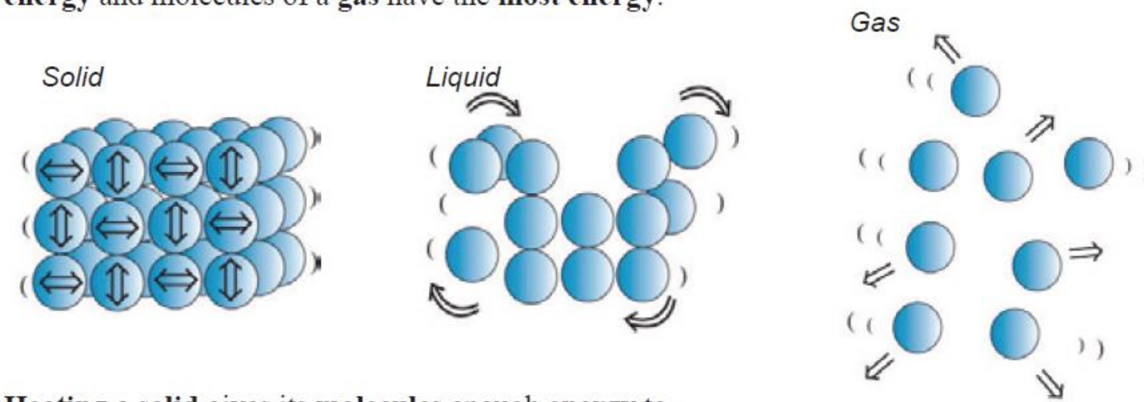
JC Chemistry

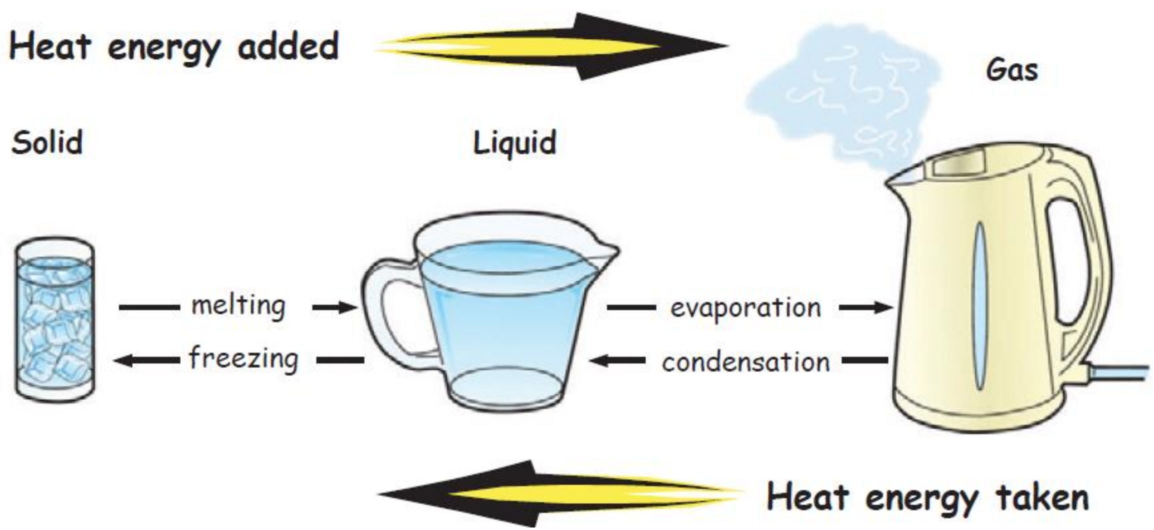
States of Matter

- **Matter** is anything that takes up **space** and has **mass**.
- **Solids** have a **definite mass**, **definite shape** and **definite volume**.
  They do not flow and can't be squeezed into a smaller space.
- **Liquids** have a **definite mass**, **definite volume** but **no definite shape**.
  They can **flow** but can't be squeezed into a smaller space.
- **Gases** have a **definite mass**, but **no definite volume** or **shape**.
  They move into the available space and can be squeezed into a smaller space.

- **Molecules** of a **solid** have the **least energy**, molecules of a **liquid** have **more energy** and molecules of a **gas** have the **most energy**.

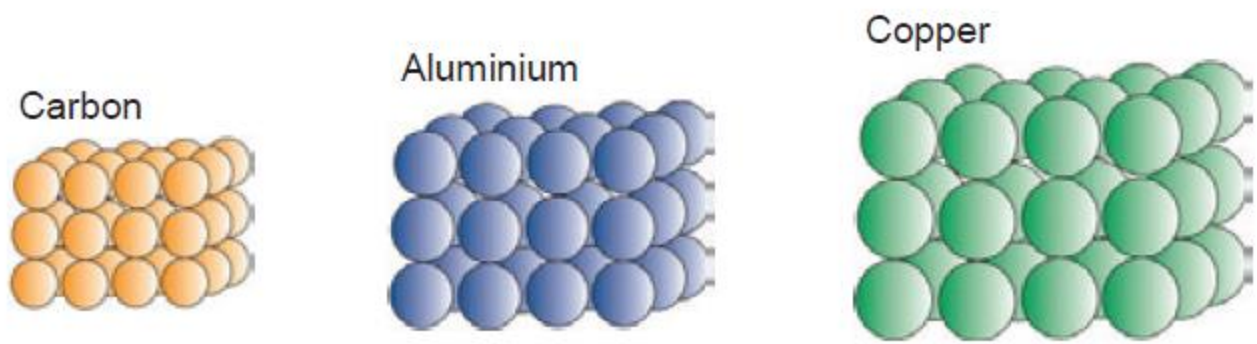


- **Heating a solid** gives its **molecules** enough **energy** to behave like a liquid (it **melts**).
- **Heating a liquid** gives its **molecules** enough **energy** to behave like a gas (it **evaporates**).
- The **melting point** of a solid is the temperature at which both the solid and liquid states of a substance exist together.
- **Evaporation** is the changing of a liquid to a gas or vapour.
- The **boiling point** of a liquid is the temperature at which evaporation begins to occur throughout the liquid.
- **Condensation** is the changing of a gas to a liquid.



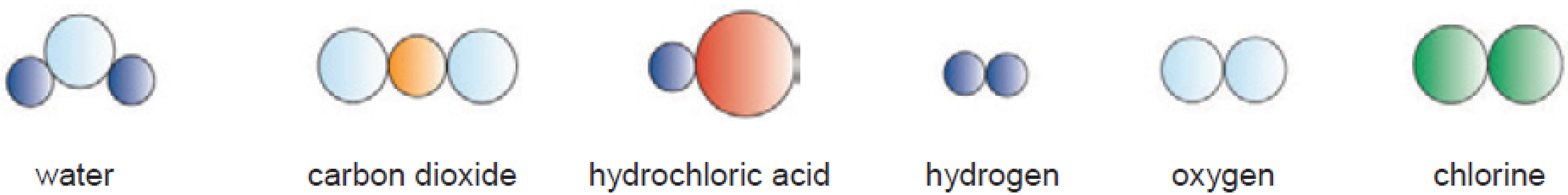
Elements, Compounds, Mixtures

- An **element** is a substance which cannot be broken down into simpler substances by chemical means.
- **Examples of elements** that are **metals** are: iron (Fe), copper (Cu), zinc (Zn), aluminium (Al), silver (Ag), gold (Au), and sodium (Na).



- **Examples of elements** that are **non-metals** are: carbon (C), sulfur (S), chlorine (Cl), oxygen (O), hydrogen (H), and nitrogen (N)
- An **atom** is the smallest part of an element that still has the properties of that element.

- A **compound** is formed when two or more elements combine chemically.
- **Examples of compounds** are: water (H_2O), carbon dioxide (CO_2), sodium chloride ($NaCl$), magnesium oxide (MgO), and iron sulfide (FeS).
- A **compound** is a completely new substance with its own properties.
- A **molecule** is the smallest part of an element or compound that can exist on its own.
- **Examples of molecules** are: H_2O, CO_2, HCl, MgO, FeS, O_2, H_2, and Cl_2.

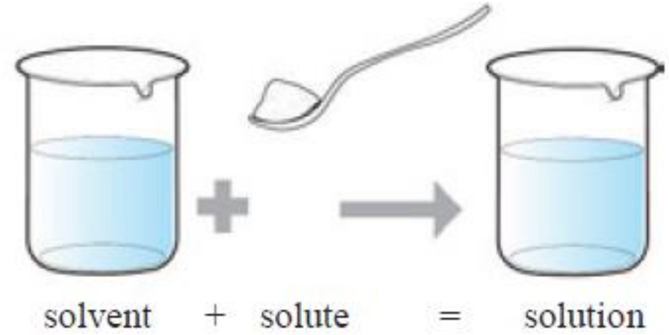


- A **mixture** consists of two or more different substances **mingled** together but **not chemically combined**.
- **Examples of mixtures** are: air (nitrogen + oxygen + carbon dioxide), seawater (water + salt), **ink** (water + various dyes), steel (iron + carbon).

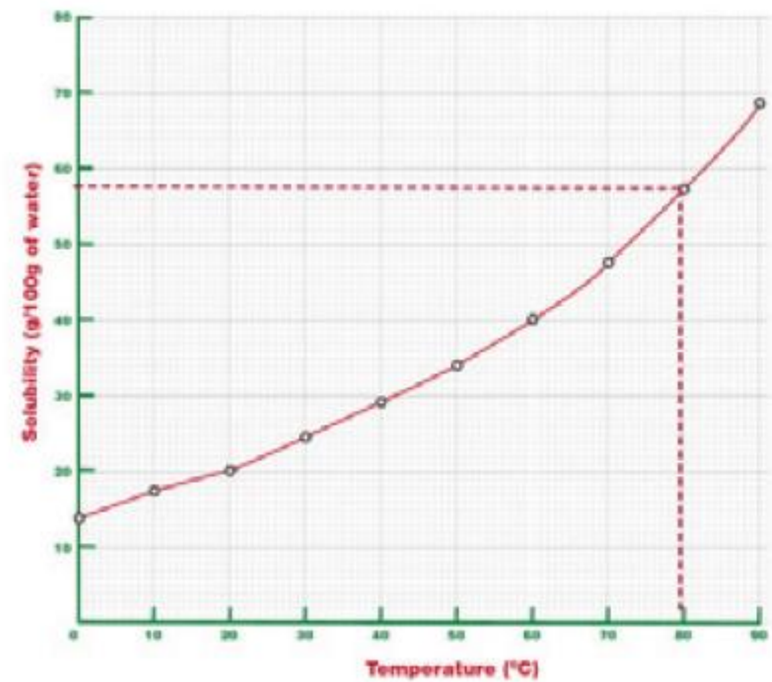
- **Differences** between mixtures and compounds:

	MIXTURE	*COMPOUND*
1.	Consists of two or more substances.	Consists of a single substance
2.	The proportion of each substance in the mixture does not matter	The elements in a compound are always in a fixed proportion.
3.	The properties are the same as those of the substances used.	The properties are very different to those of the elements used.
4.	Usually easy to separate.	Very difficult to separate.

Solutions



- A **solution** is a **mixture** of a **solute** and a **solvent**.
- A **solute** is the substance which is **dissolved**.
- A **solvent** is the **liquid** in which the **solute dissolves**.
- When blue **copper sulfate** (**solute**) is dissolved in **water** (**solvent**) a **solution** of copper sulfate in water is formed.
- The **hotter** a solution of copper sulfate is, the **more** solute it will dissolve.
- A **saturated solution** is one which contains as much dissolved solute as possible, at a given temperature.
- A **solubility curve** shows how the solubility of a substance changes with increasing temperature.
- A **concentrated solution** has a large amount of solute in a small amount of solvent.



solubility curve for copper sulfate

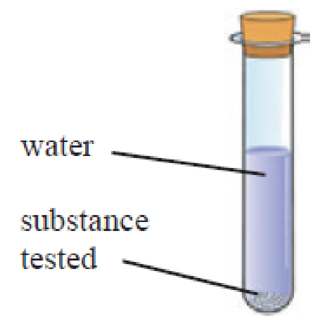
- A **dilute solution** can be made more concentrated by either **adding** more solute, or by **evaporating** off some of the solvent.
- **Crystallisation**: The forming of crystals when a hot saturated solution is cooled.
- When a **hot saturated solution** of copper sulfate is allowed to **cool**, blue **crystals** of copper sulfate are formed.

EXPERIMENTS:

24.1 To Investigate the Solubility of Different Substances in Water

A **spatula full** of each substance is added to water in a test tube.

The test tube is **stoppered** and **shaken** to see which dissolve.



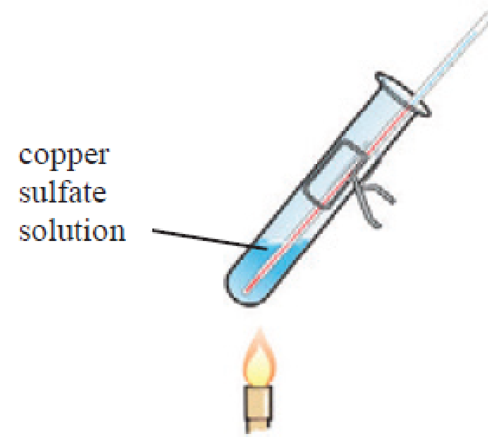
24.2 To Investigate the Solubility of Copper Sulfate at Different Temperatures

The mass of **copper sulfate** that will dissolve in **100 g** of **water** at **20°C** is found using an **electronic balance**.

The water is **heated** to **30°C** and **more copper sulfate** is added to find the mass that can dissolve at this temperature.

This is repeated for temperatures of **40°C** , **50°C** and **60°C**.

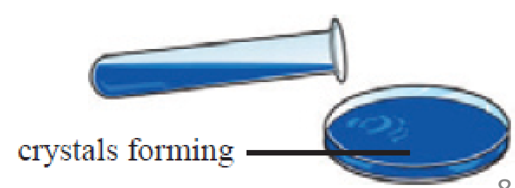
The results are **plotted** on graph paper to give a **solubility curve** for copper sulfate (see above).



24.3 Growing Copper Sulfate Crystals

A **hot**, **concentrated solution** of copper sulfate is poured onto a warm evaporating dish. The dish is left for 3 hours.

Crystals of **copper sulfate** form as the solution cools.

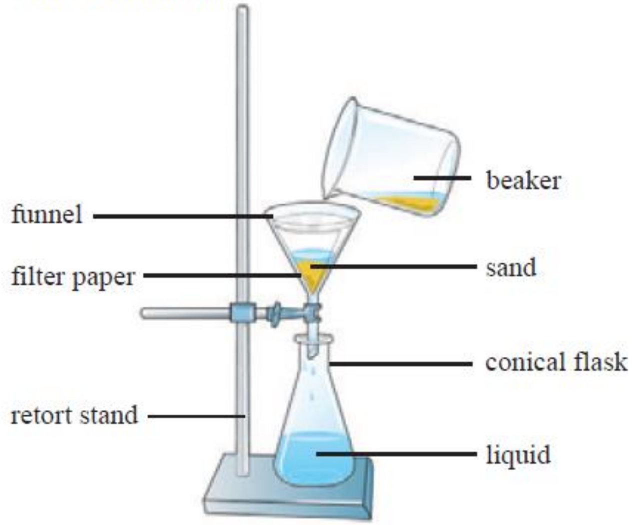


Separating Mixtures

- **Filtration** is a method used to separate small **insoluble solids** from a **liquid** by using **filter paper** to trap the solids.
- **Evaporation** is a method used to separate **soluble solids** from a **solution** by **evaporating** off the liquid to leave the solids.
- **Crystallisation** occurs when crystals appear from a solution which has been **evaporated**.
- **Crystallisation** occurs when crystals appear from a **hot**, **concentrated solution** which has **cooled**.
- **Distillation** is used to separate **two liquids** with different **boiling points** such as alcohol and water. It is also used to separate a **soluble solid** from a **liquid** (e.g. seawater) to give a pure sample of each.
- **Chromatography** is a method used to separate a mixture of dissolved substances in a solution.

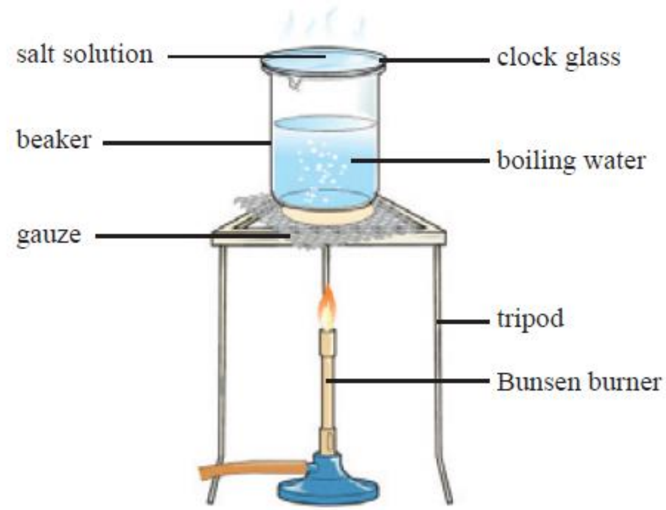
EXPERIMENTS:

24.1 Filtration



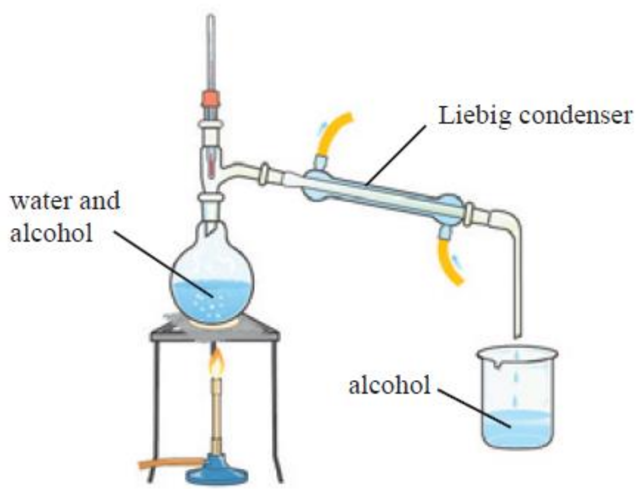
- The sand is trapped in the filter paper, the water goes through.

24.2 Evaporation



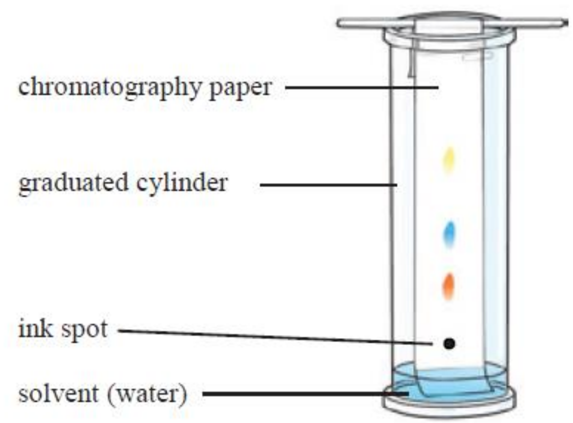
- The water evaporates, the salt remains on the clock glass and forms crystals.

24.3 Distillation



- The alcohol (boiling point 78°C) evaporates first, condenses in the Liebig condenser and is collected.

24.4 Chromatography



- The more soluble dyes stay in solution longer and get deposited further up the chromatography paper.

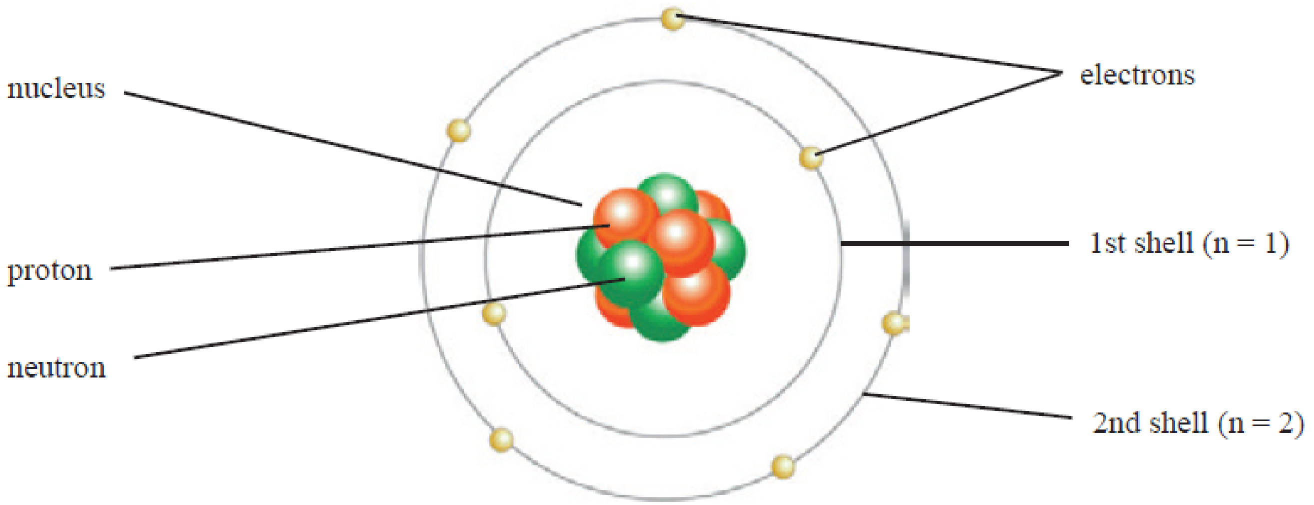
The Atom - A Closer Look

- The **atom** is made up of sub-atomic particles called **protons**, **neutrons** and **electrons**.

	MASS	CHARGE	LOCATION
PROTON	1 amu	+1	In the nucleus
NEUTRON	1 amu	0	In the nucleus
ELECTRON	1/1840 amu	–1	Orbiting the nucleus

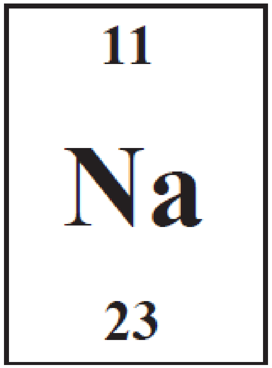
- The **atomic number** of an atom is the **number** of **protons** it has in its nucleus.
- Atoms are arranged in the **Periodic Table** according to the **number** of **protons** they have.
- Atoms are **electrically neutral** because they have the **same number** of **electrons** as **protons**.
- **Electrons** are found in **shells** around the nucleus.
  The **first shell** can only hold **2 electrons**, all the others can hold a maximum of **8**.
- Shells are filled from the **inside** (1st shell) **outwards**.

- The **atomic number** of **sodium** (Na) is **11**, **-** it has 11 protons and therefore 11 electrons. Its **electronic configuration** is **2**, **8**, **1** (i.e. three shells of electrons).
- The electronic configuration of **calcium** (atomic number = **20**) **is 2**, **8**, **8**, **1**.
- The **atom** shown below is an atom of **nitrogen** (N). How do you know?



The Bohr structure of the atom

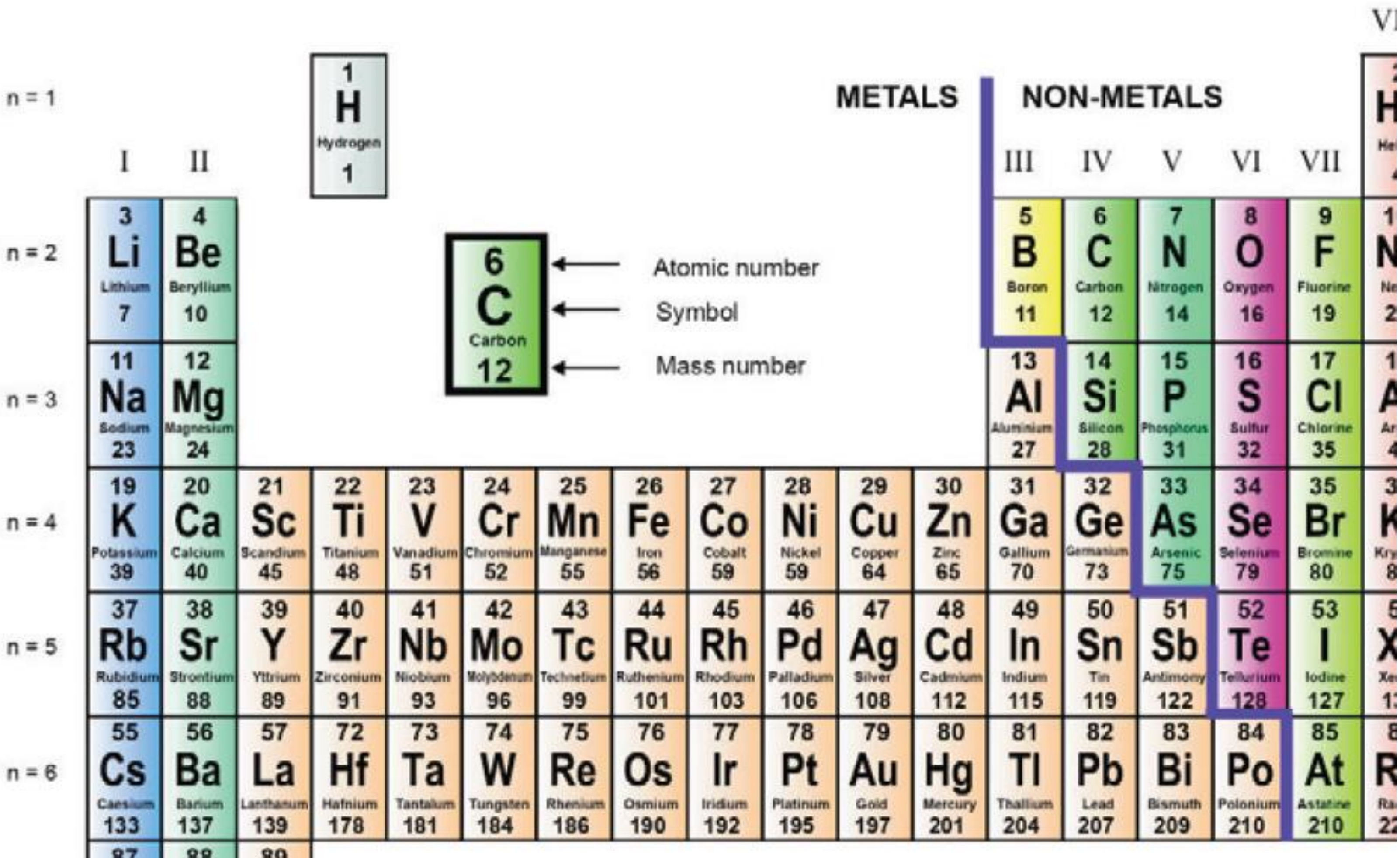
- The **mass number** of an atom is the number of **protons plus** the number of **neutrons** in its nucleus.
- **The mass number** is the **larger number**, written **below** the element symbol in the **Periodic Table**.
- **An atom of sodium** has **11 protons** (atomic number 11) and **11 electrons**, it has 23 protons plus neutrons (mass number 23), so it has **12 neutrons** in its nucleus (23 – 11 = 12).



- **Isotopes** are **atoms** of the **same element**, which have **different numbers of neutrons**.
- **Isotopes** of an element therefore have the **same atomic number,** but **different mass numbers.**

The Periodic Table

- The **atomic number** of an atom is the number of **protons** in the nucleus of that atom.
- The **Periodic Table** arranges the elements in order of **increasing atomic number**.



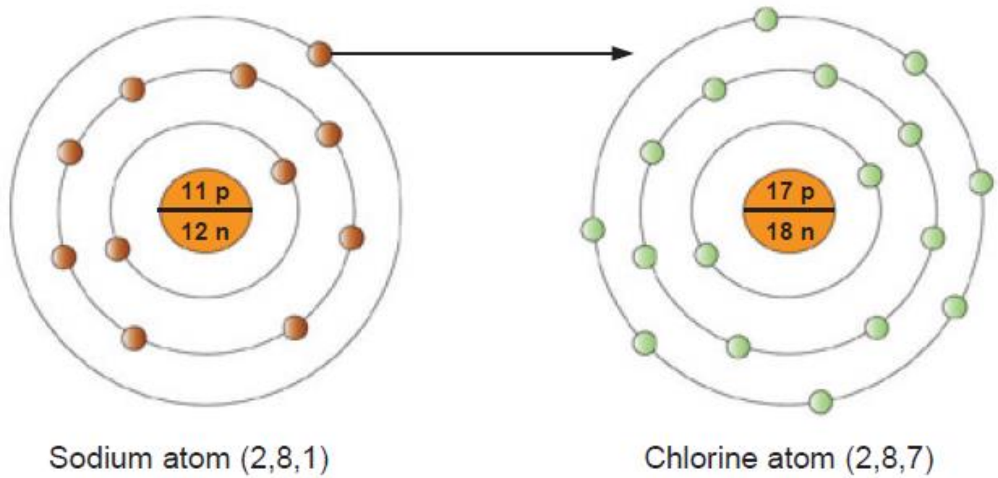
The Periodic Table of the Elements

- A **Group** is a **vertical column** of elements that all have the same number of electrons in their outside shells.
- **All the elements** in a particular Group behave in a similar manner chemically.
- Group I is the **Alkali Metals**, and includes the elements **lithium**, **sodium** and **potassium**.
- Group II is the **Alkaline Earth Metals**, and includes the elements **magnesium** and **calcium**.
- Group VII is the **Halogens**, and includes the elements **helium**, **neon** and **argon**.
- Group VIII is the **Noble Gases**.
- A **Period** is a **horizontal row** of elements in the Periodic Table.
- **Periods are numbered** n=1; n=2; n=3; etc.
- All the elements in a Period have the same number of electron shells.
- The Periodic Table can be divided into **metals** (on the left and middle), and **non-metals** (on the right).
- Common **metal elements** include: Cu, Zn, Al, Pb, Fe, Ag and Au.
- Common **non-metal elements** include: C, O, S, H and N.

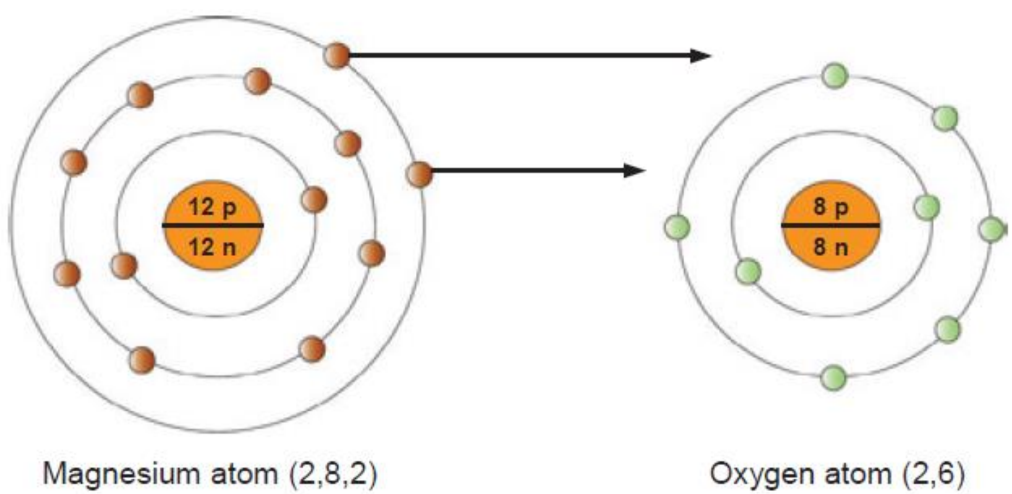
Chemical Bonding I - Ionic

- Atoms combine with each other to form **compounds**.
- A **chemical bond** holds atoms together in a compound.
- The **noble gases** (Group VIII elements) have the **most stable** electron setup. They have an outer shell of 8 electrons (i.e. they all have a **full outer shell**).
- The **Octet Rule** states that atoms bond together so that each atom ends up with an electron arrangement with 8 electrons in its outermost shell.
- An **ionic bond** is formed when electrons are **given** or **taken** by atoms.
- An **ion** is a positively or negatively charged atom or group of atoms.
- When an atom **loses electrons**, it becomes a **positively charged** ion.
- When an atom **gains electrons**, it becomes a negatively **charged** ion.

- A **sodium atom** (2, 8, 1) gives its outer electron to a **chlorine atom** (2, 8, 7) to form NaCl, **sodium chloride**, made up of **sodium ions** (Na^+) and **chloride ions** (Cl^-). An **ionic bond** is formed and both **ions** now have a **full outer shell** of 8 electrons.



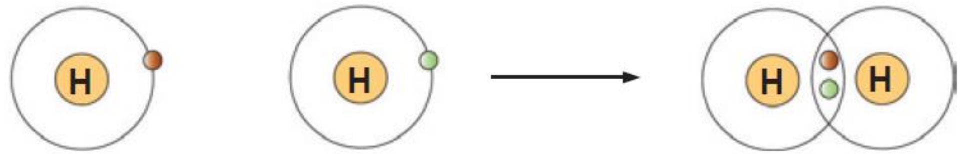
- In the same way, a **magnesium atom** forms an **ionic bond** with an **oxygen atom** by **giving** it its two outer electrons. The ionic compound, **magnesium oxide** (MgO) is formed.



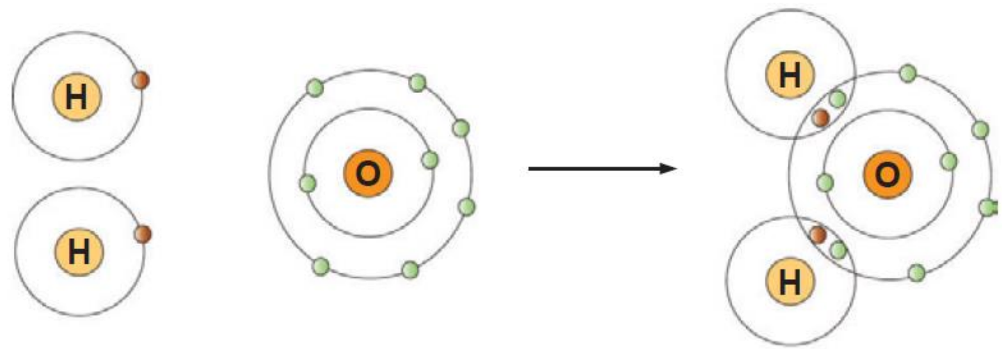
- An **ionic bond** is formed by the **force of attraction** between a **positive** and a **negative ion**.

Chemical Bonding II - Covalent

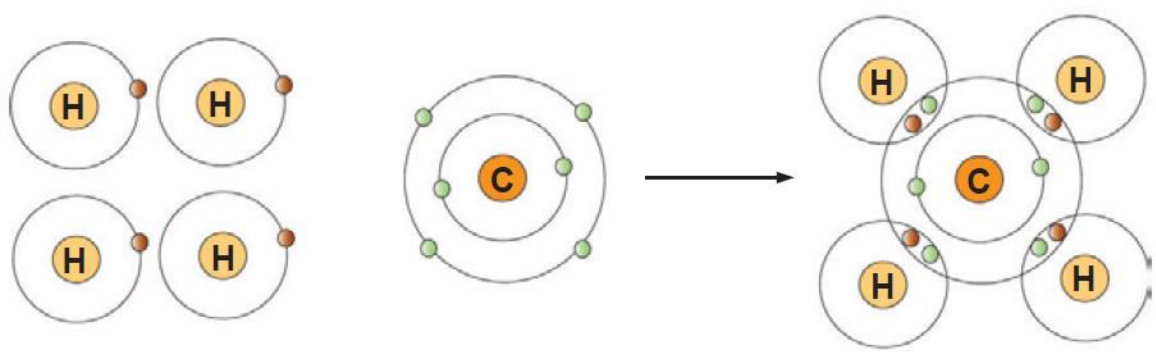
- A **covalent bond** is formed when atoms combine by **sharing electrons** so that each atom has a stable outer shell of electrons.
- A **single** covalent bond is formed when atoms share **one pair** of electrons.
- **Single covalent bonds** occur in the following molecules:



Hydrogen atom — Hydrogen atom — Hydrogen molecule (H_2)

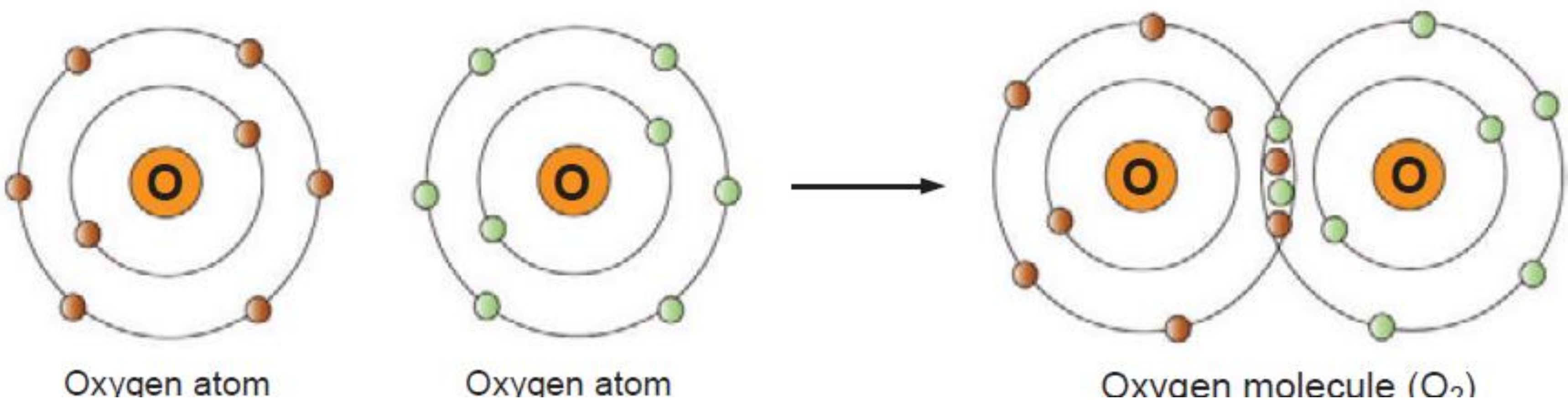


Hydrogen atoms — Oxygen atom — Water molecule (H_2O)



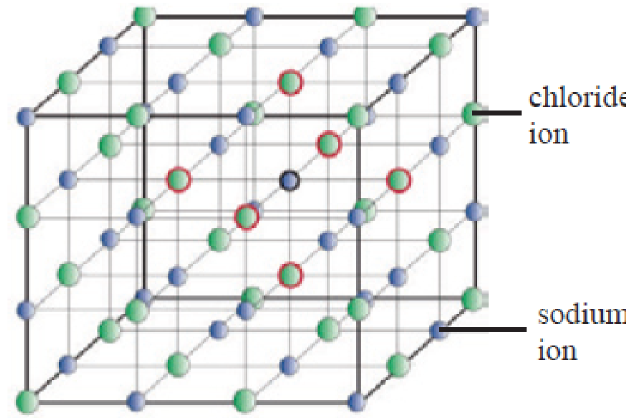
Hydrogen atoms — Carbon atom — Methane molecule (CH_4)

- A **double covalent bond** is formed when atoms share **2 pairs** of electrons.
- A **double bond** occurs in the **oxygen molecule**:



Ionic & Covalent Compounds

- **Ionic bonds** are formed between **metals** and **non-metals**.
- **Ionic compounds** consist of many oppositely charged ions held together strongly to form a **crystal lattice**.
- Examples of ionic compounds include: **sodium chloride** (NaCl), and **magnesium oxide** (MgO).
- A **covalent compound** is made up of separate, **single molecules**.
- In a covalent compound, there is a fairly **weak attraction *between*** the individual molecules.
- Examples of **covalent compounds** include: **water** (H_2O), **carbon dioxide** (CO_2), and **methane gas** (CH_4).



The sodium chloride crystal lattice

ionic compounds	covalent compounds
Consist of crystal lattices	**Consist of separate molecules**
Usually crystalline solids	**Usually liquids or gases**
High melting and boiling points	**Low melting and boiling points**
Usually soluble in water	**Usually insoluble in water**
Conduct electricity when melted or in solution (see experiment below)	**Do not conduct electricity (see experiment below)**

EXPERIMENT:

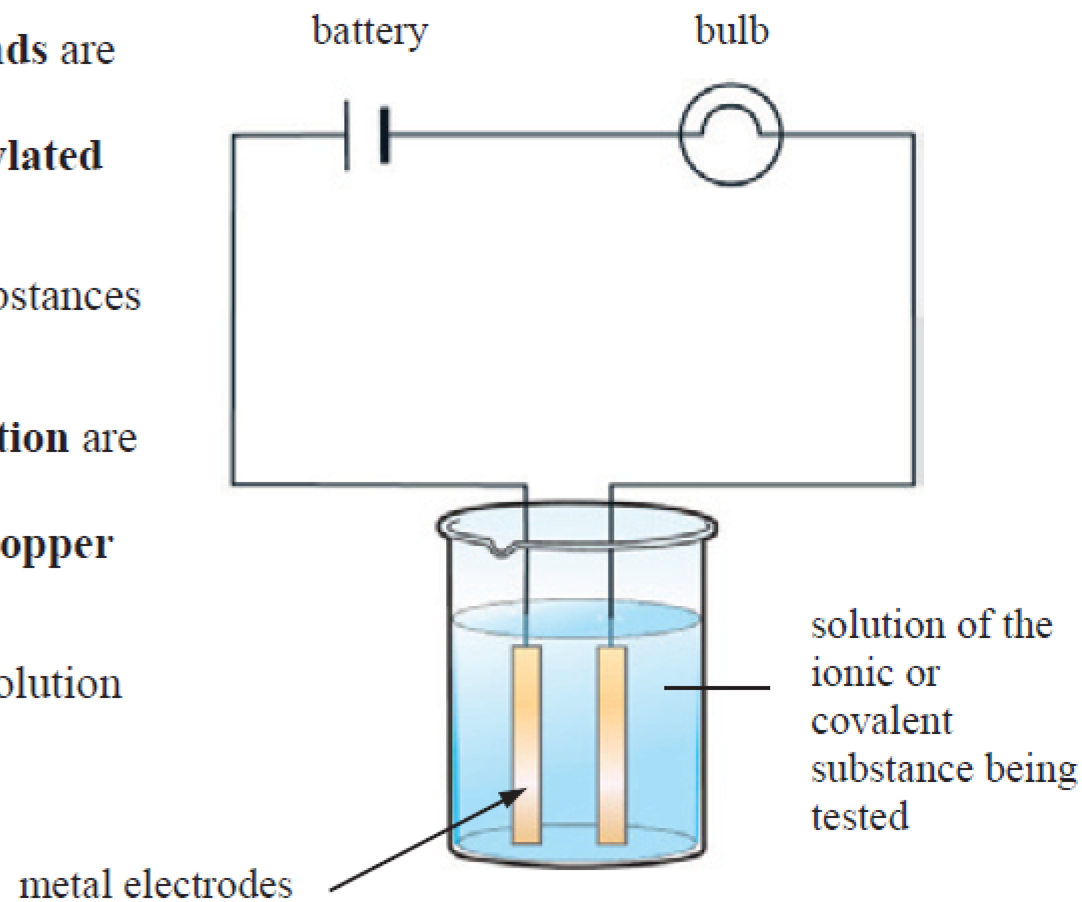
29.1 To Investigate the Ability of Ionic and Covalent Compounds to Conduct Electricity

A circuit is set up as shown.
A variety of **covalent liquid compounds** are poured into the beaker.
Examples include: **paraffin oil**, **methylated spirit**, **distilled water**.

The **bulb does not light**. **Covalent** substances will **not conduct electricity**.

A variety of **ionic compounds in solution** are poured into the beaker.
Examples include: **sodium chloride**, **copper sulfate**, **sodium hydroxide**.

The bulb lights. **Ionic** substances in solution will **conduct electricity**.



ACIDS & BASES

- An **acid** is a substance that **turns blue litmus** paper **red**.
- **Strong acids** are **hydrochloric acid** (HCl) and **sulfuric acid** H_2SO_4.
- **Weak acids** include **vinegar**, **lemon juice**, and **acid rain**.
- A **base** is a substance that **turns red litmus** paper **blue**.
- **Bases** that are **soluble in water** are called **alkalis**.
- **Strong bases** are **sodium hydroxide**, NaOH, and **calcium hydroxide**, $Ca(OH)_2$.
- **Weak bases** include **toothpaste**, **soap** and **window cleaner**.
- An **indicator** is a chemical which shows, by means of a colour change, whether a substance is an acid or a base. **Litmus** is an indicator.
- The **pH scale** goes from **0 to 14** and measures the strength of an acid or a base.
- **Universal indicator paper** is used to **measure the pH** of a substance.
- **Neutral solutions** have a pH of **7**.
  Acids have a pH of **less than 7**.
  Alkalis have a pH of **greater than 7**.

- A **neutralisation reaction** occurs when an **acid** and a **base** react together and neutralise each other to form a **salt** and **water**.
- An experiment to neutralise an acid (HCl) with a base (NaOH) is called a **titration**. A **burette**, **pipette**, **conical flask** and **white tile** are used.

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

- A **salt** (e.g. NaCl) is formed when the **hydrogen in the acid** is replaced by a **metal**.
- The **salt** can be seen by **evaporating off** the water in a **clock glass**.
- Acids are involved in the following reactions:

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$
$$HCl + NaOH \longrightarrow NaCl + H_2O$$

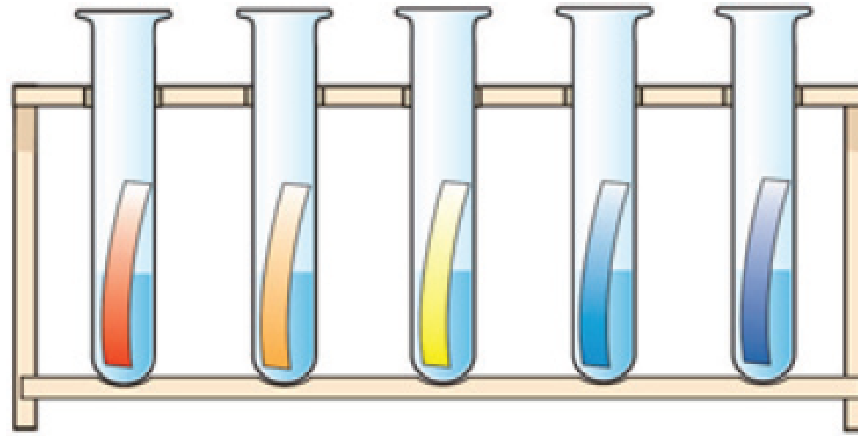
$$\text{Acid} + \text{Carbonate} \longrightarrow \text{Salt} + \text{Water} + \text{Carbon dioxide}$$
$$2HCl + CaCO_3 \longrightarrow CaCl_2 + H_2O + CO_2$$

EXPERIMENTS:

30.1 Testing the pH of Various Chemicals

Various **acids** and **alkalis** are placed in test tubes and tested with strips of **universal indicator paper**.

The **colour change** of the paper is then compared with a **pH colour chart** to find the **pH** of the solution.



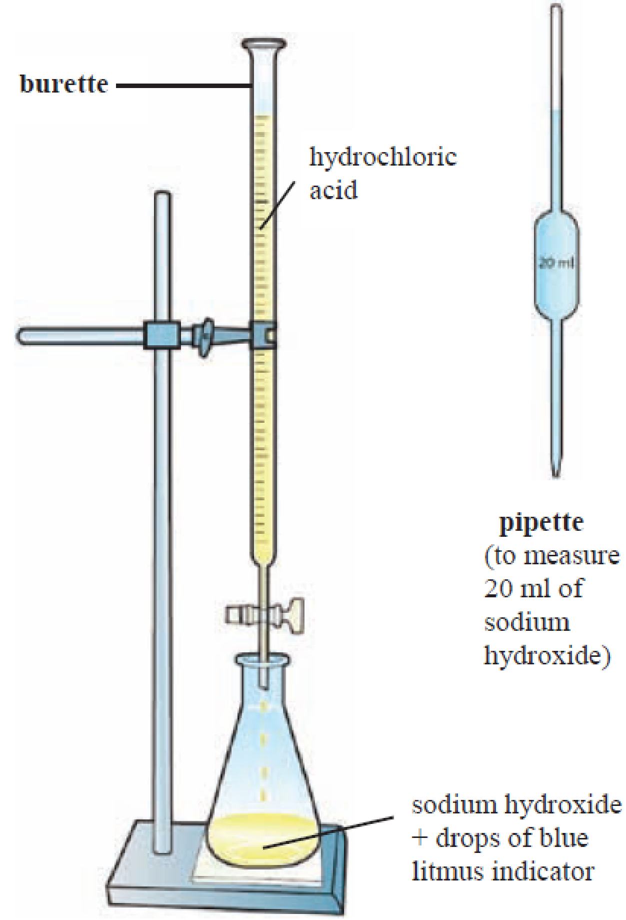
30.3 To Neutralise an Acid with a Base by Titration

A fixed volume of **hydrochloric acid** is placed in a **burette**. A **pipette** is used to measure 20 ml of **sodium hydroxide** into a **conical flask**.
A few drops of **litmus indicator** is added to the flask. A **white tile** is placed under the flask to see the colour change easily.

The **acid** is **added slowly** while the **flask is shaken**. When the solution in the flask **just begins** to change colour, **neutralisation** is complete, and the **volume of acid used** is noted. This is repeated and an **average volume** of acid needed is calculated.

The **titration** is repeated again, without litmus indicator and using the volume of acid calculated before.

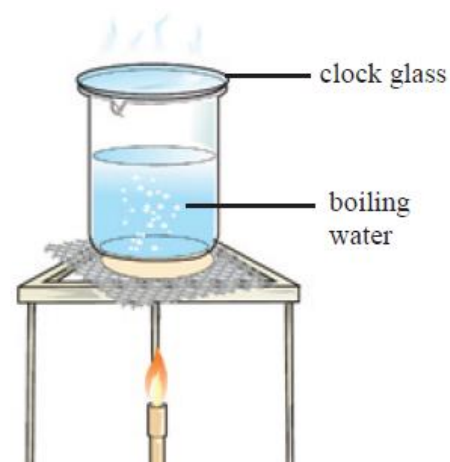
The flask now contains the salt, **sodium chloride** and **water**.



30.4 To Show the Salt formed from the Neutralisation Reaction

A sample of the **salt solution** from the flask (***30.3***) is placed in a clock glass.
The **water** in the **clock glass** is **evaporated off** as shown.

This leaves crystals of **sodium chloride** on the clock glass.



AIR

- Air consists of the **elements** nitrogen (78%), oxygen (21%), argon (<1%), and the **compounds** carbon dioxide (0.03%) and water vapour (0 - 4%).

- Air is a **mixture** because:
  - Its **composition varies** from place to place.
  - The different gases **condense** as liquids at **different temperatures**.

- **Cobalt chloride paper** is used to test for the presence of **water**. It is **blue** when **dry**, **pink** when **wet**.

- **Limewater** is used to test for **carbon dioxide**. Carbon dioxide turns limewater **milky**.

Limewater + Carbon dioxide ⟶ Chalk + Water

$$Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$$

- **Carbon dioxide** is prepared by the reaction between dilute **hydrochloric acid** (HCl) and **marble chips** (calcium carbonate, $CaCO_3$).

Calcium carbonate + hydrochloric acid ⟶ calcium chloride + water + carbon dioxide

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

- **Oxygen** is prepared by the breakdown of **hydrogen peroxide** (H_2O_2) in the presence of **manganese dioxide** (MnO_2).
- **Manganese dioxide** is a **catalyst - it speeds up** the breakdown of the hydrogen peroxide into water and oxygen.
- A **catalyst** is a substance that **speeds up** a chemical reaction.
- A **glowing splint** is used to test for **oxygen**. **Oxygen relights** a **glowing splint**.

Properties of Oxygen	
Physical	**Chemical**
1. colourless, odourless, tasteless **2**. slightly heavier than air **3**. slightly soluble in water	**1**. supports burning (combustion) **2**. very reactive element, easily forming oxides: $2Mg + O_2 \longrightarrow 2MgO$ $C + O_2 \longrightarrow CO_2$ **3**. neutral to litmus paper

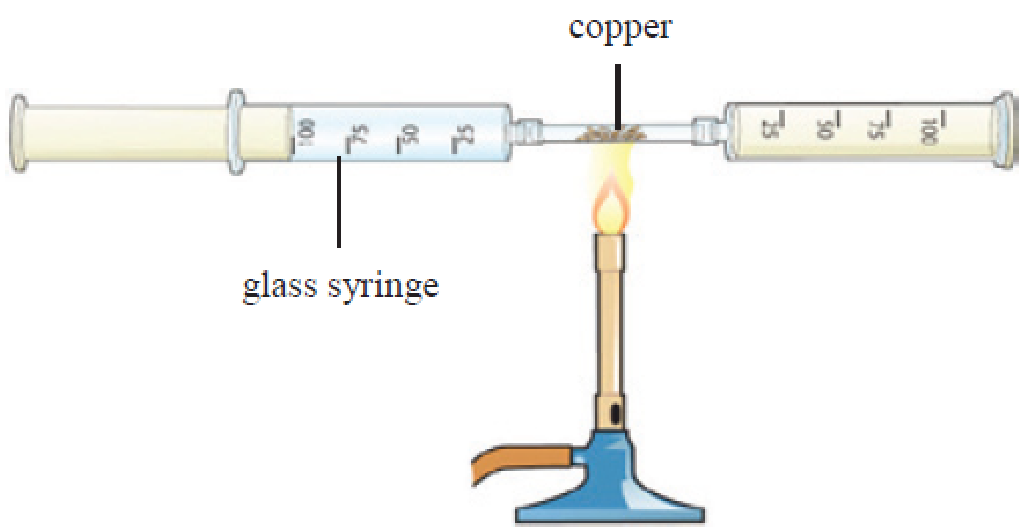
- **Oxygen** is used for **breathing**, **welding** and **burning**.

Properties of Carbon dioxide	
Physical	**Chemical**
1. colourless, odourless, tasteless 2. heavier than air 3. moderately soluble in water	1. does not supports burning 2. turns limewater milky 3. forms carbonic acid in water $CO_2 + H_2O \longrightarrow H_2CO_3$ 4. acidic to litmus paper

- **Carbon dioxide** is used in **photosynthesis**, **fire extinguishers** and **fizzy drinks**.

EXPERIMENTS:

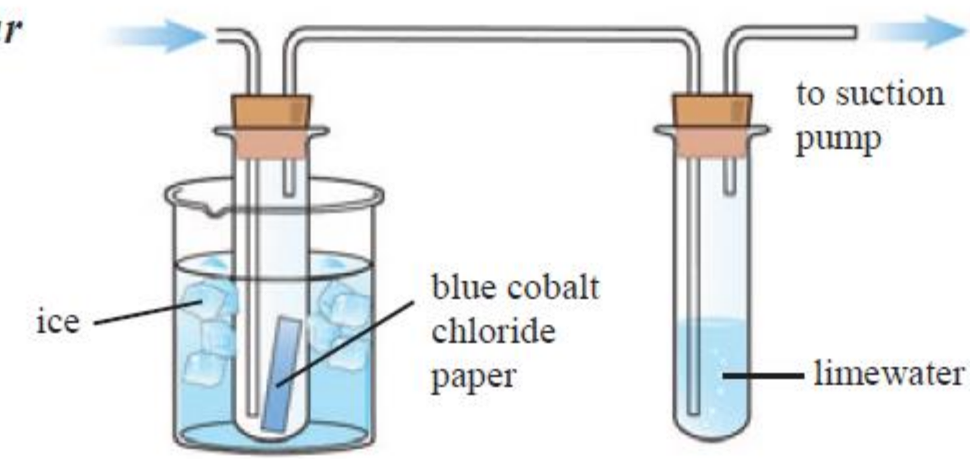
31.1 To Measure the Percentage of Oxygen in Air



100 cm^3 of air is repeatedly passed from one syringe to the other.
Copper is heated strongly in the tube.

The oxygen in the air reacts with the **copper**, forming **copper oxide**.
Only 79 cm^3 of air remains.
Oxygen comprises **21%** of air.

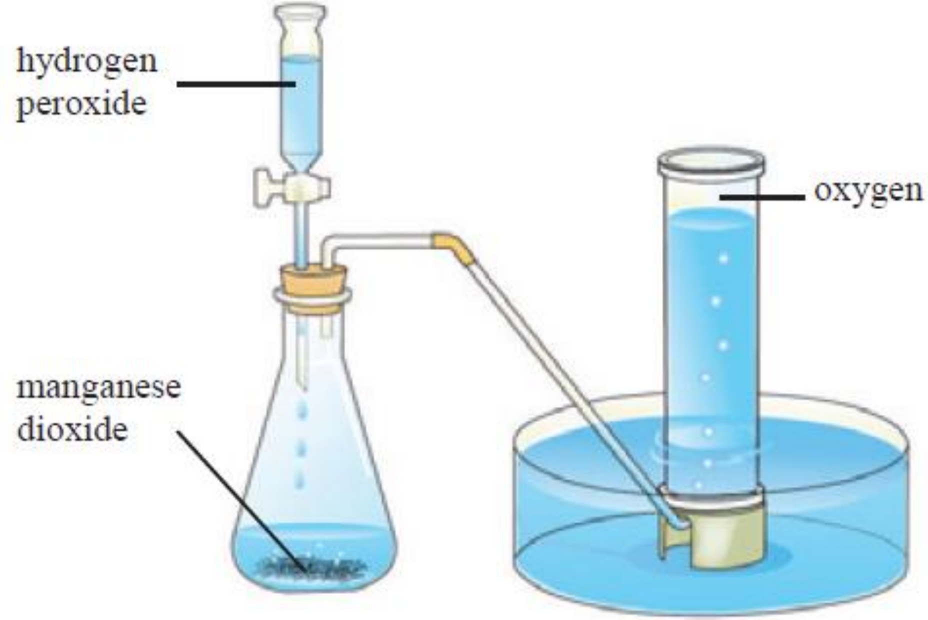
31.2 To Show the Presence of Water Vapour and Carbon Dioxide in Air



Water in the air turns the **blue cobalt chloride paper pink**.
Carbon dioxide in the air turns the **limewater milky**.
Air contains water vapour and carbon dioxide.

31.3 To Prepare Oxygen Gas

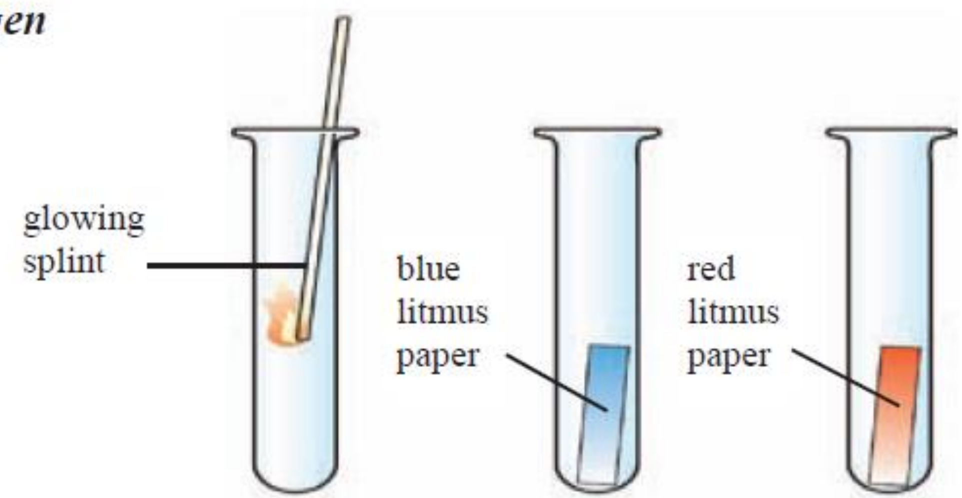
Oxygen is prepared by the breakdown of **hydrogen peroxide** in the presence of the **catalyst**, **manganese dioxide**. Oxygen is collected in the gas jar.



31.4 To Investigate the Properties of Oxygen

Oxygen relights a glowing splint.
Oxygen has no effect on moist **blue litmus paper**.
Oxygen has no effect on moist **red litmus paper**.

Oxygen **supports combustion** and is a **neutral** gas.

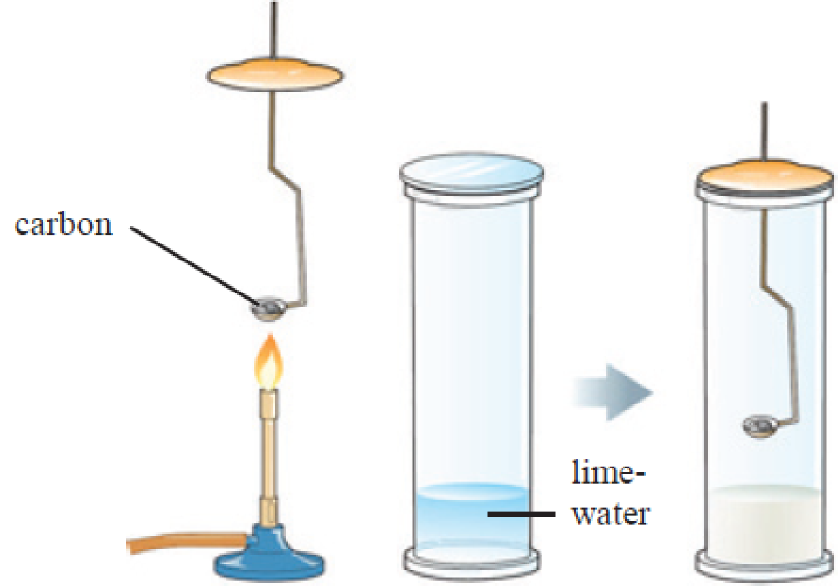


31.5 Burning Carbon and Magnesium in Oxygen

Carbon or **magnesium** are heated on a deflagrating spoon and then plunged into a jar of **oxygen**.

Carbon reacts to form **carbon dioxide** which turns some **limewater milky**.

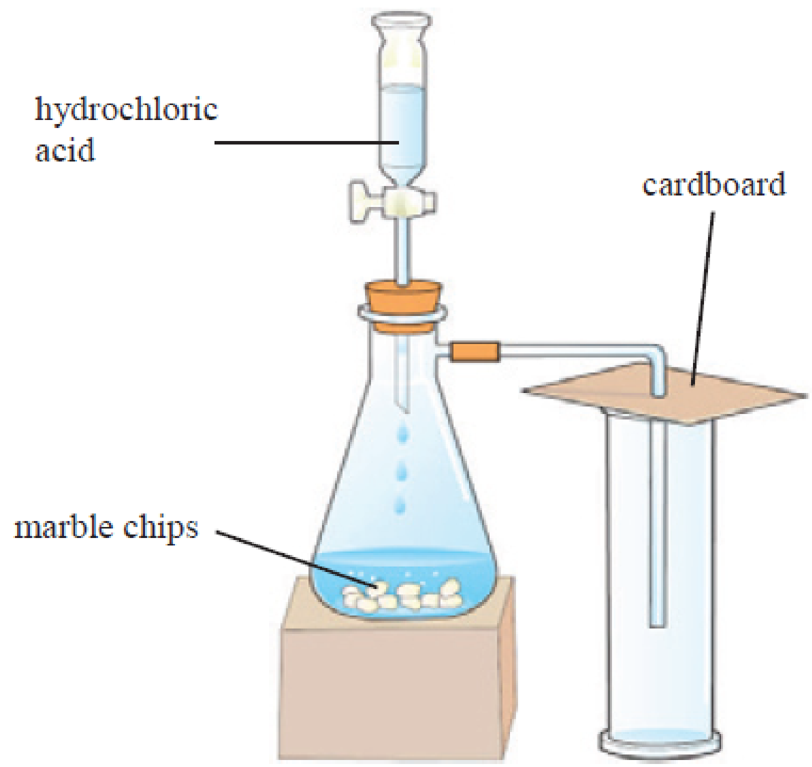
Magnesium forms white specks of **magnesium oxide - a basic oxide** which turns moist red **litmus paper blue**.



31.6 To Prepare Carbon Dioxide Gas

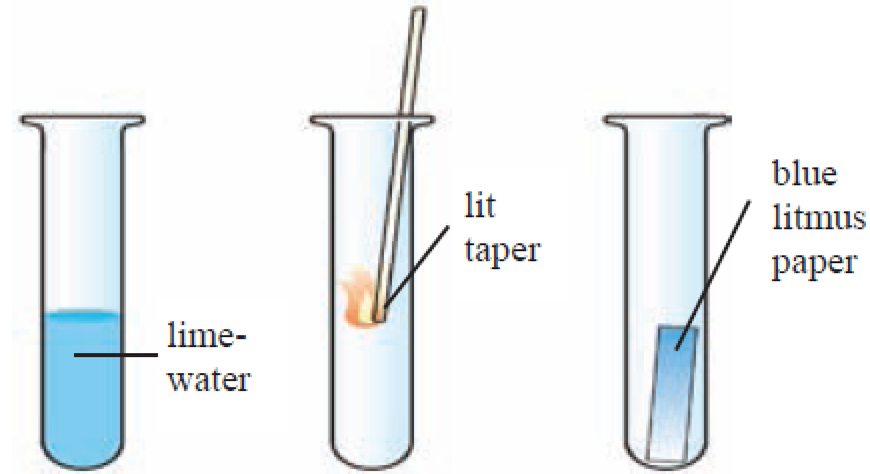
Carbon dioxide is prepared by the reaction between dilute **hydrochloric acid** (HCl) and **marble chips** ($CaCO_3$).

Carbon dioxide is **denser** than air and is therefore collected as shown.



31.7 To Investigate the Properties of Carbon Dioxide

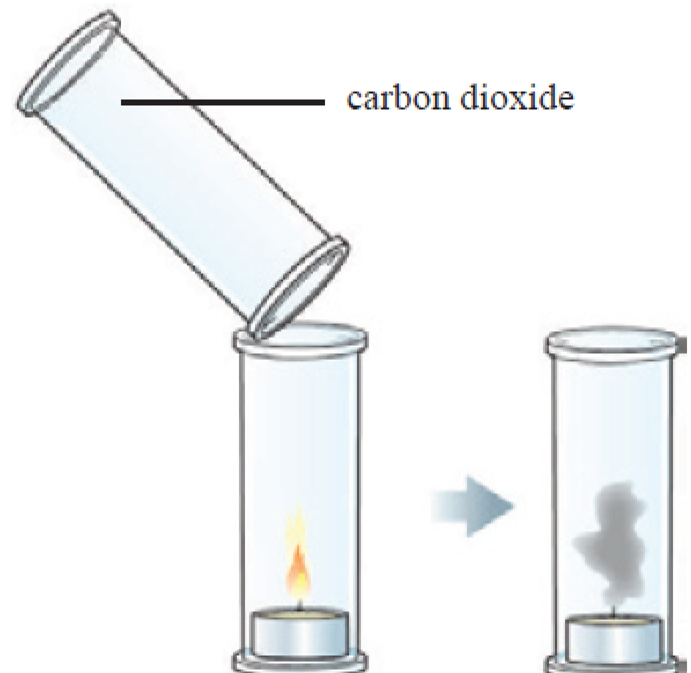
Carbon dioxide turns **limewater milky**.
Carbon dioxide does **not** support combustion.
Carbon dioxide turns moist **blue litmus paper red** - it is an **acidic oxide**.



31.8 To Show that Carbon Dioxide is Denser than Air

Carbon dioxide can be **poured** from one gas jar to another - this shows that it is **denser than air**.

The **candle** is **extinguished**, showing that carbon dioxide does **not** support combustion.



Water

- Blue **cobalt chloride paper** is used to test for water. It is **blue** when **dry -** water turns it **pink**.
- The **freezing point** of pure water is **0°C**. Its **boiling point** is **100°C**.
- **Ice** is **less dense** than water, therefore it **floats** on water.
- **Water** is an **excellent solvent**. Many substances dissolve in it.
- The **water cycle** is how water is recycled on the planet.
- **Water treatment** consists of 5 stages:
  screening (wire mesh), **settling** (settling tanks), **filtration** (sand and gravel filter beds), **chlorination** (adding chlorine to kill bacteria), and **fluoridation** (fluoride for teeth).

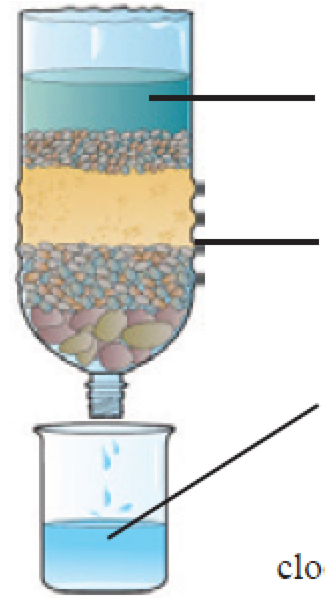
- **Hard water** is water that does **not easily** form a **lather** with soap.
- **Soft water** is water that forms a **lather easily** with soap.
- **Hardness** in water is caused by the presence of **calcium ions** dissolved in the water.
- **Calcium carbonate** (limestone) + **rainwater** (acidic) $\longrightarrow$ **calcium ions** in the water
- **Hardness** can be **removed** by passing the water through an **ion exchange resin**, or by **distillation**.
- **Ion exchange resin** exchanges **calcium ions** for **hydrogen ions** to remove hardness.
- Water can be broken down into its elements, **hydrogen** and **oxygen** by **electrolysis**.
- **Electrolysis** is the production of a **chemical reaction** by using **electricity**.
- A **Hoffman voltameter** is used in the **electrolysis** of water to form $\mathbf{H_2}$ and $\mathbf{O_2}$.
- **Hydrogen** gas forms at the **negative electrode**, **oxygen** forms at the **positive electrode**.
- **Twice as much hydrogen** as **oxygen** is formed because water is H_2O.

EXPERIMENTS:

32.1 To Use a Sand/Gravel Filter Model to Purify Water

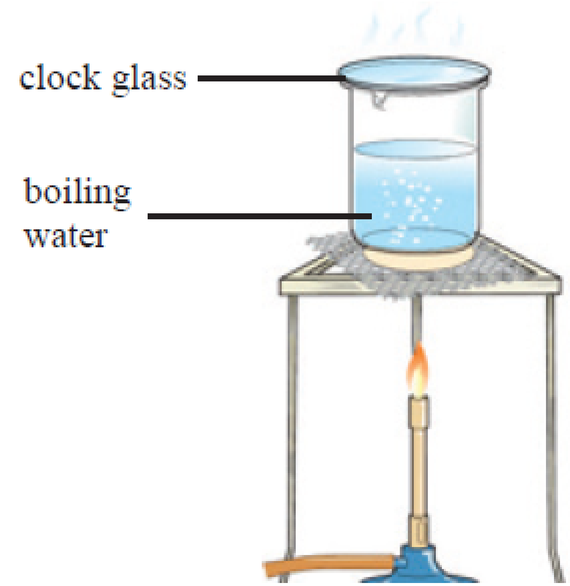
Muddy water is poured through the **filter**, set up as shown.

The water gets **filtered** and is **clear** when it enters the **beaker**.



32.2 To Show that Water Contains Dissolved Solids

Various **water samples** are **evaporated** in the **clock glass**.
Dissolved solids remain on the clock glass.
Hard waters such as **mineral water** and water from **limestone areas** contain the **most** dissolved solids. **Rainwater** contains **least**.

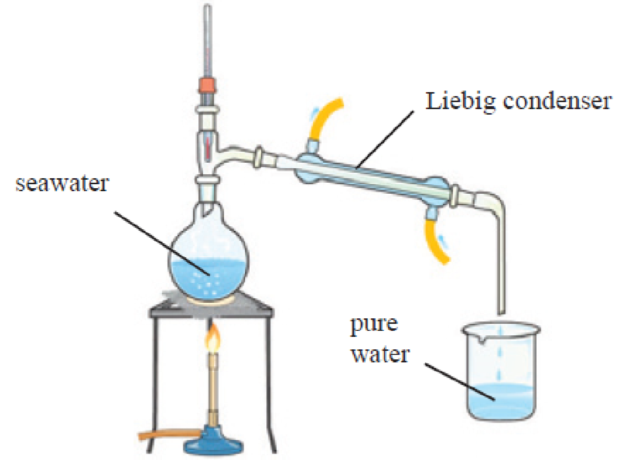


32.3 To Obtain Pure Water from Seawater

Seawater is placed in the **flask** of the **Liebig condenser** apparatus.

Pure water evaporates and condenses in the Liebig condenser and is collected in the beaker. The **salt remains** in the **flask**

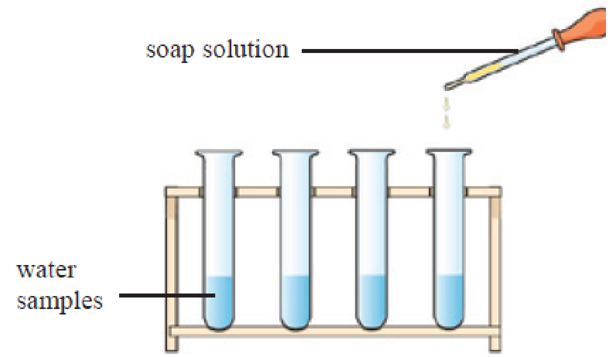
Pure water collects in the **beaker**.



32.4 To Test Various Water Samples for Hardness

Equal amounts of various **water samples** are placed in test tubes.
Soap solution is added to each test tube, a **drop** at a time, and the tube **shaken** to see if a **lather forms**.

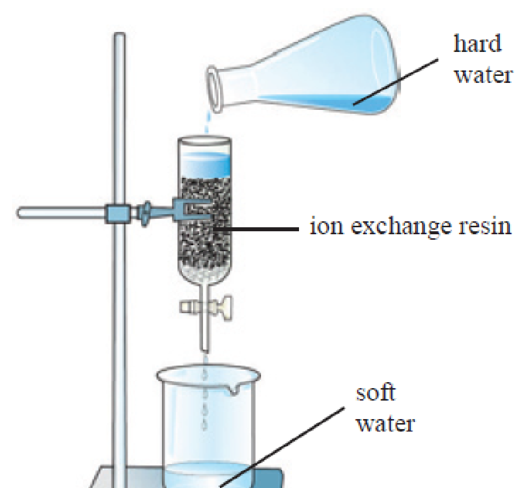
Hard water samples will need **more** drops of **soap solution** to form a **lather** than soft water samples.



32.5 To Remove Hardness from Water Using Ion Exchange Resin

A **tap funnel** is filled with ion **exchange resin**. **Hard water** is then poured through the resin as shown.

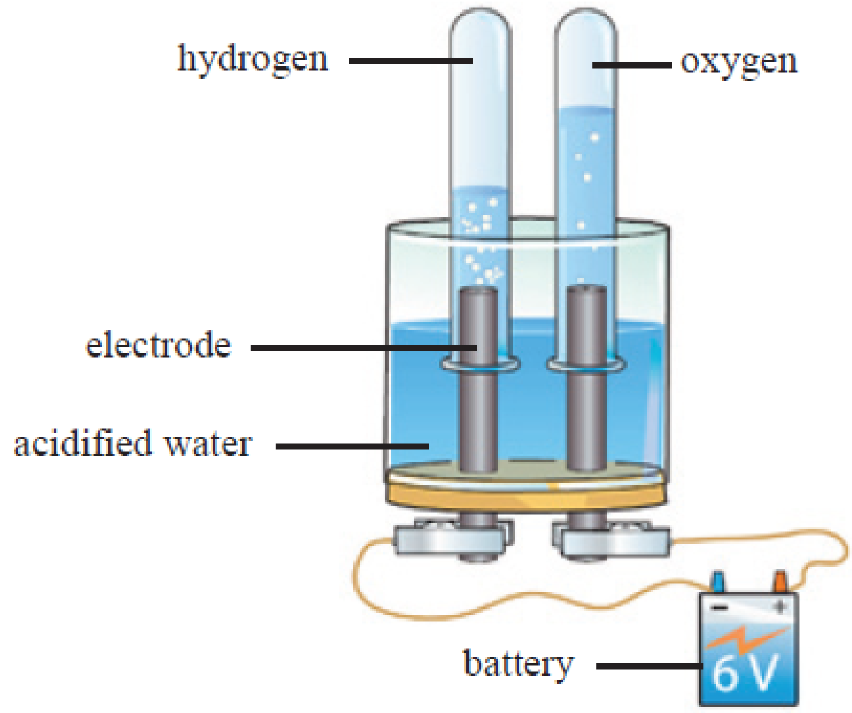
The water collected in the **beaker** is **soft water**. This can be **tested** by using **soap solution** as described in Experiment 32.4 above.



32.6 To Show the Composition of water by Electrolysis

The **Hoffman voltameter** and the two test tubes are filled with water to which a few **drops of acid** have been added to help **conduct** the **electricity**.

Oxygen collects in the test tube above the **positive electrode -** it can be **tested** for by using a **glowing splint**, which **relights**.
Twice as much **hydrogen** collects in the test tube above the **negative electrode -** it can be **tested** for by using a **lit taper**.
Hydrogen gives a '**pop**' sound when lit.



Groups of Elements

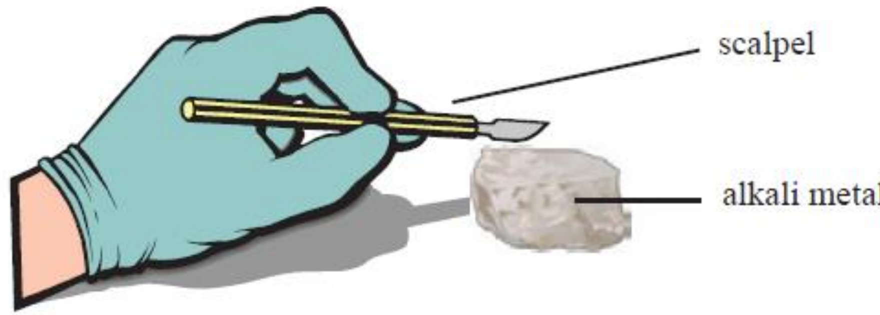
- **Atoms** of **elements** in the **same Group** all have the **same number** of **electrons** in their **outermost shells**.
- All **elements** in the **same Group** show **similar chemical properties**.
- **Group I** elements, the **alkali metals** are **soft** and **shiny**, with **low densities**.
- The first three alkali metals are: **lithium**, **sodium**, and **potassium**.
- The alkali metals react with **oxygen** to form **oxides**.
  Sodium metal gets a white coat of **sodium oxide** when it is cut with a knife.
- The alkali metals react with **water** to form **hydroxides** and **hydrogen** gas.
  Sodium metal in water produces hydrogen gas which bursts into flame.
- **Reactivity increases** going **down** the Group in the alkali metals.
- **Lithium** is used to make watch and camera batteries.
- **Sodium** is used in street lights that give a soft orange glow.
- **Potassium** is used in plant fertilisers.

EXPERIMENTS:

33.1 To React an Alkali Metal with Air

A sample of **sodium** or **lithium** is cut with a **scalpel**.

The shiny cut surface gains a **dull** layer of **sodium** or **lithium oxide**.
Alkali metals react easily with **oxygen** in the air to form the **metal oxide**.

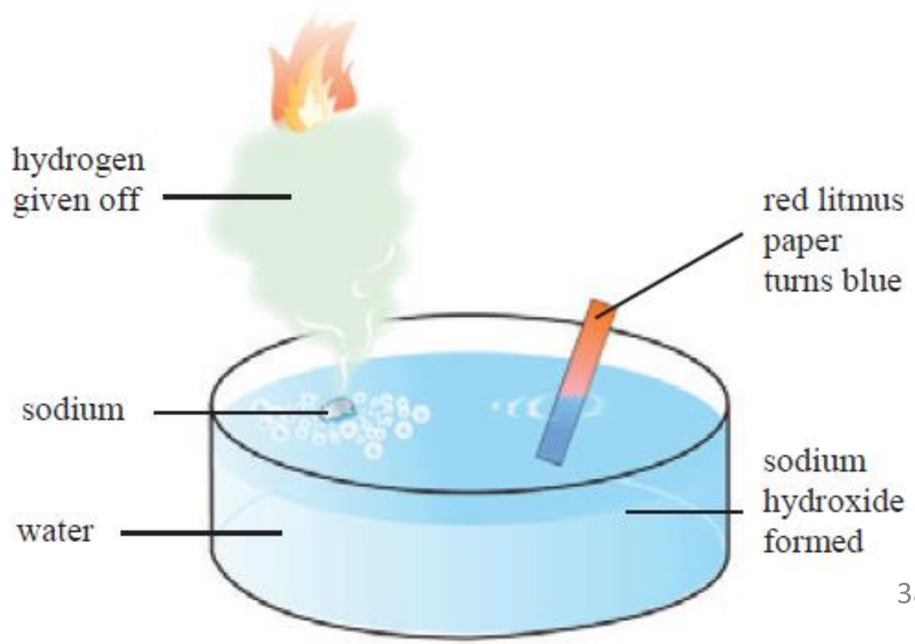


33.2 To React an Alkali Metal with Water

A sample of **sodium**, lithium or potassium is placed in water.

Hydrogen gas, which sometimes lights and explodes, is given off.
The **metal hydroxide** (e.g. NaOH) is formed and **dissolves** in the **water**.

The **alkali** formed turns **red litmus** paper **blue**.



Metals

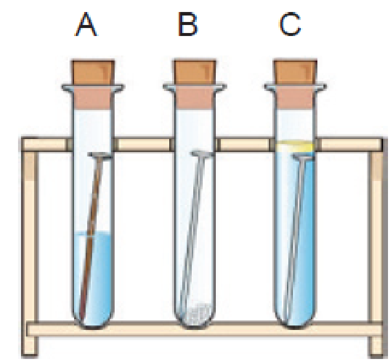
- Metals are usually **dense**, **hard**, **shiny solids** with **high** melting points.
- Metals are **ductile** and **malleable**, are **good conductors** of **heat** and **electricity**, and most **corrode**.
- **Corrosion** is a reaction in which a metal **slowly reacts** with **oxygen** or some other element in the air, to form an **oxide** or some other compound.
- The **corrosion** of **iron** or **steel** is called **rusting**.
- **Rusting** takes place in the presence of **water** and **oxygen** (air).
- Corrosion is **prevented** by **painting**, **greasing**, **galvanising** or **chrome plating**.
- **Galvanising** involves coating **iron** with a **layer of zinc**.
- An **alloy** is a **mixture** of **metals**.
- **Bronze** (copper and tin), **brass** (copper and zinc), **steel** (iron and carbon) and **solder** (lead and tin) are all metal alloys.

- The **Activity Series** is a list of metals placed in order of **how reactive** they are.
- When **metals** react with an **acid**, **hydrogen** gas is given off.
- **Zinc** reacts with **hydrochloric acid** to form **zinc chloride** and **hydrogen**.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

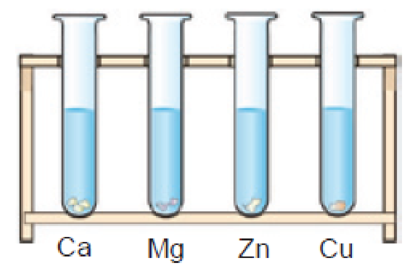
EXPERIMENTS:

34.1 To Investigate the Conditions Necessary for Rusting



Tube **A** has both **water** and **oxygen** - this is the **control**.
Tube **B** has **no water** - removed by **calcium chloride**.
Tube **C** has **no oxygen** - cooled, **boiled water** was used.
Only the nail in test tube **A** will rust.

34.2 To Compare the Reactivity of Metals

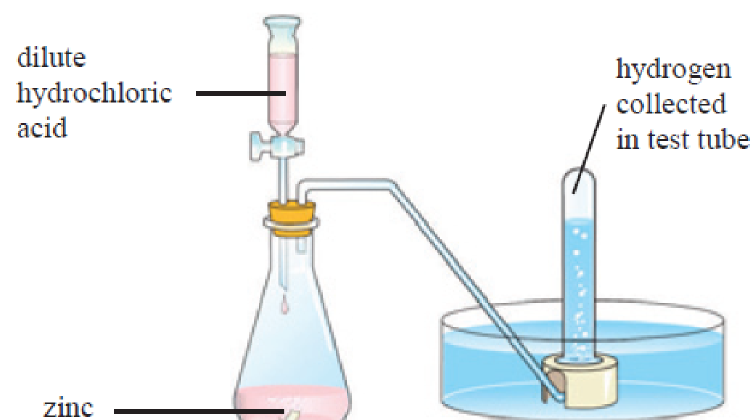


When tested with **water** and dilute **hydrochloric acid**, **calcium** is found to be **most reactive**, then **magnesium**, **zinc** and finally, **copper**.

34.3 To React Zinc with Hydrochloric Acid and Test for Hydrogen

Dilute hydrochloric acid reacts with **zinc** to form **zinc chloride** and **hydrogen** gas.

The hydrogen is collected in a test tube and tested with a **lit taper** - it goes **'pop'**.



Chemistry in Everyday Life

- **Fossil fuels** are formed from the **remains** of **plants** and **animals** that lived millions of years ago.
- **Coal**, **oil** and **gas** are fossil fuels.
- All **fossil fuels** contain the elements **carbon** and **hydrogen**.
- **Natural gas** is mainly **methane** (CH_4).
- **Fossil fuels** produce **carbon dioxide** (CO_2) and **water** (H_2O) when burned.
- Some **fossil fuels** (coal and oil) contain **sulfur compounds** which release the gas **sulfur dioxide** (SO_2) on **burning**.
- **Sulfur dioxide** combines with **rainwater** to form **sulfuric acid**, causing **acid rain**.
- Acid rain **kills fish**, is harmful to **plants** and erodes **limestone buildings**.

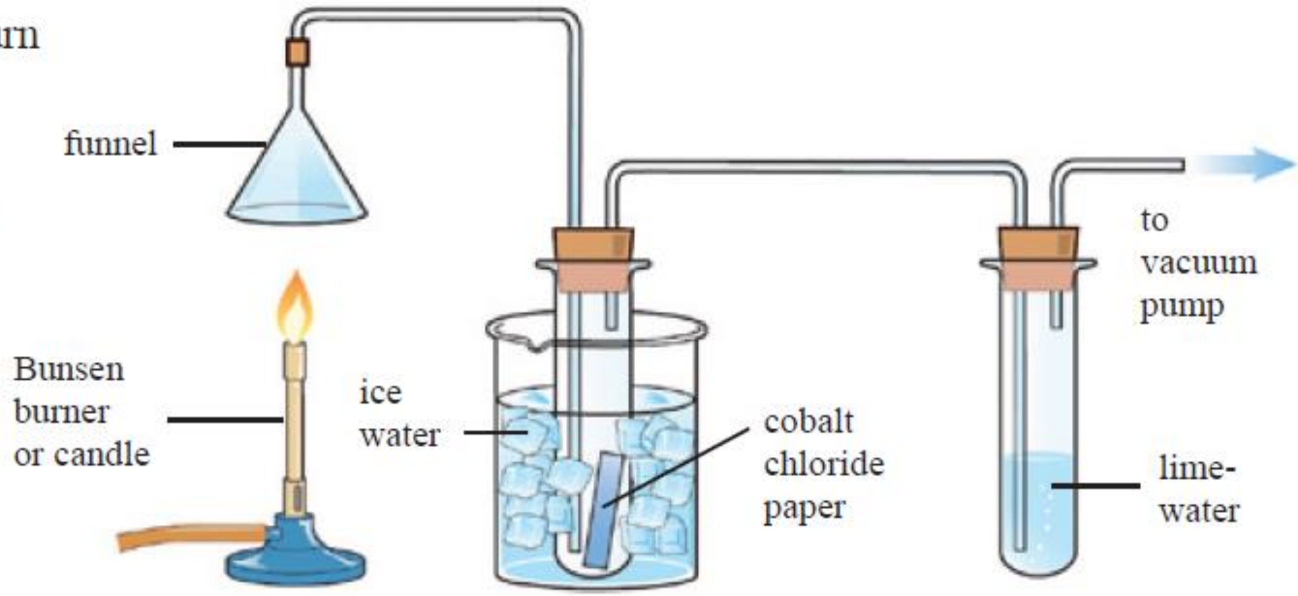
- **Plastics** are man-made materials made from chemicals called **monomers** extracted from **crude oil**.
- The **monomers** are **reacted together** to make long chains of molecules called **polymers** which are used to make plastics.
- **Plastics** are **lightweight**, **durable**, **waterproof**, **easy** to **clean** and can be easily **moulded** into different shapes.
- **Plastics** are **non-biodegradable** which means they cannot be broken down by living organisms, such as bacteria and fungi in the soil.

EXPERIMENTS:

35.1 To Investigate the Products of Combustion

Use a very low flame to burn **methane gas** as shown.

Water from the burning fuel turns the **blue cobalt chloride paper pink**. **Carbon dioxide** from the burning fuel turns the **limewater milky**.



35.2 To Show the Effect of Acid Rain on Limestone

Dilute sulfuric acid is dropped onto **limestone**. This acid is found in **acid rain**.

The **limestone fizzes** as the sulfuric acid reacts with it.