copper metal is formed.

<u>Silicates</u>

Though most rocks contain small amounts of sulphides, oxides and carbonates, the main substances in rocks are **Silicates** e.g. granite is composed of three silicates : Mica (black specs), Feldspar (white or pink opaque crystals) and Quartz (white sparkling crystals).
Silicates have very strong Silicon - Oxygen bonds and are thus very stable. They are unreactive. They do not react with acid : glass (a mixture of Sodium silicate and Calcium silicate) can be used to store most acids. They do not decompose on heating : asbestos (a silicate containing Magnesium) is used in fire blankets etc and Mica is used as windows on paraffin stoves. They do not dissolve in water : buildings made from red sandstone (mostly quartz) do not dissolve when it rains !

UNIT 3

The Sea

There are 300 million cubic miles of sea water on the earth. Each cubic mile holds 171 million tons of dissolved solids.

Experiment

Object : To find out what substance is present in the dissolved solid.

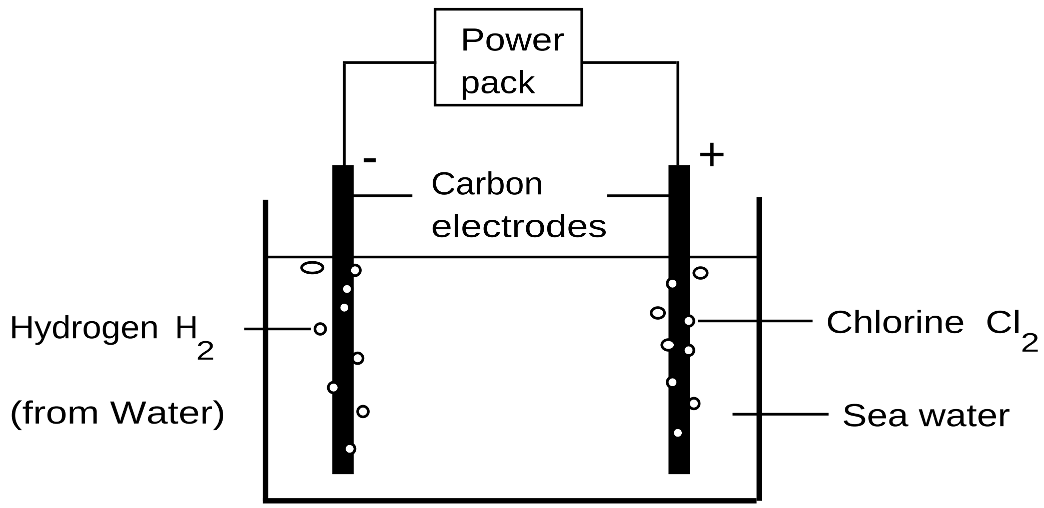
Method :

1. Evaporate 50 cm^3 sea water (a white solid was obtained)

   Weight of evaporating basin =
   Weight of evaporating basin + white solid =
   Weight of white solid =

2. Flame test the white solid (a yellow colour was obtained showing that **Sodium** is present)

3. Pass an electric current through some sea water (Chlorine is given off at the positive electrode showing that **Chlorine** is also present) :



Conclusion :

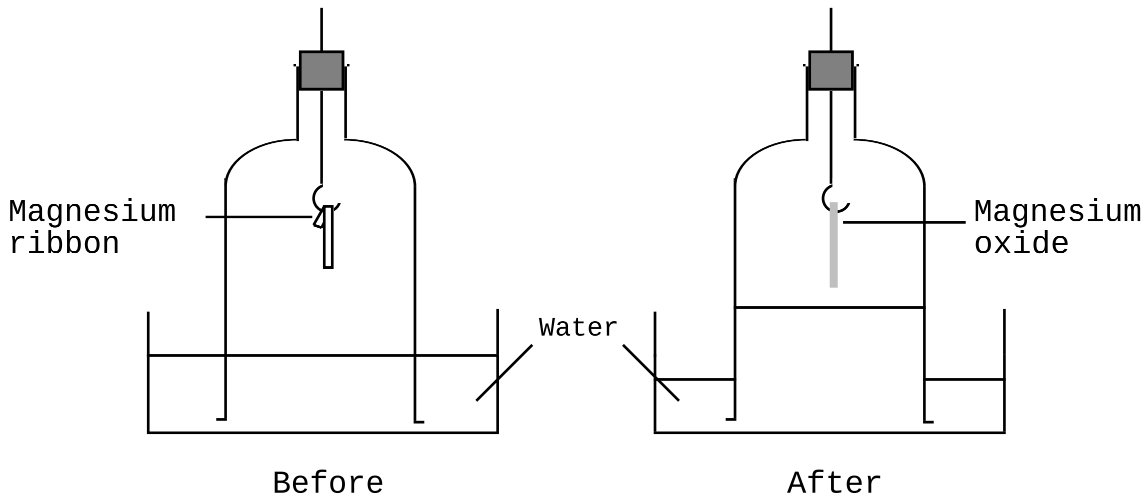
The main dissolved substance is **Sodium chloride** from which can be made : Sodium hydroxide (used in making soap etc) and Chlorine (used in making bleaches, plastics etc).

Evaporation of inland seas, millions of years ago, has lead to vast deposits of Sodium chloride or rock salt.

Other dissolved substances are also present. The dissolved substances are deposited in the sea by rivers which dissolve the soluble minerals in the rocks they flow over. They also come from the dissolving of minerals in rocks on the ocean floor.

The Gases in the Atmosphere

Experiment 1. Burn Magnesium in air as shown below :



The Magnesium burns forming Magnesium oxide :

Magnesium	+	Oxygen	->	Magnesium oxide
Mg		O_2		MgO

This uses up all the **Oxygen** in the jar. The water rises up into the jar to take the place of the Oxygen. The water fills about 1/5 of the jar so 1/5 of the air must be Oxygen. Plunge a burning splint into the gas remaining - it goes out. The gas remaining is mostly **Nitrogen**, an unreactive gas in which substances do not burn. About 4/5 of the air is Nitrogen N_2.

Experiment 2 : Pour some Limewater into a beaker and leave it to sit in the air for 30 minutes. The surface, in contact with the air, turns milky showing that air contains a trace of **Carbon dioxide** CO_2.

When fossil fuels like coal and oil are burned, Carbon dioxide gas is given off into the air. Carbon dioxide allows the sun's rays to penetrate the atmosphere and warm the surface of the Earth but traps the heat produced. This gradual warming of the Earth is known as the 'Greenhouse Effect' and will result in melting of the ice caps and changes in climate.

Experiment 3 : Leave some dry Cobalt chloride paper lying in the air for 30 minutes. The paper turns from blue to pink because the air contains a trace of **Water** vapour.

Composition of the Air :

Nitrogen	78 %
Oxygen	21 %
Argon	1 %
Carbon dioxide	trace
Water vapour	trace

The Moon

The first men landed on the moon in the Sea of Tranquillity near the crater Moltke on July 20 1969 at 2118 BST (Apollo 11).
During its formation the Moon's low gravity allowed all gases coming through the crust from the interior to escape into space with the result that the Moon has no atmosphere. The astronauts Oxygen requirement, 630000 cm^3 per day, was carried in cylinders. The Carbon dioxide breathed out was absorbed in Lithium hydroxide. The lunar rocks are composed mainly of Aluminium and Magnesium silicates. The most common minerals are Plagioclase (a feldspar), Pyroxenes and Silica (occurring as glassy globules).

Many features are not present due to the lack of an atmosphere. The minerals contain no hydroxides (showing the lack of Water vapour). The rocks contain no carbonates (because of the lack of Carbon dioxide) and no fossils (lack of living things).

Copper and its Compounds

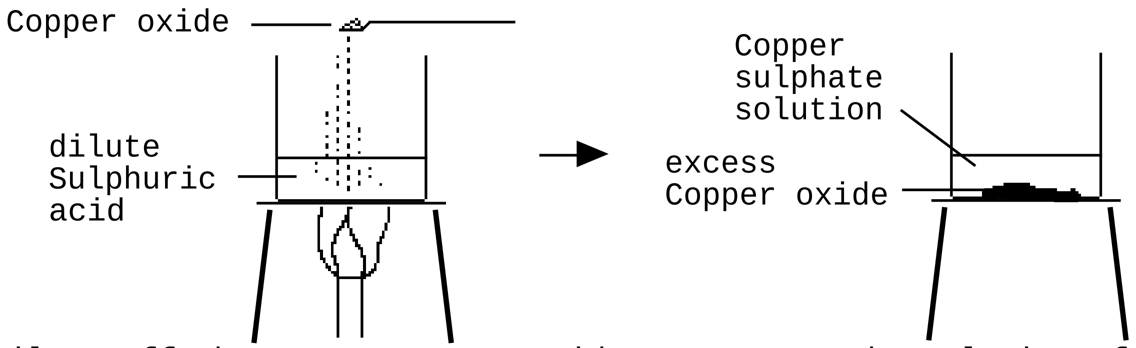
Copper Cu is a soft reddish-brown metal, melting point 1083 ^{0}C. It is a good electrical conductor and so is used for wiring. It is very unreactive. It does not react with Water and is therefore used for water pipes. Most acids do not react with Copper ; it does however react with strong Nitric acid.

Experiment 1

Object : To prepare Copper sulphate from Copper oxide.

Method : Add Copper oxide (a black powder), a little at a time, to warm, dilute Sulphuric acid in a beaker until no more will react. A blue solution of Copper sulphate is formed :

Copper oxide	+	Sulphuric acid	->	Copper sulphate	+	Water
CuO		H_2SO_4		$CuSO_4$		H_2O



Filter off the excess Copper oxide. Evaporate the solution of Copper sulphate part way then set aside to crystallise.

Result : Blue, diamond-shaped crystals of Copper sulphate were formed.

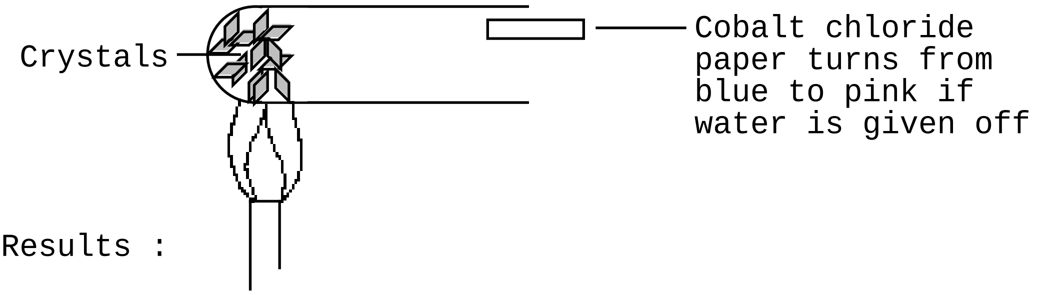
Water of Crystallisation

In the previous experiment, as the Water evaporates and the crystals begin to grow, Water is trapped inside the crystals : this is called 'Water of Crystallisation'. Some substances contain water of crystallisation ; some do not.

Experiment 2 :

Object : To find out if substances contain water of crystallisation.

Method : Heat the crystals in a test tube. The Water, if present, is driven out and sometimes the crystals change colour.



Results :

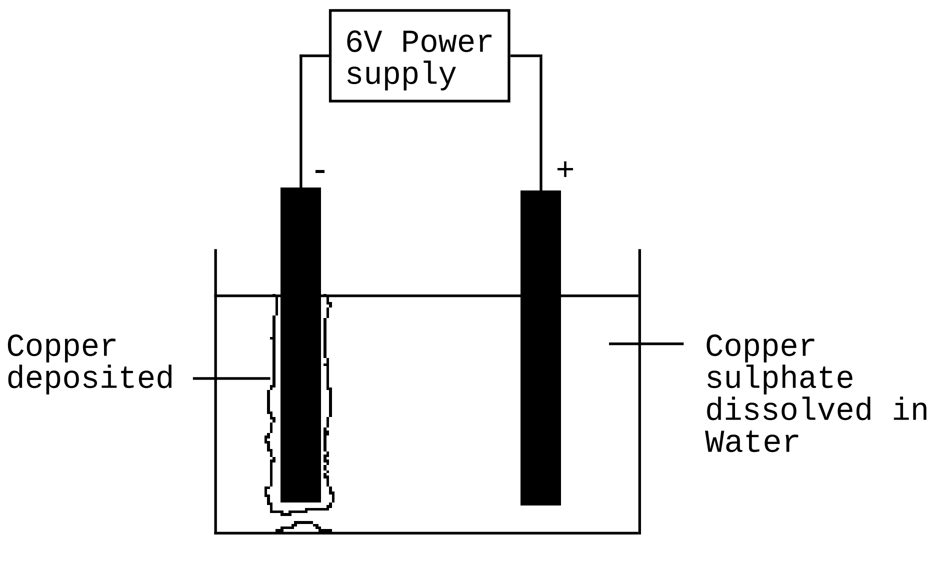
Substance	Colour	Colour after heating	Water driven off	Any 'Water of Crystallisation'
Copper sulphate	Blue	Colourless	Yes	Yes
Magnesium sulphate	Colourless	Colourless	Yes	Yes
Potassium nitrate	Colourless	Colourless	No	No
Cobalt chloride	Pink	Blue	Yes	Yes

Experiment 3 : Since Copper is so unreactive it is easy to obtain it from its compounds.

Object : To obtain Copper from Copper sulphate by two methods

Method 1 : by electrolysis

On passing an electric current through a solution of Copper sulphate, Copper is formed at the negative electrode :

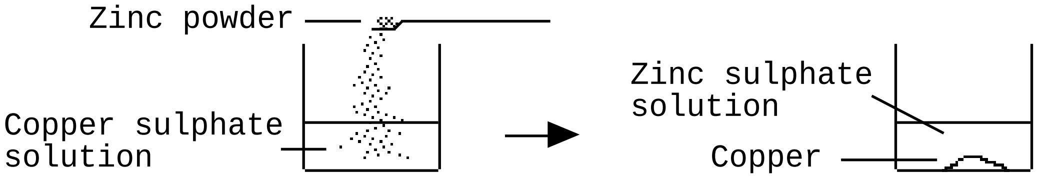


Method 2 : by reaction with Zinc

Since Zinc is more reactive than Copper, Zinc will displace Copper from Copper sulphate and, itself, become bonded to the sulphate group :

Zinc	+	Copper sulphate	->	Zinc sulphate	+	Copper
Zn		$CuSO_4$		$ZnSO_4$		Cu

Add Zinc powder to a solution of Copper sulphate until the blue colour becomes colourless. Filter off the Copper produced.



Results : Brown Copper powder was produced by both methods.