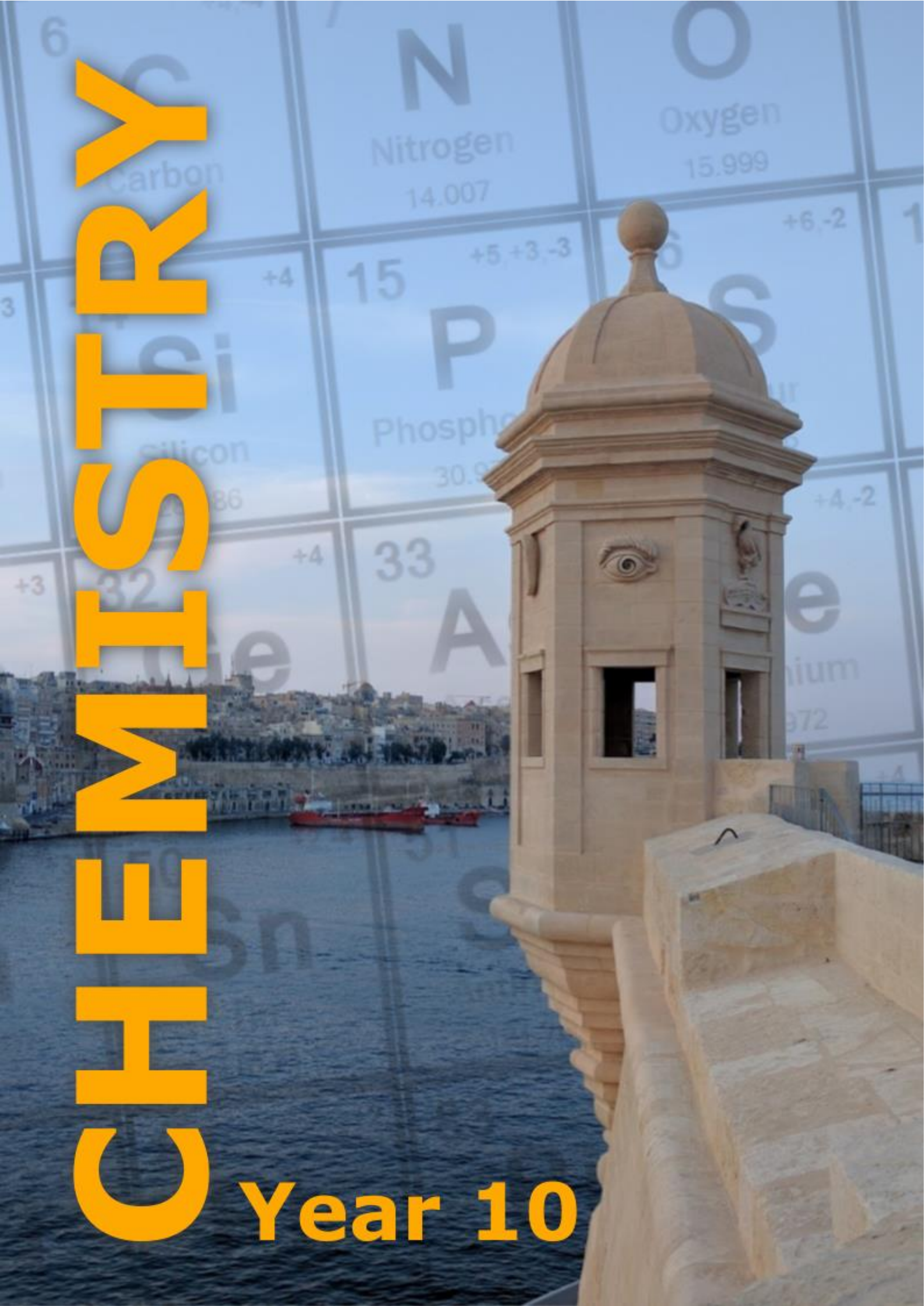

CHEMISTRY

Year 10

This textbook will be available in digital format (PDF) only. It is intended to be used by secondary school students and teachers in Malta and Gozo. It leads students to sit for the SEC06 Chemistry examination from 2025 onwards.

Version history

- Version 1.0 distributed in September 2023.

Acknowledgements

Michael Mercieca (EO) contributed towards the frontpage and typesetting.

Dr Doreen Mizzi (HoD), Robert Zammit (HoD), Dennis Zarb (HoD), and Michael Mercieca (EO) edited this textbook.

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How to use this e-book

This e-book is closely linked to the SEC 06 (2025) Chemistry syllabus which can be downloaded at: SEC 06 syllabus (um.edu.mt)

Each chapter represents the work expected in one learning outcome. Each chapter starts with a graphic that indicates the subject focus of which the learning outcome is part of.

Chapters are split into sections that deal with groups of related assessment criteria. These assessment criteria are declared so that students can focus on what needs to be learned. Some sections are accompanied by a series of questions that are related specifically to the content in that section.

Hyperlinks link to content that is online and serve as further reading. Online content is not examinable however students are encouraged to follow these links to further their understanding of the role of chemistry in the real world.

An internet connection is required to access video clips, animations, and online content. When using the digital version of this e-book, QR codes can be clicked to access the online resources.

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Throughout this e-book you will find one of five types of textboxes shown below:

This textbox shows **exercises** related to content that students can work on their own or with their teachers.

This textbox shows **snippets** of information that are not meant to be studied for your examination, but which give extra information related to discoverers, discoveries, and further relevant content.

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Also,

Definitions and **important information** is presented on a blue background.

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1 Physical and chemical changes – LO 9

1.1 Distinguishing between physical and chemical changes

By the end of this section, you should be able to:

- Explain that some substances are useful in their native state and that other substances need to be changed by chemical reactions to be more useful. (9.2a)
- Compare chemical reactions with physical changes. (9.1b)
- Name the changes that take place when chemical reactions occur. (9.2b)
- Explain that when chemical reactions happen mass is conserved. (9.2e)

Substances can undergo changes which can be classified as either **physical** or **chemical**.

1.1.1 Usability of materials

Some elements and compounds are naturally found and used in their native state. Such examples are gold which is used in jewellery and electronic contacts whereas limestone (calcium carbonate) is used in the building industry.

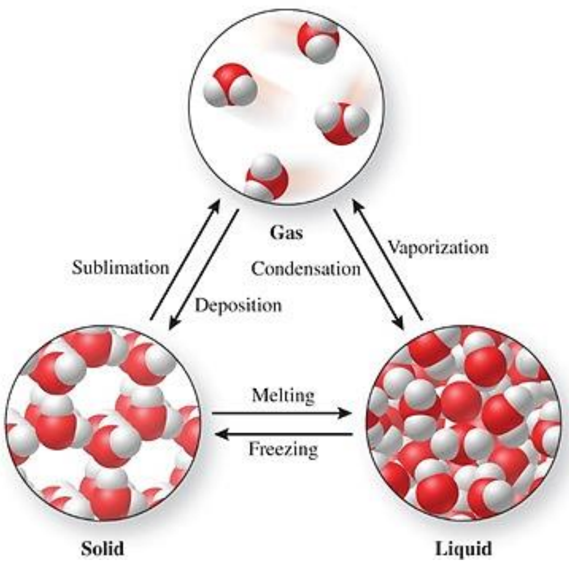
Other substances are obtained from ores or other chemicals by using different chemical reactions. For example, iron(III) oxide , also known as haematite, contains iron which has many uses in various alloys and structures. However, iron needs to be extracted from iron(III) oxide before it can be of any use. Therefore, iron(III) oxide needs to be chemically changed to iron metal.

1.1.2 Physical changes

When a physical change happens, **no new substances are formed**.

These physical changes are often changes in state such as melting, freezing, boiling, and condensing. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the material does not. Physical changes that involve a change of state are **reversible**.

For example, when liquid water is heated, it changes to water vapour. Even though the physical properties have changed, the molecules are the same as before. This change is reversible. Likewise, melted ice may be refrozen, so melting is a reversible physical change as well. The diagram shows that the molecules of water do not change when it undergoes changes of state.



Also, dissolving a salt in water is a reversible change since the mixture can be easily separated using physical separation techniques.

Other situations like the cutting of a ribbon or the breaking of glass are irreversible physical changes since although no new substance is formed it is difficult to reverse the process to regain the original object.

1.1.3 Chemical changes

When a chemical change happens, **matter is neither created nor destroyed**.

Instead, the particles of a substance rearrange themselves **to form new chemicals**.

This means that a substance with a certain set of properties (such as melting point, colour, density, etc.) changes into a different substance with different physical and chemical properties. Substances that react together (also known as **reactants**) form new substances (also known as **products**).

Chemical changes are frequently harder to reverse than physical changes. A chemical change is usually accompanied by one or more of the following:

- Production of light (flame).
- Effervescence (production of gas bubbles).
- Change in smell.
- Change in colour.
- Temperature change.
- Products formed may have a different state of matter than the reactants.

An example of a chemical change is burning LPG on a stove. This results in the formation of new chemicals (carbon dioxide and water vapour).

1.1.4 The law of conservation of mass

During chemical or physical changes mass is conserved, meaning that **matter is neither created nor destroyed**.

In a chemical reaction,

the total mass of the reactants is equal to the total mass of the products.

This is why in a chemical reaction the total number of elements on the left-hand side is equal to the total number of elements on the right-hand side. For example, when magnesium is burned in air, the total mass of magnesium and oxygen used in the reaction is equal to the total mass of magnesium oxide produced.

$$\begin{array}{c} 2Mg(s) + O_2(g) \rightarrow 2MgO(s) \\ 48\text{ g} \quad + 32\text{ g} \rightarrow \quad 80\text{ g} \end{array}$$

This is called the **law of conservation of mass**.

The law of conservation of Mass

Before the principle of the conservation of mass was proposed through the work of Mikhail Lomonosov (in 1756) and Antoine Lavoisier (in 1773), it was believed that in certain reactions the mass of the initial reactant is lost. This was concluded from various observations, like the decrease in mass of a log of wood after it burns.

However, this was disproven when reactions (like rusting) where held in a sealed glass container. The theory of the conservation of mass helped chemists understand that a chemical can undergo a transformation and produce new substances, while maintaining the same mass throughout the reaction.

1.2 The three states of matter

By the end of this section, you should be able to:

- State the physical properties of the three states of matter. *(Limited to compressibility, ease of flow, and shape.)* (9.1c)
- Describe using diagrams, the arrangement, movement of particles, and forces of attraction between particles in the three states of matter. *(Forces of attraction limited to strong and weak forces.)* (9.2c)
- Interpret the physical properties *(E.g. compressibility, ease of flow, shape)* of the three states of matter in terms of the kinetic theory. (9.3c)

A substance can exist in one of three states of matter that is a solid, a liquid, or a gas. The following table shows the differences in physical properties between these three states of matter in terms of shape, ease of flow, and compressibility.

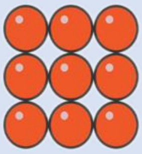
	SOLIDS	LIQUIDS	GASES
Shape	Solids have a fixed shape and volume at a particular temperature. The volume can be affected by a change in temperature.	A fixed volume of liquid takes the shape of a container occupying the bottom part of it. The volume can be affected by a change in temperature.	Gases do not have a fixed volume and take the shape of the whole container. As the temperature of the gas increases the volume increases as long as the gas can expand freely.
Ease of flow	Solids do not flow.	Liquids flow easily and they can be poured from one container to another. They are also called fluids.	Gases flow easily as they can occupy the whole container, they are in. Like liquids they are also called fluids.
Compressibility (Compression is a force that is applied on a material.)	Solids cannot be compressed.	Liquids cannot be compressed.	Gases are easy to compress.

Matter is made up of tiny moving particles, invisible to the naked eye. We can better understand this using **THE KINETIC THEORY** which states that:

All matter consists of particles in constant random motion.

The kinetic theory explains the properties of the different states of matter. The particles in solids, liquids, and gases have different amounts of kinetic energy. They are arranged differently and move in different ways.

The following table shows how the kinetic theory of matter describes and explains the three states of matter:

Properties	Solid	Liquid	Gas
Arrangement of particles.	Particles are arranged very close together with very little space in between the particles. They have a regular pattern of arrangement. 	The particles are still very close together, however there is no defined structure. 	The particles in a gas are very far apart from each other. 
Movement of particles.	Particles vibrate about fixed positions. The more energy they have, the more they vibrate.	Particles have more kinetic energy than in a solid and move about in different directions over short distances.	Particles have more kinetic energy than in a liquid and move randomly and independently at very high speeds. They move in all directions over longer distances.
Forces of attraction between particles.	Particles are held by **strong** forces of attraction keeping them tightly packed.	The forces of attraction between particles are **weaker** than in a solid due to a higher kinetic energy of the particles.	The forces of attraction between particles are so weak that they are considered as **negligible**.

The physical properties of the three states of matter can be explained using the kinetic theory.

	SOLIDS	LIQUIDS	GASES
Shape and ease of flow	Solids have a fixed shape and volume and cannot flow because particles cannot move from place to place due to strong forces of attraction.	A fixed volume of liquid flows and takes the shape of a container because the attraction between particles is no longer capable of holding the particles in their fixed positions.	Gases do not have a fixed volume and flow into the shape of the whole container because the forces of attraction between particles is negligible.
Compressibility	Solids cannot be compressed because particles are tightly packed and the space between particles is limited.	Liquids cannot be compressed because although particles can move over each other, they are still closely packed with limited space between them.	Gases can be compressed because the space between particles is relatively large, and the particles can pushed closer together.

The States of Matter

In the Periodic Table we can find 2 elements only that are liquids at room temperature. These are bromine and mercury. There are four solid elements (Rb, Fr, Cs and Ga) that have a melting point slightly above room temperature. These can easily melt into a liquid when some heat is applied.

Moreover, liquids cannot exist where there is no pressure, such as in outer Space. In the Universe we can only find liquids on surfaces or interiors of planets and moons. When liquids are exposed to conditions similar to those in outer Space (where no pressure is found) they will immediately boil if the temperature is high or freeze if the temperature is low.

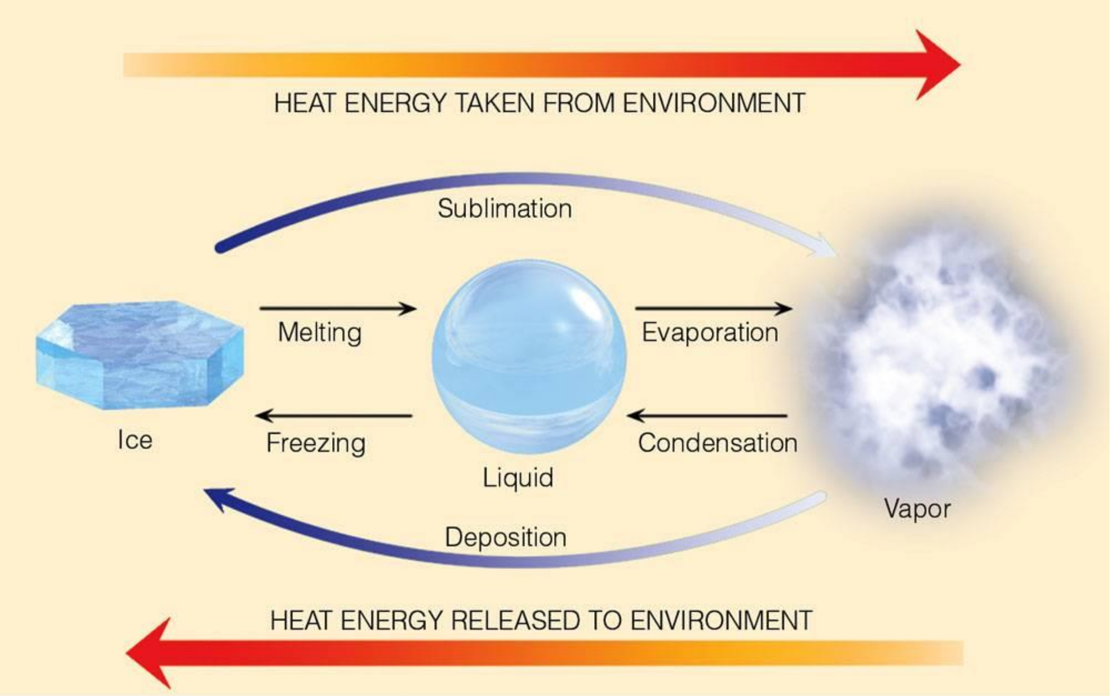
Further reading on the three states of matter.

1.3 Changes of state

By the end of this section, you should be able to:

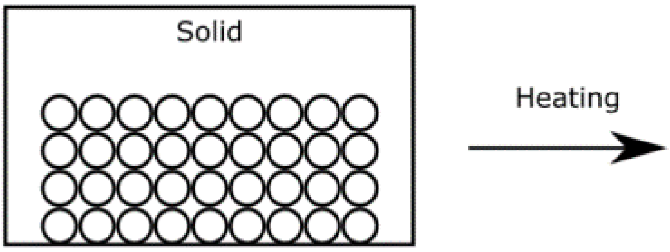
- Name the six changes of state. *(Melting, freezing, evaporation/boiling, condensation, sublimation, and deposition.)* (9.1d)
- Interpret the shape of heating/cooling curves. *(Without reference to the kinetic theory.)* (9.2d)
- Explain energy changes accompanying changes of state using the kinetic theory of matter. (9.3d)

The following picture shows all the changes of state and their relationship with heat energy.



1.3.1 Melting

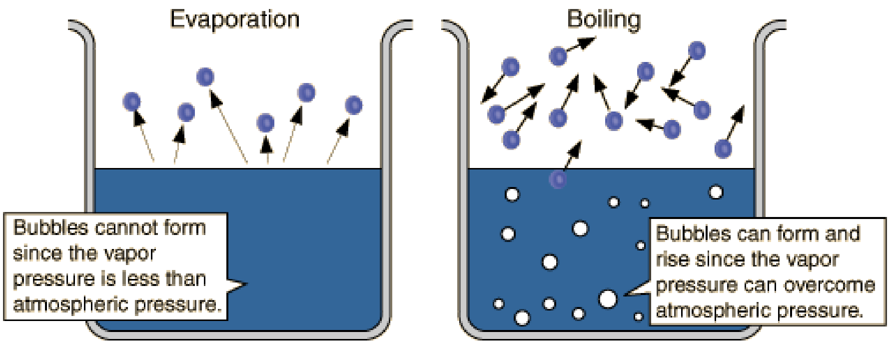
When a solid is heated the particles vibrate faster as they gain kinetic energy and the attractive forces between particles are weakened.



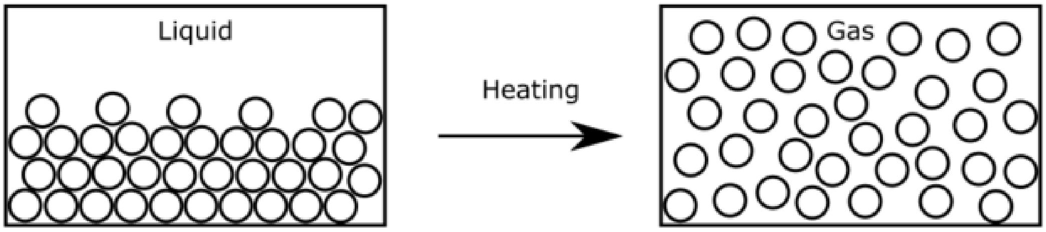
Eventually, at the **melting point**, the attractive forces are no longer able to hold the particles in an ordered way. So the solid melts to change the state and become a liquid as the particles lose their fixed arrangement. Different substances have different melting points. In fact this property can be used to distinguish substances from each other.

1.3.2 Evaporation and boiling

During **evaporation**, the particles of the liquid become free to form a gas (or vapour). Evaporation occurs **on the surface** of a liquid at **any** temperature. The process occurs slowly and cannot be observed by the naked eye. Evaporation occurs when there is exposure of a liquid to air and liquid molecules change into vapour.



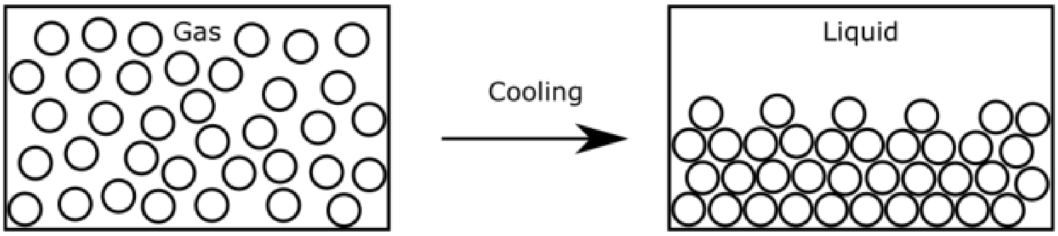
Boiling occurs in the entire mass of a liquid and bubbles of gas can be observed throughout the liquid. For pure liquids, boiling occurs at a fixed temperature (boiling point), while evaporation can happen at any temperature depending on varying environmental conditions.



Both evaporation and boiling require heat energy.

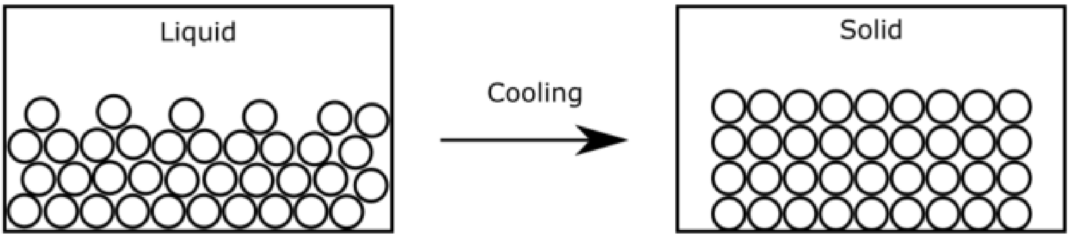
1.3.3 Condensation

On cooling, gas particles lose kinetic energy and as they slow down, attractive forces cause the particles to come closer together to form a liquid.



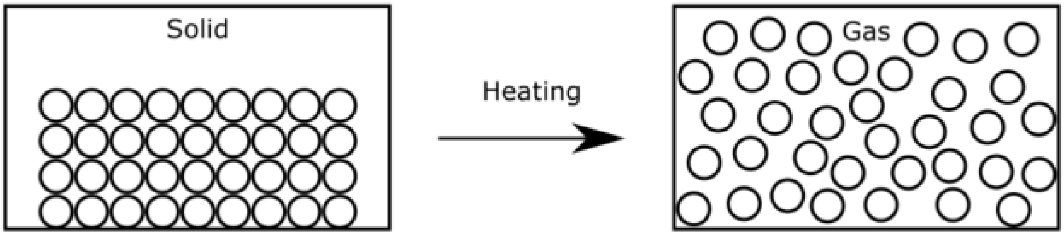
1.3.4 Freezing

On cooling, liquid particles lose kinetic energy and so can become more strongly attracted to each other. Eventually at the **freezing point** the forces of attraction are sufficient to remove any remaining freedom of movement and the particles come together to form an ordered solid arrangement. The freezing point of different substances has a different value. This value may be used to distinguish between substances.



1.3.5 Sublimation

Under certain conditions, some solids turn straight into a gas when heated. This process is called **sublimation**.



An example is solid carbon dioxide, also called 'dry ice'. At atmospheric pressure, it turns straight into gaseous carbon dioxide. Liquid carbon dioxide can only exist under high pressure, such as in fire extinguishers. Iron(III) chloride also sublimes at atmospheric pressure.

On heating, particles in a solid gain kinetic energy and as they move faster, they overcome the attractive forces causing the particles to separate from each other to form a gas.

Sublimation

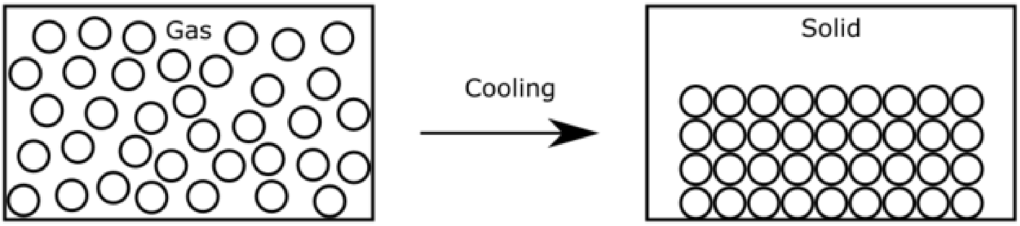
Have you ever heard of Mothballs? These are small white balls usually placed next to clothes, to protect them from insects that would otherwise damage the fabric. These mothballs are made of a chemical called naphthalene, a white solid which at room temperature, slowly sublimes into a toxic gas.

In 1821 John Kidd, an English chemist, documented his findings related to the production of naphthalene. He was able to produce a list of its properties, like its crystalline appearance, its highly pungent and aromatic odour, and its sublimation into gas. It was later, in 1948, that naphthalene was registered in the United States, as a pesticide.

In 2008 it was banned in the EU due to its cancer risks; mothballs are known to cause liver and kidney damage.

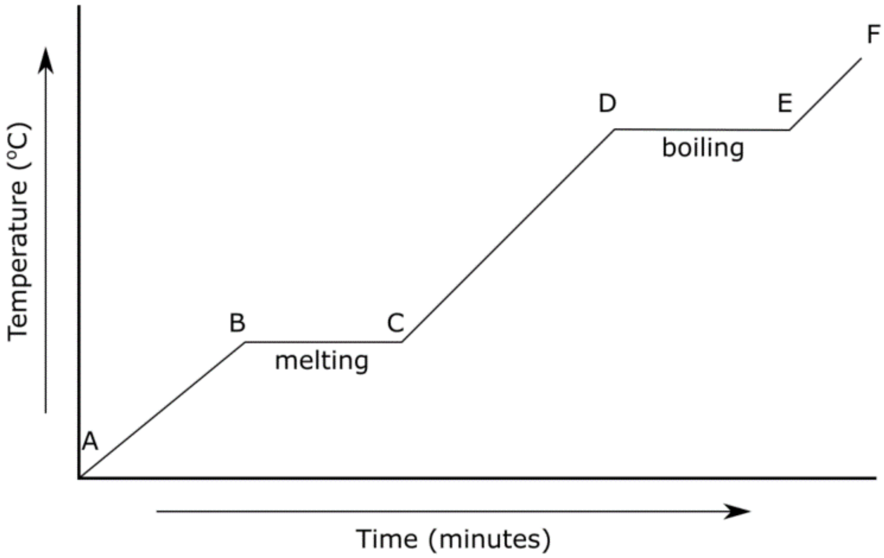
1.3.6 Deposition

During deposition a substance that is found in the gaseous state will deposit straight into the solid phase skipping the liquid phase. On cooling, gas particles lose kinetic energy and as they slow down, attractive forces cause the particles to come closer together to form a solid. An example of a substance that undergoes deposition is carbon dioxide.



1.3.7 Heating curves

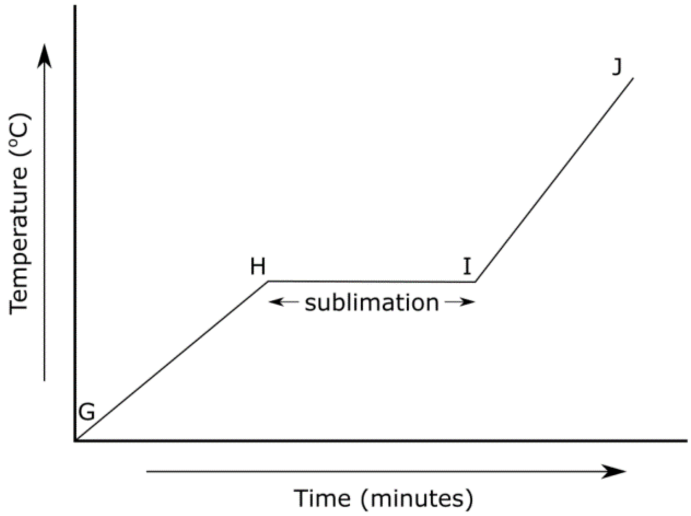
A heating curve is the graph that is plotted for a substance that is heated such that it changes from a solid to a liquid and then from a liquid to a gas. Consider the heating curve of a pure substance below:



The heating curve has letters from A to F that shows different stages. These stages are described below:

- **A – B:** Substance exists in the solid state, although the temperature is increasing with time.
- **B - C**: Point B is called the **melting point**. The graph is flat showing that the temperature remains constant, although time is passing. This is the melting phase of a substance. The fact that stage B - C is flat shows that the substance is pure. If point C is higher than point B then this would mean that the substance is impure. At point C all of the solid has changed to a liquid.
- **C – D**: Substance exists as a liquid although the temperature increases as time passes.
- **D – E**: Point D is called the **boiling point**. The graph is flat showing that the temperature is constant. As long as heat continues to be provided, the liquid continues to boil until all of the liquid turns into a gas.
- **E – F**: The temperature of the gas increases as more energy is provided to the substance.

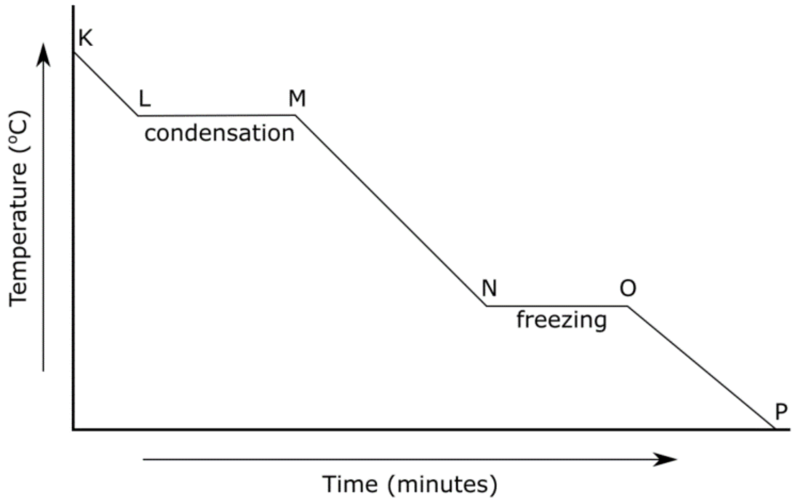
Consider the heating curve shown below for a substance that sublimes.



The heating curve for a substance that sublimes only shows one change of state. In the diagram above, G – H shows the substance in the solid state while I - J shows the substance in the gaseous state. Sublimation occurs at point H. During stage H – I the solid substance sublimes and changes to a gas. The temperature is constant until the change of state takes place.

1.3.8 Cooling curves

Consider the cooling curve shown below:



The cooling curve has letters from K to P that shows different stages. Cooling curves show how the temperature changes as a substance is cooled.

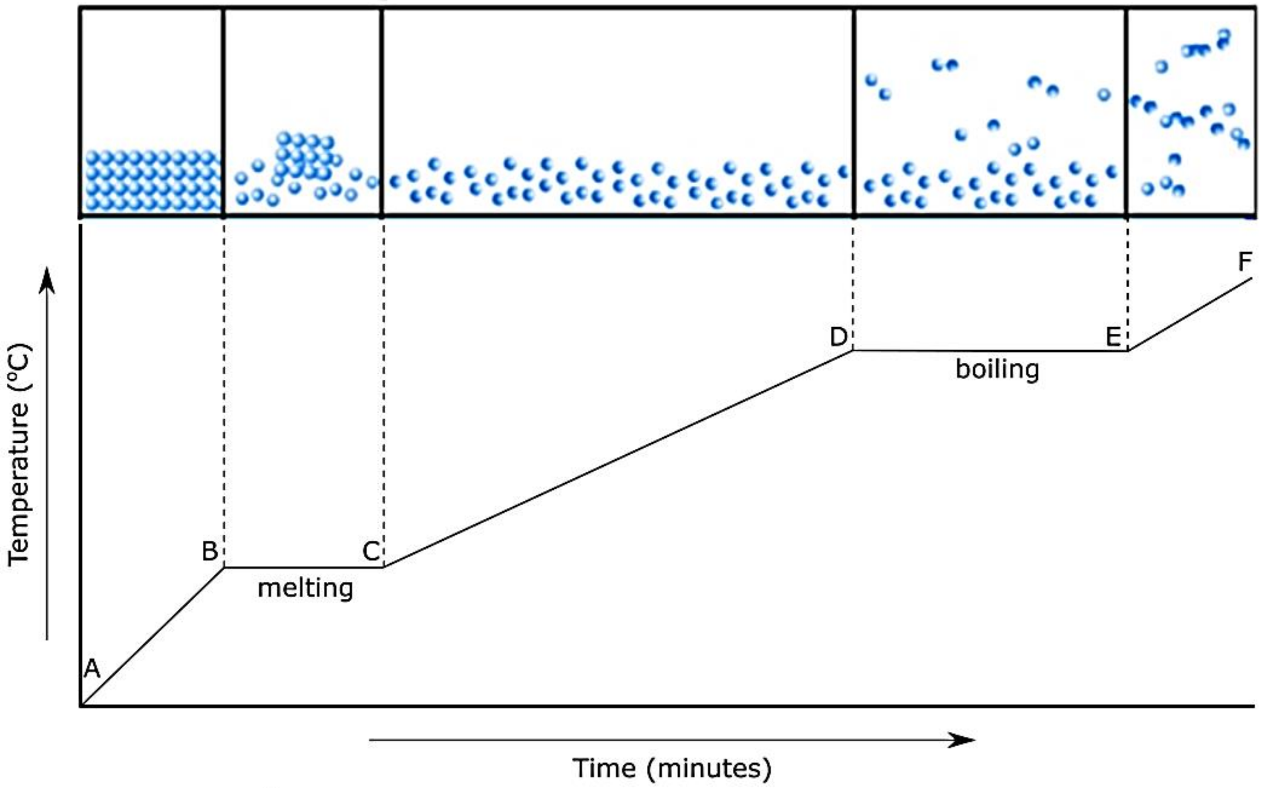
Stage K – L shows the substance in the gaseous state, stage M – N shows the substance in the liquid state while stage O – P shows the substance in the solid state.

For pure substances, just like heating curves, cooling curves have horizontal flat parts where the state changes from gas to liquid (stage L – M) known as **condensation** and from liquid to solid (stage N – O) known as **freezing**. The condensation point/temperature would be the same as the boiling temperature and the freezing point/temperature would be the same as the melting temperature.

1.3.9 The kinetic theory of matter and the energy changes accompanying changes of state

The changes of state can be explained with reference to the kinetic theory in terms of energy and intermolecular forces between the particles.

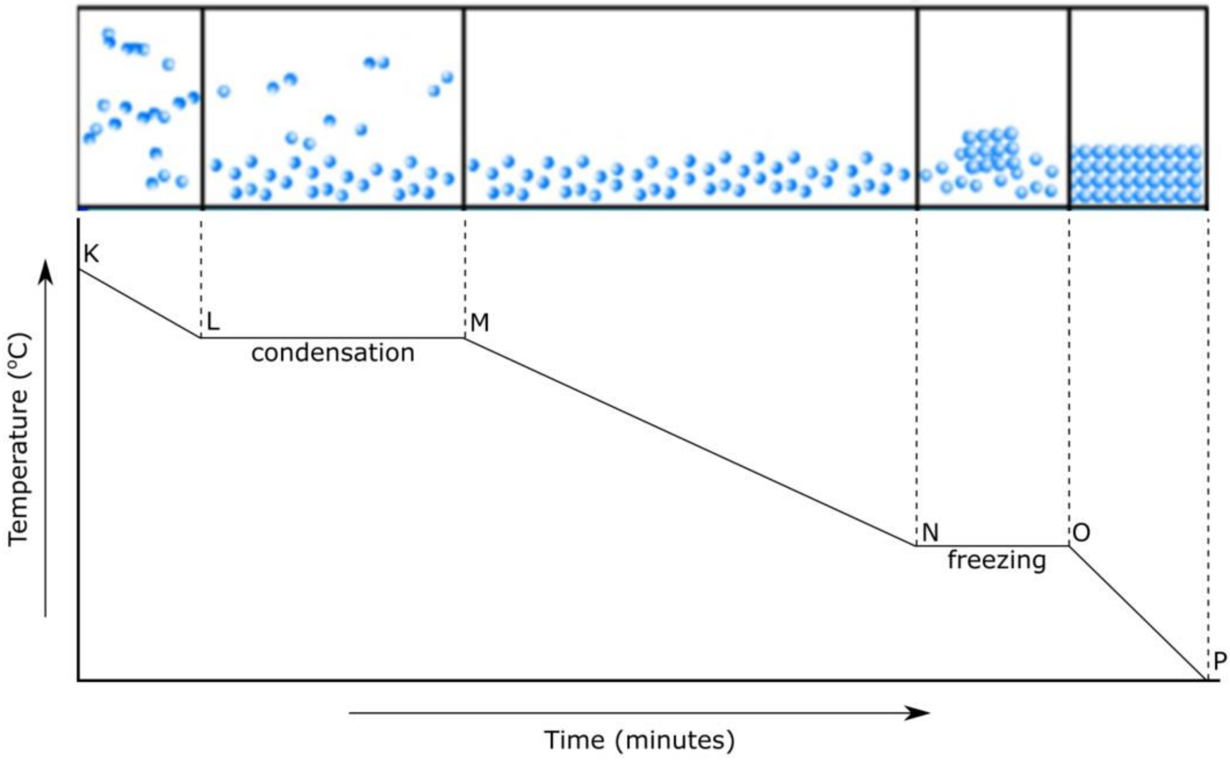
The diagram below shows the behaviour of particles in the different stages of the heating curve of a substance that is being heated.



- **A – B**: The substance exists in a solid state. As the solid is heated its temperature increases. This happens because the particles absorb the heat energy provided, their kinetic energy increases resulting in more particle movement. However, the particles continue to vibrate in their fixed positions.
- **B - C**: During the melting stage the heat energy provided is used to separate the particles by weakening the forces of attraction between them and they can slide over each other forming a liquid. The heat provided is not directed towards increasing the kinetic energy of the particles but to change the state of matter.

- **C – D**: The substance exists in a liquid state. As the liquid is heated its temperature increases. This happens because the particles absorb the heat energy provided, their kinetic energy increases resulting in more particle movement.
- **D – E**: During the boiling stage the heat energy provided is used to overcome the forces of attraction between particles, so they are free and independent from each other as the gas is formed. The heat provided during this stage is not directed to increase the kinetic energy of the particles but to change the state of matter.
- **E – F**: The substance exists as a gas. The particles are moving at a high speed in random motion. As the gas continues to be heated its temperature increases. This happens because the particles absorb the heat energy provided, their kinetic energy increases resulting in faster particle movement.

Cooling curves are the reverse of heating curves. During cooling, the particles of a substance lose their kinetic energy. The movement of the particles decreases and the forces between the particles become stronger as a gaseous substance condenses to a liquid, and then cools further until the liquid freezes into a solid.



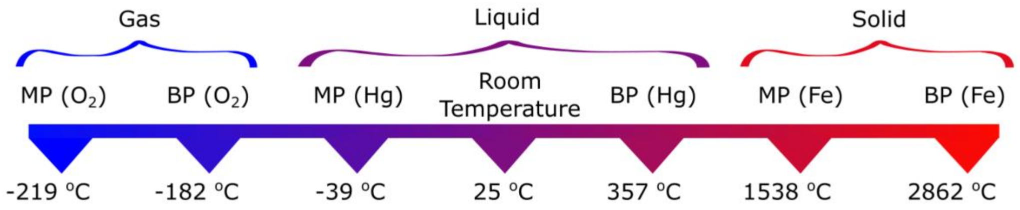
Further reading on the kinetic theory of matter and the energy changes accompanying changes of state.

1.3.10 Predicting the physical state of a substance

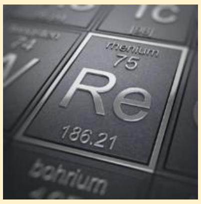
To predict the physical state of a substance at a given temperature, one should compare the given temperature with the melting and boiling points of the substance. The following table summarises how to work this out.

Temperature	State of substance
Given temperature is less than the melting point	Solid
Given temperature is between the melting and the boiling point	Liquid
Given temperature is higher than the boiling point	Gas

Room temperature is taken as 25 °C. At this temperature all elements and compounds exist in a state of matter that is familiar to us. For example oxygen (O_2) is a gas, mercury (Hg) is a liquid while iron (Fe) is a solid. The diagram below shows why these elements are found in these particular states of matter. For a substance to be in the **liquid** state of matter, room temperature must fall between the melting and boiling point of the material, as shown in the diagram below for mercury.



Predicting the state of matter



Do you know that the element with the lowest melting point is Helium, with a boiling point of -272 °C and that the element with the highest boiling point is Rhenium with 5590 °C? This means that at a temperature below -272 °C all elements will be in their solid state. And at a temperature above 5590 °C all elements would have changed to their gaseous state.

However, it is important to understand that these two extreme temperatures are not easily found on Earth. For example, the calculated coldest temperature in our universe is that of -270.14 °C and the calculated temperature of the Sun’s outer shell (photosphere) is that of 5504 °C.

Substances from the Earth

Aquatic Environments

The Atmosphere

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

2 Rates of chemical reactions – LO 11

2.1 Introducing rates of chemical reactions

2.1.1 What is the rate of a chemical reaction?

By the end of this section, you should be able to:

- State that rate of reaction is the increase in amount of product or decrease in amount of reactant with time. (11.2a)

The rate of a chemical reaction is an indication of how rapidly the reactants are changed into products or the products are formed from the reactants.

Some reactions are very slow like the rusting of iron, and some proceed at a moderate rate like glue setting. Others are extremely fast like fireworks.

Fireworks

Rusting

The rate of a chemical reaction can be defined as:

the change in concentration of reactants or products per unit time.

2.1.2 Performing experiments to determine the rates of chemical reactions

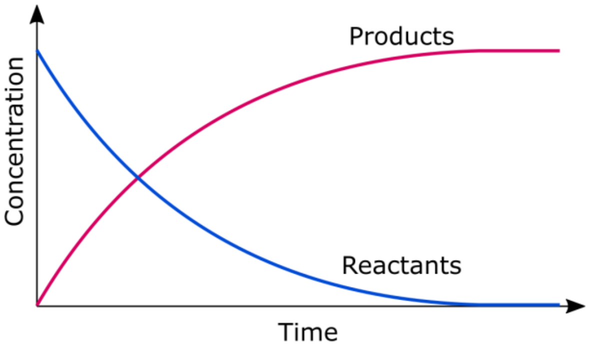
By the end of this section, you should be able to:

- Perform experiments to measure the rate of a reaction. *(E.g. Between an acid and different metals; between limestone and acid; precipitation reactions such as the reaction of thiosulfate with an acid. No chemical equation required for the latter.)* (11.2b)
- Investigate methods to follow the rate of a reaction. *(E.g. Between an acid and different metals; between limestone and acid; precipitation reactions such as the reaction of thiosulfate with an acid.)* (11.3b)

The rate of a reaction can be followed by:

- measuring how quickly the reactants are used up **or**
- measuring how quickly the products are formed.

The sketch shows how the concentration of reactants / products change with time.



The rate of different reactions can be followed by observing certain changes at regular time intervals.

Such observable changes can be:

- the volume of gas produced.
- a change in mass.
- the amount of precipitate produced.

Reaction equations only give the final picture and do not show how the reaction proceeds. In reality a reaction is made up of a series of steps and a reaction is only as fast as the slowest step.

Simple experiments can be carried out in the laboratory to monitor the rate of the reaction. These reactions include:

1. An acid reacting with different metals.
2. An acid reacting with limestone (marble chippings).
3. A precipitation reaction between sodium thiosulfate and an acid.

In experiments where a gas is released, such as in reactions 1 and 2 above, the volume of gas produced or the decrease in mass is measured at regular time intervals.

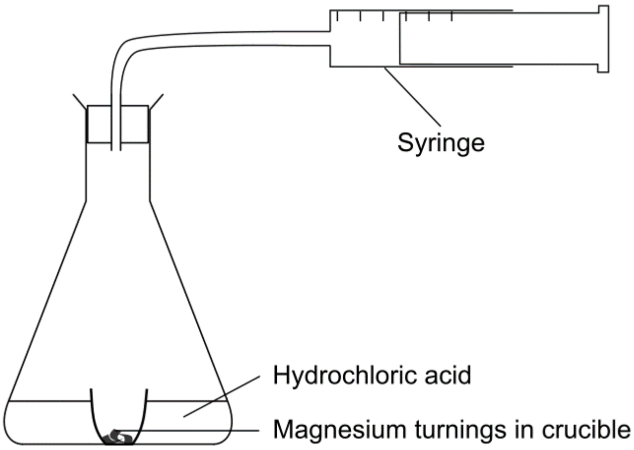
In an experiment where a precipitate is formed, such as in reaction 3, the time taken for a liquid to become 'cloudy' due to the formation of a precipitate is measured.

2.1.2.1 Reactions between an acid and different metals

When a metal reacts with an acid, salt and hydrogen are produced.

Metal + Acid → Salt + Hydrogen

It is very practical to monitor the rate of a reaction by recording the volume of gas given off at regular time intervals. The following apparatus can be set up.



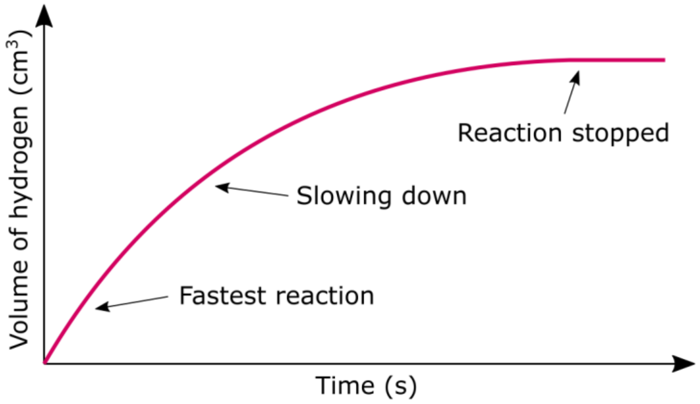
1. The reaction of the acid with the metal starts when the flask is shaken/swirled. At the same time the stopwatch is started.

2. The gas given off is collected in the gas syringe.

3. The volume of hydrogen liberated is recorded every 30 s interval until the reaction stops. Results are recorded in the table below.

Volume of hydrogen (cm^3)											
Time (s)	0	30	60	90	120	150	180	210	240	270	300

A graph is plotted with volume of hydrogen (cm^3) on the y-axis and time (s) is plotted on the x-axis.



The rate of the reaction at any one time is given by the slope or gradient of the graph.

- When the slope is steep, a lot of hydrogen is given off and the reaction is fast.
- When the slope is less steep, less hydrogen is being produced, showing that the reaction is slowing down due to the consumption of reagents.
- When the graph is horizontal, it means that the reaction has stopped, and no more hydrogen is being produced.

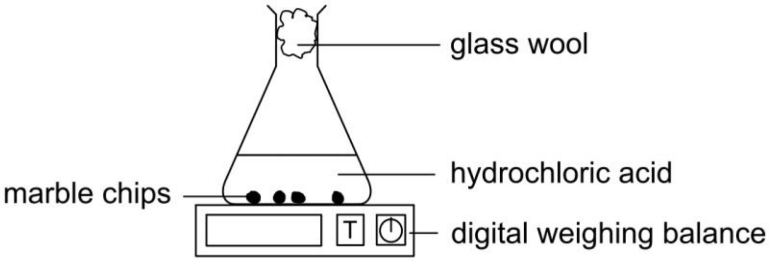
2.1.2.2 Reaction between limestone and acid

The reaction between limestone and acid will produce salt, water, and carbon dioxide gas.

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

The reaction rate can be monitored by recording the volume of carbon dioxide given off at regular time intervals, just like the reaction between metal and acid.

Another possible way to record the reaction rate is by monitoring the change in mass of the contents in the reaction vessel (conical flask) at regular time intervals, as carbon dioxide is allowed to diffuse out into the atmosphere. The following apparatus is set up:



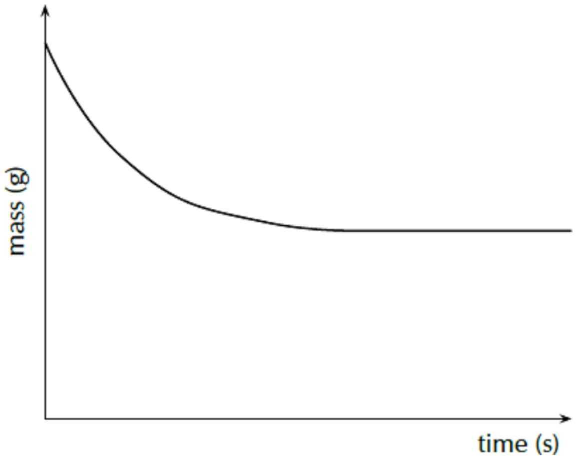
Method:

1. A clean empty conical flask is placed on a top pan balance.
2. A known volume of acid is transferred into a conical flask.
3. A known mass of limestone chunks (marble chippings) are added.
4. A piece of glass wool is placed in the neck of the conical flask to allow carbon dioxide to escape whilst preventing loss of acid spray.
5. The initial mass of the conical flask, limestone, acid, and glass wool are recorded.

6. The conical flask is swirled from time to time, and mass is recorded every 30 seconds until reaction stops.
7. Results are recorded in the table below.

Mass of flask and contents (g)											
Time (s)	0	30	60	90	120	150	180	210	240	270	300

The readings obtained can be plotted on a graph paper, with mass of flask and contents (g) on the y-axis versus time (s) on the x-axis. The graph shows that the reaction is fastest at the beginning. It then slows down until it eventually stops when one or both reagents are used up.

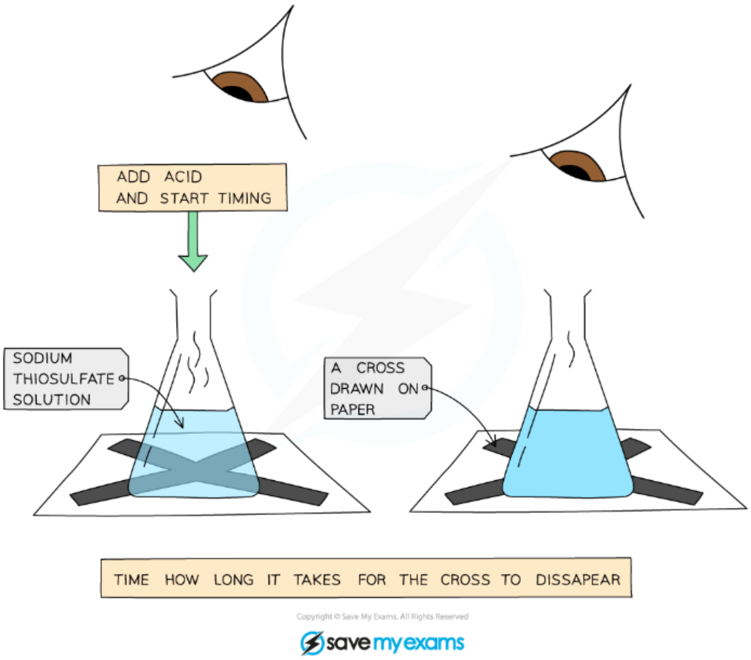


2.1.2.3 Reaction between sodium thiosulfate and acid

The reaction between sodium thiosulfate and acid, produces solid sulfur as a precipitate. Both reactants are colourless solutions.

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow S(s) + 2NaCl(s) + SO_2(g) + H_2O(l)$$

The sulfur forms a cloudy yellow-white precipitate during the reaction. A paper marked with a cross is placed underneath the conical flask as shown in the setup. When the precipitate deposits the solution becomes cloudy or opaque and the cross is no longer visible. The time for the cross to disappear is noted.



Method:

1. Using a measuring cylinder, add 50 cm^3 of dilute sodium thiosulfate solution to a conical flask.
2. Place the conical flask on a piece of paper with a black cross drawn on it.
3. Using a different measuring cylinder, add 10 cm^3 of dilute

hydrochloric acid to the conical flask. Immediately swirl the flask to mix its contents and start a stop clock.

4. Look down through the reaction mixture. When the cross is no longer visible, record the time.

A simulation on investigating the effect of changing the sodium thiosulfate concentration on its reaction with hydrochloric acid.

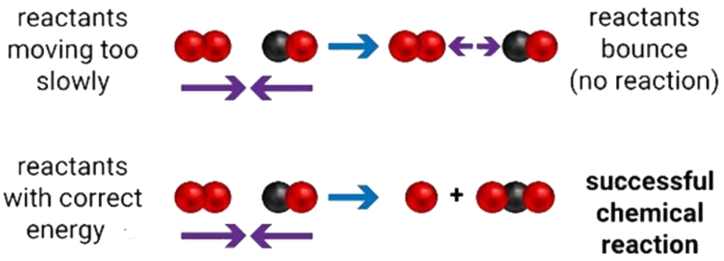
2.1.3 The Collision Theory

By the end of this section, you should be able to:

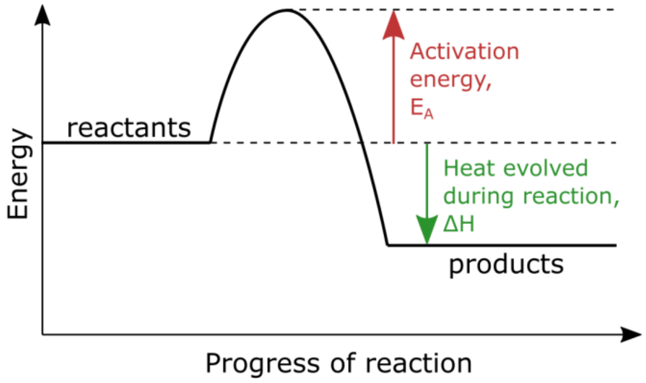
- Use the kinetic and collision theories to explain how factors such as state of subdivision, concentration, temperature, and pressure affect the rate of a reaction. (11.3g)

Reaction rates are explained by the collision theory.

The particles have to **collide** with **enough energy** to break the existing chemical bonds and reform new bonds.



The minimum energy needed for the bonds to be broken is called the **activation energy, E_A**. Collisions that have the required quantity of activation energy and that result in a chemical reaction are called **successful collisions**.



The **rate of reaction** depends on the **number of successful collisions per second**. Note that once a successful collision happens and a reaction occurs, the amount of activation energy provided will be given back for an exothermic reaction as shown in the diagram above.

2.2 The factors that affect rates of reaction

By the end of this section, you should be able to:

- Identify conditions that may affect the rate of a given reaction. *(Limited to state of subdivision of reactants, and temperature.)* (11.1c)
- Identify conditions that may affect the rate of a given reaction. *(Limited to concentration, catalyst, light, and pressure in gases.)* (11.2c)
- Investigate how the rate of reaction may be affected by surface area of reactants/catalysts. (11.2d)
- Investigate how the rate of reaction may be affected by various factors. *(E.g. Surface area of reactants, concentration of reactants, temperature, light, and the use of a catalyst.)* (11.3d)
- Plot a single series of data using experimental results. (11.2e)
- Plot multiple series of data using experimental results. (11.3e)
- Interpret results/graph containing single series of data related to rates of reactions. (11.2f)
- Interpret results/graphs containing multiple series of data related to rates of reaction. (11.3f)

The factors that affect the rates of chemical reactions are:

- State of subdivision (surface area) of the solid reactant
- Temperature of the reactants
- Concentration of the reactants
- Pressure (in reactions involving gases)
- Catalyst
- Light

2.2.1 Surface area

The following two experiments can be used to investigate the effect surface area has on the rate of reaction when at least one of the reactants is in the solid state.

2.2.1.1 Experiment 1

Let's consider the reaction between magnesium with hydrochloric acid. The experiment is first performed using magnesium turnings and then the experiment is repeated using magnesium powder. The rate of each reaction can be monitored by measuring the rate of increase in volume of hydrogen gas produced.

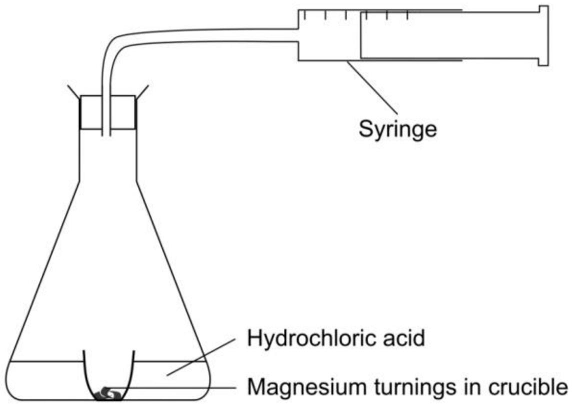
$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Aim: To investigate how the surface area of magnesium affects the rate of reaction when it reacts with an acid.

Variables:

Independent variable	Surface area
Dependent variable	Volume of hydrogen (cm^3)
Controlled variables	Volume and concentration of acid Mass of magnesium Temperature

Method:



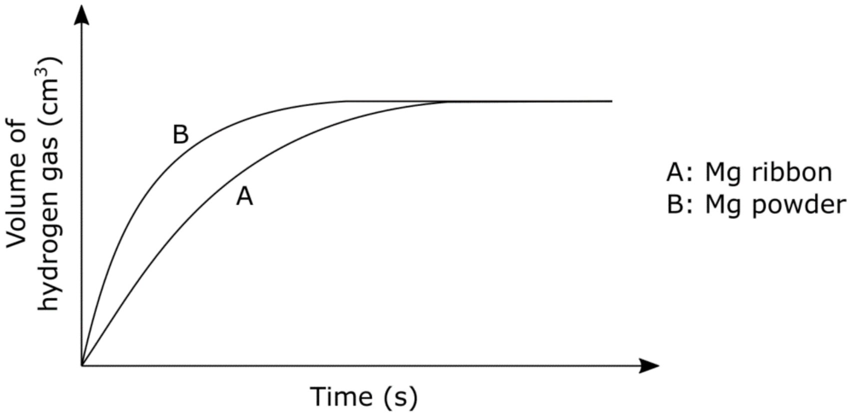
1. 0.1 g of magnesium turnings are placed in a crucible which is lowered into a conical flask filled with 25 cm^3 of 0.5 mol dm^{-3} hydrochloric acid as shown in the diagram.
2. The reaction is started by swirling the conical flask to mix the magnesium with the hydrochloric acid.
3. The volume of hydrogen liberated is recorded at 30 s intervals until three unchanged readings are obtained.
4. The experiment is repeated using the same mass of magnesium powder added to 25 cm^3 of 0.5 mol dm^{-3} hydrochloric acid. The volume of hydrogen liberated is again recorded every 30 s interval until three unchanged readings are obtained.
5. The results of both experiments are plotted on the same graph of volume of hydrogen (cm^3) liberated against time (s).

Results:

Time (s)	Volume of gas collected (cm^3)	
	Magnesium ribbon	Magnesium powder

Graphs:

When the results obtained are plotted, the following graph should be obtained.



Conclusion:

From the graph it can be concluded that the reaction is faster between magnesium powder and hydrochloric acid since the graph is steeper. Therefore, the larger the surface area the faster the rate of reaction.

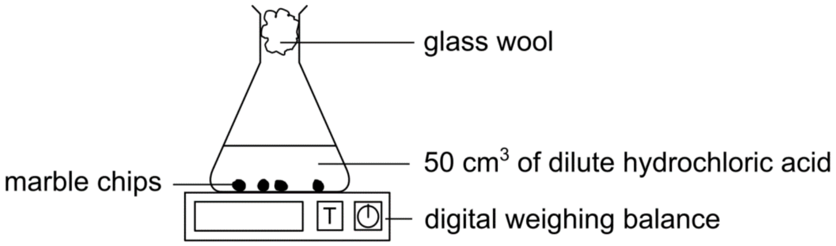
Both experiments produce the same volume of hydrogen gas since the mass of magnesium and concentration and volume of acid were not changed.

2.2.1.2 Experiment 2

An alternative experiment which could be carried out is the reaction between marble chips and hydrochloric acid. In this case, the change in mass of the reaction vessel is monitored at regular time intervals.

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

Aim: To investigate how the surface area of marble chips affects the rate of reaction when it reacts with an acid.



Variables:

Independent variable	Surface area
Dependent variable	Mass of contents and flask (g)
Controlled variables	Volume and concentration of acid Mass of calcium carbonate Temperature

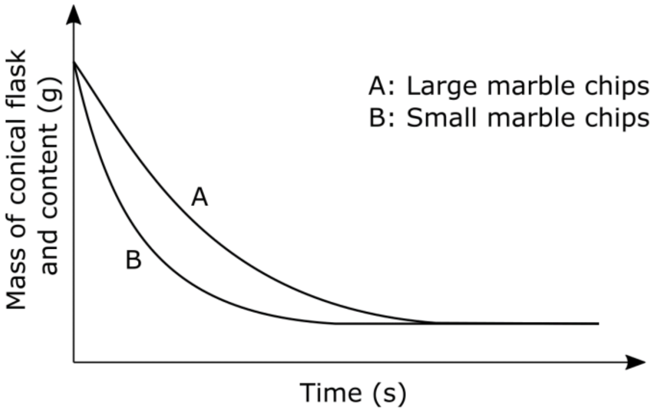
Method:

1. 50 cm^3 of 1 mol dm^{-3} hydrochloric acid are poured into a conical flask and placed on the electronic balance.
2. 1 g of large marble chips is added.
3. The neck of the flask is plugged with a piece of glass wool and the initial mass is recorded.
4. The mass readings are noted every 30 seconds until three unchanged readings are obtained.
5. The experiment then is repeated with 1 g of small marble chips which are reacted with 50 cm^3 of 1 mol dm^{-3} hydrochloric acid. The mass readings are again noted every 30 seconds until three unchanged readings are obtained.
6. The results of both experiments are plotted on the same graph of mass of contents and flask (g) against time (s).

Results:

Time (s)	Mass of contents and flask (g)	
	Large marble chips	Small marble chips

Graph



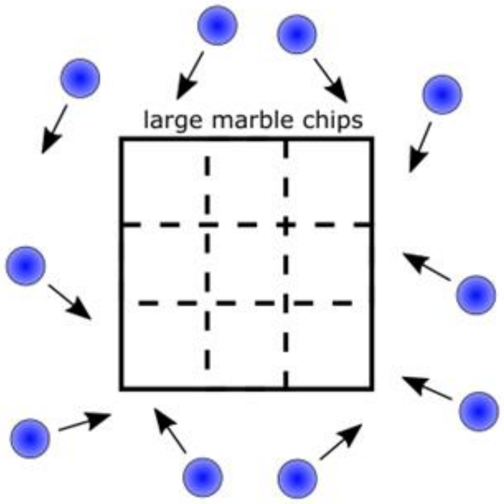
Conclusion:

From the graph it can be concluded that the steeper the slope, the faster the rate of reaction. Therefore, the larger the surface area the faster the rate of reaction.

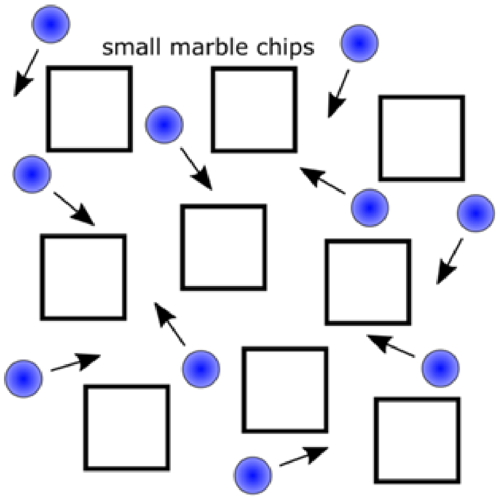
Both curves level at the same point as the same mass of marble chips, concentration and volume of acid was used in both experiments. The final mass of the flask and contents is the same in each, as the same mass of carbon dioxide was liberated.

Explaining the effect of surface area on rate of reaction in terms of the collision theory

Reactions can only happen when particles collide with enough energy.



Particles in the acid collide with the outer surface of the solid reagent (magnesium ribbon or large marble chips). Rate of reaction is slow due to less successful collisions.



When a crushed/ powdered solid reagent is used (e.g., powdered magnesium or crushed marble chips) the reaction is faster because more surface area of reagent is in contact with the acid. This results in more successful collisions.

2.2.2 Temperature

Temperature is a measure of the kinetic energy that the particles in a substance have. The greater the temperature, the greater is the value of the kinetic energy of the particles in a substance.

Since a specific quantity of kinetic energy is needed for a fruitful (successful) collision to happen, then temperature is a very important factor in determining whether a chemical reaction will happen or not. Consequently, temperature affects the rate of a chemical reaction.

At low temperatures particles move relatively slowly as they have a low kinetic energy. Also, few collisions between the particles lead to the formation of the product at a slow rate of reaction.

As the temperature increases the particles move rapidly as the kinetic energy of the particles increases. The frequency and energy of collisions increase resulting in a faster rate of reaction.

Experiment: To find how temperature affects the rate of a reaction

The reaction which will be used to investigate the effect of temperature on the rate of reaction is the reaction between sodium thiosulfate with hydrochloric acid. This reaction produces a yellow precipitate of sulfur.

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$$

The rate of this reaction is measured by looking at the rate at which the product (sulfur) is formed. The solid sulfur makes the colourless solution go cloudy.

This reaction is usually carried out in a flask placed on a piece of white paper with a black cross on it.

The time taken for the cross to disappear is used as an indication for the rate of reaction.

Variables:

Independent variable	Temperature
Dependent variable	Time for the cross to disappear
Controlled variables	Volume and concentration of acid Volume and concentration of thiosulfate

Method:

1. 10 cm^3 of 0.5 mol dm^{-3} sodium thiosulfate solution was measured and placed into a conical flask.
2. 40 cm^3 distilled water was added.
3. 50 cm^3 of 0.2 mol dm^{-3} hydrochloric acid was measured in a second conical flask.
4. The temperature of both flasks was maintained at 30 °C using a water bath.

5. A conical flask was prepared and when the two solutions were at the same temperature, they were mixed and placed on a paper marked with a cross.

6. A stopwatch was started immediately, and the time taken for the cross to disappear was recorded.

7. The experiment was repeated several times such that more readings were taken over a temperature range of 25 – 80 °C.

8. A table was used to record the results obtained.

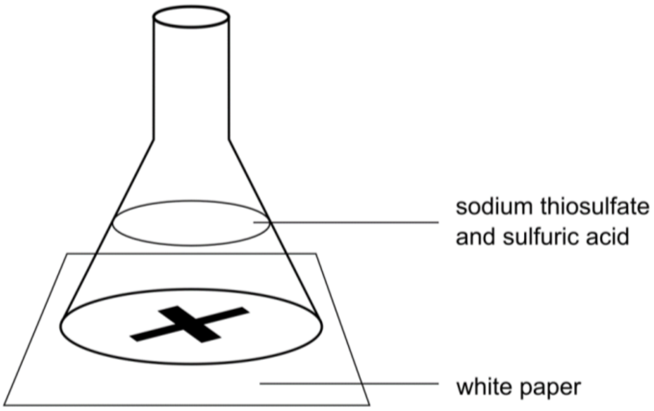


Table of Results:

Temperature (°C)	Time taken for cross to disappear (s)

Results are plotted on a graph of **time** for cross to disappear (seconds) on the y-axis against **temperature** (°C) on the x-axis.

Plotting a single series of data from the table of results (11.2e)

Use a graph paper to plot a line graph to represent the results obtained below.

Temp (°C)	Time taken for cross to disappear (s)
10	195
20	107
30	61
40	35
50	17
60	9

Interpretation of results:

As temperature increases the time taken for the cross to disappear becomes shorter indicating a faster reaction. Usually, for every 10 °C rise in temperature, the rate of reaction approximately doubles, for example if a reaction lasts 200 s at 20 °C, at 30 °C, it will last 100 s.

Precautions taken in this experiment:

- Same acid and sodium thiosulfate concentration is used in all experiments.
- The same person will determine the time taken for the cross to disappear.

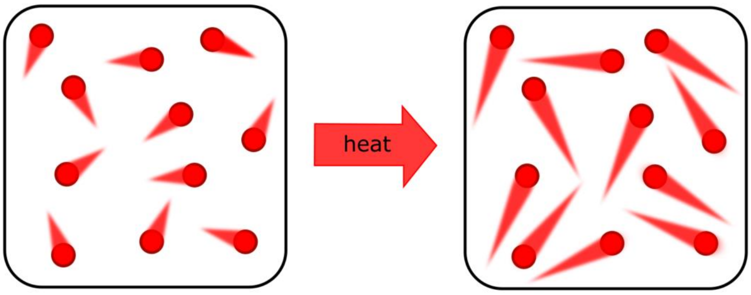
Safety considerations:

- Perform experiment in a well-ventilated area due to production of sulfur dioxide. The sulfur dioxide can be neutralised after the experiment is over by pouring the contents of the flask into a solution of sodium carbonate.

Collision Theory: temperature and rate of reaction

There are two ways we can explain why increasing the temperature increases the rate of reaction.

1. An increase in temperature will increase the kinetic energy in particles leading to more frequent collisions. This results in an increased rate of reaction.



low temperature

high temperature

2. Temperature also increases the energy at which particles collide. To react, particles must collide with a minimum amount of energy. This is called activation energy. High temperatures will increase the chances of particles colliding with the right amount of energy. The number of successful collisions in a given time increases as the temperature increases.

Collision theory and temperature

Two main founders of the collision theory were Max Trautz (a German chemist) and William Lewis (a British physical chemist) who published their work between 1916-1918.

The interesting thing is that since their work emerged during World War I, both chemists were unaware of each other’s publications. Trautz worked mostly on areas related to chemical kinetics, and he focused his interest on the activation energy of molecules.

Max Trautz: William Lewis:

2.2.3 Concentration

The reaction between magnesium and hydrochloric acid is a useful reaction to monitor the rate of a reaction. The volume of hydrogen produced over a given time is used as an indication of the rate of the reaction.

When investigating the concentration on the rate of a reaction, the temperature, mass, and size of magnesium are kept the same. The volume of hydrochloric acid used is kept the same and is present in excess so that the quantity of magnesium limits the reaction.

The only variable which will be altered is the concentration of the acid used.

As the concentration of the acid increases, the rate of reaction increases.

Investigating how the rate of reaction may be affected by concentration of hydrochloric acid

Variables:

Independent variable	Concentration of acid
Dependent variable	Volume of hydrogen (cm^3)
Controlled variables	Temperature Mass and surface area of magnesium Volume of acid

Method:

1. 0.1 g magnesium ribbon is placed in a conical flask with side arm connected to a syringe. This quantity of magnesium is used so as not to exceed the volume that can be collected by the syringe.
2. 20 cm^3 of 0.5 mol dm^{-3} hydrochloric acid is added to the magnesium ribbon. The volume of hydrogen given off is then recorded every 30 seconds until three unchanged readings are obtained.
3. The experiment is then repeated using 20 cm^3 of 1 mol dm^{-3} hydrochloric acid and the same mass of magnesium ribbon.
4. The results are recorded in a table.
5. A graph of **volume of hydrogen** (in cm^3) against **time** (in seconds) is plotted.

Table of Results:

Volume of hydrogen (cm^3)	Time (seconds)

Results may be presented in the form of a graph of volume of hydrogen produced (cm^3) against time (seconds).

Interpretation of results:

The higher the concentration of the acid, the faster the rate of reaction.

The final volume of hydrogen released is the same in each experiment as mass of magnesium ribbon used is always in the same.

Collision Theory: concentration and rate of reaction

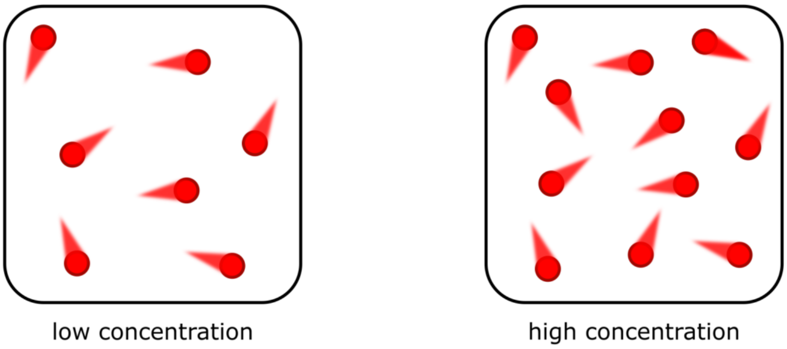
Concentration is equal to the amount of reactant particles in a given volume of water.

A concentrated solution has more particles of solute per unit volume than a dilute solution.

Rate of reaction depends on:

- Number of successful collisions per second
- Energy in the collision (activation energy)

The higher the concentration the more particles present. As a result, there is a higher chance of particles colliding with other reactant particles with enough energy.



If there are more collisions per second, the rate of reaction is faster.

Other reactions which could be used to investigate how concentration affects the rate of reaction include:

- Sodium thiosulfate with hydrochloric acid (changing the concentration of thiosulfate).
- Calcium carbonate and hydrochloric acid (change acid concentration).

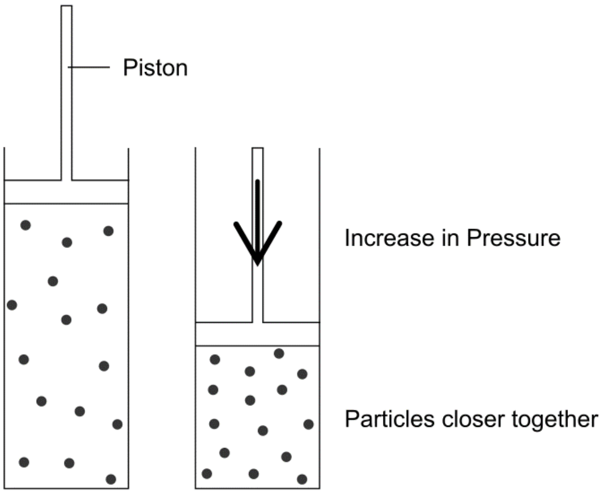
2.2.4 Pressure

The rate of reaction involving gases as the reactants, will be affected by a change in pressure. This ONLY applies when gases are involved.

A higher pressure will increase the rate of a reaction, while a lower pressure will decrease it.

Reaction rates for solids and liquids remain unaffected by increases in pressure.

Collision Theory: pressure and rate of gaseous reactions



The pressure generated by a gas in an enclosed container is the result of the collisions of the gas particles with the wall of the container.

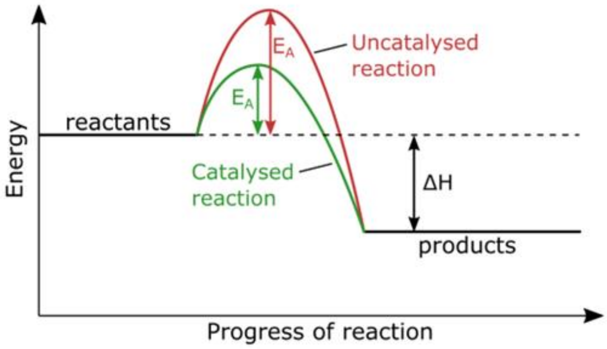
Therefore, for a given mass of a gas, the way to increase its pressure is to compress it into a smaller volume. Since the number of particles per unit volume is higher, this leads to a higher number of collisions with the wall of its container thereby increasing pressure.

The higher the pressure, the higher the concentration of the reactant gases in a given volume. This results in an increase in successful collisions between gaseous reactants, hence a higher rate of reaction follows.

2.2.5 Catalysts

A **catalyst** is a substance that alters the rate of a chemical reaction but remains chemically unchanged.

- Catalysts are usually transition metals and/or their compounds.
- Catalysts are used in small amounts.
- Catalysts act by providing an alternative chemical pathway with a lower activation energy.
- Since catalysts are usually in the solid state, the higher the surface area of the catalyst the more effective it becomes.
- The mass of the catalyst does not change and so is reused indefinitely.



Investigating the effect of different catalysts on the rate of a reaction

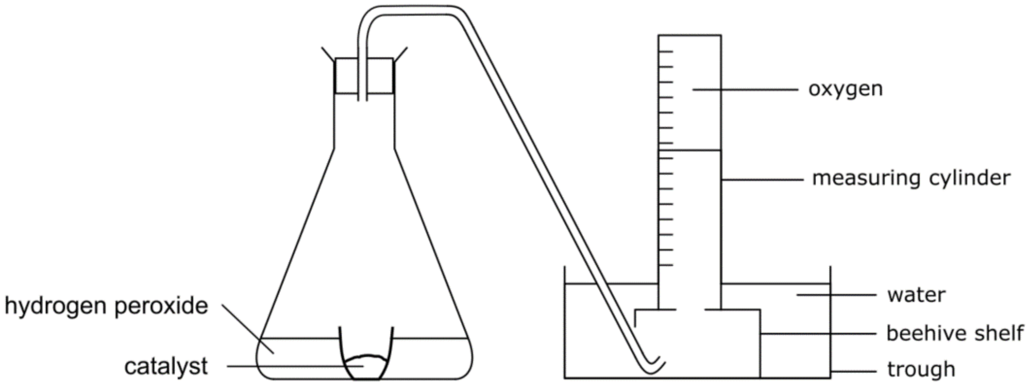
Hydrogen peroxide solution decomposes slowly at room temperature to oxygen and water.

hydrogen peroxide → oxygen + water

$$2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$$

The volume of oxygen given off can be measured using a gas syringe, or by collecting in an inverted measuring cylinder. The effectiveness of the catalyst can be judged from the initial rate of the oxygen production.

Setup of apparatus:



Variables:

Independent variable	Type of catalyst
Dependent variable	Volume of oxygen (cm^3)
Controlled variables	Temperature Mass and surface area of catalyst Volume and concentration of hydrogen peroxide

Method:

1. Set up the apparatus as shown in the diagram.
2. Measure 0.5 g of a catalyst and place it in a crucible.
3. Using a measuring cylinder, add 50 cm^3 of 6% hydrogen peroxide solution (20 volumes/pharmacy grade) to a conical flask.
4. Lower the crucible including the 0.5 g of a catalyst into the flask such that the crucible floats on the hydrogen peroxide.
5. The reaction starts when the reaction vessel is shaken to allow the catalyst to mix with the hydrogen peroxide and the stopwatch is switched on simultaneously.
6. Record the volume of gas given off every 10 seconds. Continue timing until no more oxygen appears to be given off.
7. Clean the apparatus.
8. Repeat steps 1 to 7 for another two catalysts keeping all other variables unchanged.

Results:

Catalyst	Time (s)					
	10	20	30	40	50	60

Plotting graphs using experimental data (11.3f)

Use a graph paper to plot line graphs to represent the results obtained.

Catalyst	Time (s)					
	10	20	30	40	50	60
Manganese(IV) oxide	59	81	92	100	100	100
Iron	34	48	54	57	59	60
Copper(II) oxide	13	20	25	28	30	31

2.2.6 Light

Some reactions are sensitive to light and are called **photochemical reactions**. The brighter the light the faster the reaction. Examples of photochemical reactions:

Photography

A silver halide salt (e.g. silver chloride) is used in black and white photography.

Silver chloride is sensitive to light and breaks down to form metallic silver, which appears black. The brighter the light falling on the photographic film or paper, the faster the reaction and so show as darker parts on the negative.

Negative

Actual photograph

Adapted from: Remembering the Dead: Discovering Century-Old Dry Plate Photos | PetaPixel

Photosynthesis

An important chemical reaction in plant cells, whereby light energy is trapped and changed into chemical energy for the plant cells. Light energy is absorbed by the green pigment, chlorophyll.

Further reading on factors affecting rate of reaction.

Catalyst and Light on rate of reactions

In 1794, an English chemist, Elizabeth Fulhame invented the concept of catalysis and photoreduction (light dependant reactions). This was done after she conducted numerous experiments, spanning over a period of 14 years, involving oxidation-reduction reactions.

Her interest started with investigating ways of using heavy metals to stain cloth, under the influence of light. She used various metal salts which included salts of gold, silver, platinum, mercury, copper, and tin. Her work was considered by some the starting point to photography.

Watch lesson 1, part 1 by Mr D. Bugeja on rates of reaction.

Watch lesson 1, part 2 by Mr D. Bugeja on rates of reaction.

Watch lesson 2, part 1 by Mr D. Bugeja on rates of reaction.

Watch lesson 2, part 2 by Mr D. Bugeja on rates of reaction.

Substances from the Earth

Aquatic Environments

The Land

Making New Materials: How fast? How far? How much?

Carbon Compounds. Meeting our energy needs

The Atmosphere

3 Quantitative calculations – LO 10

3.1 Calculations related to relative mass and percentage composition by mass.

By the end of this section, you should be able to:

- Calculate relative formula mass or relative molecular mass of a compound from relative atomic masses. (10.2a)
- Work out percentage by mass calculations. *(E.g. Percentage by mass of an element in a compound and the value of xH_2O in a hydrated compound.)* (10.2b)

3.1.1 Calculating RMM/RFM for compounds

When referring to relative masses for substances, the term relative molecular mass (RMM) is used for covalent molecules (e.g. H_2, O_2, NH_3) while the term relative formula mass (RFM) is used for ionic compounds (e.g. NaCl, $MgCl_2$, $CaCO_3$).

RMM and RFM can be obtained by adding the relative atomic masses of the atoms in the molecule or formula. As with RAM, RMM and RFM **do not** have units.

The **relative molecular mass (RMM)** of an element or a compound is the mass of one molecule of that element or compound compared with $\frac{1}{12}$th the mass of an atom of carbon (^{12}C).

The **relative formula mass (RFM)** of an ionic compound is the mass of one formula of that compound compared with $\frac{1}{12}$th the mass of an atom of carbon (^{12}C).

Worked examples.

Substance	Formula	RAM	RMM / RFM
nitrogen	N_2	N = 14	14 × 2 = **28**
ammonia	NH_3	N = 14 H = 1	14 + (1 × 3) = **17**
magnesium carbonate	$MgCO_3$	Mg = 24 C = 12 O = 16	24 + 12 + (16 × 3) = **84**
calcium hydroxide	$Ca(OH)_2$	Ca = 40 O = 16 H = 1	40 + 2(16 + 1) = **74**
hydrated copper(II) sulfate	$CuSO_4.5H_2O$	Cu = 63.5 S = 32 O = 16 H = 1	63.5 + 32 + (16 × 4) + 5[(1 × 2) + 16] = **249.5**

1. Calculate the relative molecular mass of:
    a. O_2
    b. NO_2
    c. C_2H_5OH
    d. CH_3COOH
2. Calculate the relative formula mass of:
    a. $MgCl_2$
    b. $(NH_4)_2SO_4$
    c. $Ca(HCO_3)_2$
    d. $FeSO_4.7H_2O$

3.1.2 Percentage composition by mass

3.1.2.1 Percentage composition by mass of an element in a compound

Percentage composition by mass is the percentage by mass of **each element** in a compound.

In order to calculate the percentage by mass of an element in a compound:

1. Write down the formula of the compound.
2. Calculate the RMM / RFM of the compound using the list of RAMs.
3. Write the mass of each element in the compound as a fraction of the total mass.
4. Multiply the fraction by 100 to give a percentage.

Worked examples

1. Calculate the percentage composition of **sulfur** and **oxygen** in sulfur dioxide, SO_2. *(RAM of S = 32; O = 16)*

    RMM of SO_2: $32 + (16 \times 2) = 64$

    Sulfur in SO_2: 32

    Oxygen in SO_2: $16 \times 2 = 32$

    Therefore, % composition of sulfur = $\frac{32 \times 100}{64}$ = **50%**

    and % composition of oxygen = $\frac{32 \times 100}{64}$ = **50%**

    NOTE: The two percentages add up to 100%.

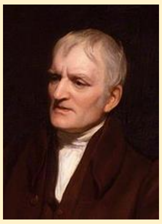
2. Calculate the percentage composition of **nitrogen** in ammonium nitrate, NH_4N0_3. *(RAM of N = 14; H = 1; O = 16)*

    RMM of NH_4N0_3 : $14 + (1 \times 4) + 14 + (16 \times 3) = 80$

    Nitrogen in NH_4N0_3 : $14 \times 2 = 28$

    Therefore, % composition of nitrogen = $\frac{28 \times 100}{80}$ = **35%**

Of elements from compounds



The first table of relative atomic weights was presented by John Dalton (as English chemist, in 1803), and it had only 6 elements. In this table Dalton related oxygen, nitrogen, carbon, sulfur, and phosphorus to hydrogen, which was given a conventional mass of 1. In his presentation Dalton did not explain how he came up with the relative masses of these elements.

However, later in one of his notebooks, it was discovered that these relative atomic masses were obtained from his analysis of work published by other chemists, whose work focused on compounds such as: water, carbon dioxide, and ammonia.

3.1.2.2 Percentage composition by mass of the water of crystallisation in a hydrated compound

In order to calculate the percentage composition of water of crystallisation in hydrated salts, the total relative mass of the water molecules present must be taken as multiples of 18 depending on the formula.

Worked example

Calculate the percentage composition of **water** in hydrated magnesium sulfate.

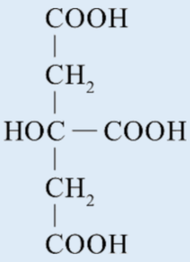
$MgSO_4.7H_2O$. (RAM of Mg = 24; S = 32; O = 16; H = 1)

RMM of $MgSO_4.7H_2O$: $24 + 32 + (16 \times 4) + 7[(1 \times 2) + 16] = 246$

Water in $MgSO_4.7H_2O$: $7[(1 \times 2) + 16] = 126$

Therefore, % composition of water: $\frac{126 \times 100}{246} = \mathbf{51.2\%}$

1. Calculate the percentage by mass of iron in iron(III) oxide, Fe_2O_3.
2. Work out the percentage composition of water in hydrated calcium chloride $CaCl_2.2H_2O$
3. Citric acid is found in citrus fruit and has the following formula.



   a. Write the chemical formula of citric acid.
   b. Calculate the RMM of citric acid.
   c. Calculate the percentage by mass of oxygen in citric acid.

3.2 Calculations to determine the amount of reacting substances

By the end of this section, you should be able to:

- Perform an experiment to determine the empirical formula of a substance. *(Limited to binary compounds and finding the value of xH_2O in a hydrated compound.)* (10.3c)
- Calculate the formula of reacting masses from experiment and relate empirical and molecular formulae of simple substances. (10.3d)
- Calculate the amount of products formed from given amount of one reactant in a reaction and vice versa. *(In moles, number of particles, masses, and volumes of gases at STP. Concept of limiting reagent will not be assessed. Use of Avogadro's constant and Avogadro's law.)* (10.3e)
- Calculate the theoretical and percentage yield of product for a given reaction. (10.3f)

3.2.1 The mole

The mole represents a specific number of particles (atoms, molecules, or ions). This number is given by **Avogadro's constant** that is **6.02×10^{23}** and is used as the main unit to measure quantities of substances. Therefore,

one mole of **any** substance contains 6.02×10^{23} particles.

The **molar mass** of a substance is equal to
the relative atomic/molecular/formula mass of that substance in grams.

For example:

- the RAM of **carbon** (C) is 12. The number of **atoms** that are present in 12 g of carbon is 6.02×10^{23}.
- the RMM of **oxygen** (O_2) is 32. The number of **molecules** that are present in 32 g of O_2 is 6.02×10^{23}.
- the RMM of **water** (H_2O) is 18. The number of **molecules** that are present in 18 g of H_2O is 6.02×10^{23}.
- the RFM of **sodium chloride** (NaCl) is 58.5. The number of **NaCl** units that are present in 58.5 g of NaCl is 6.02×10^{23}. Since NaCl is made up of Na^+ and Cl^- ions, there are 6.02×10^{23} Na^+ ions and 6.02×10^{23} Cl^- ions in one mole of NaCl.
- the RFM of **magnesium chloride** ($MgCl_2$) is 95. The number of **$MgCl_2$** units that are present in 95 g of $MgCl_2$ is 6.02×10^{23}. Since $MgCl_2$ is made up of Mg^{2+} and twice as many Cl^- ions, there are 6.02×10^{23} Mg^{2+} ions and $2 \times (6.02 \times 10^{23})$ Cl^- ions in one mole of $MgCl_2$.

Simple calculations on moles

1. Calculate the mass of 0.5 moles of lithium atoms. (RAM of Li = 7)

1 mole of lithium (Li) weighs 7 g
0.5 moles of lithium (Li) weigh ?

(0.5 moles × 7 g) / 1 mole = **3.5 g**

2. Calculate the mass of 0.5 moles of nitrogen atoms. (RAM of N = 14)

1 mole of nitrogen atoms (N) weighs 14 g

0.5 moles of nitrogen atoms (N) weigh ?

(0.5 moles × 14 g) / 1 mole = **7 g**

3. Calculate the mass of 0.5 moles of nitrogen molecules. (RAM of N = 14)

(First find the RMM of a nitrogen molecule, N_2)

RMM of N_2 : 14 × 2 = 28

Therefore, 1 mole of nitrogen molecules (N_2) weighs 28 g

0.5 moles of nitrogen molecules (N_2) weigh ?

(0.5 moles × 28 g) / 1 mole = **14 g**

4. Calculate the mass of 0.2 moles of calcium carbonate, $CaCO_3$.
(RAM of Ca = 40; C = 12; O = 16)

(First find the RFM of calcium carbonate, $CaCO_3$)

RFM of $CaCO_3$: 40 + 12 + (16 × 3) = 100

Therefore, 1 mole of calcium carbonate ($CaCO_3$) weighs 100 g

0.2 moles of calcium carbonate ($CaCO_3$) weigh ?

(0.2 moles × 100 g) / 1 mole = **20 g**

5. How many moles of oxygen molecules are there in 64 g of oxygen, O_2 ? (RAM of O = 16)

(First find the RMM of an oxygen molecule, O_2)

RMM of O_2 : 16 × 2 = 32

Therefore, in 32 g of oxygen molecules (O_2) there is 1 mole

in 64 g of oxygen molecules (O_2) there is ?

(64 g × 1 mole 32 g) = **2 mol**

6. How many moles are there in 54 g of water, H_2O? (RAM of H = 1; O = 16)

(First find the RMM of a water, H_2O)

RMM of H_2O : (1 × 2) + 16 = 18

Therefore, in 18 g of water (H_2O) there is 1 mole

in 54 g of water (H_2O) there is ?

(54 g × 1 mole) / 18 g = **3 mol**

7. Given that Avogadro's constant is 6.02×10^{23}, calculate the number of atoms in 2.8 g of silicon. (RAM of Si = 28)

In 28 g of silicon (Si) is 1 mole

in 2.8 g of silicon is ?

(2.8 g × 1 mole) / 28 g = **0.1 mol**

Therefore, if 1 mole of silicon contains 6.02×10^{23} atoms,

0.1 moles of silicon contain ? atoms

(0.1 moles × 6.02×10^{23} atoms) / 1 mole = **6.02×10^{22} atoms**

8. Given that Avogadro's constant is 6.02×10^{23}, calculate the number of ions in 28 g of copper(II) hydroxide, $Cu(OH)_2$. (RAM of Cu = 63.5, O = 16, H = 1)

RFM ($Cu(OH)_2$): 63.5 + 2(16 + 1) = 97.5

In 97.5 g of $Cu(OH)_2$ there is 1 mole

in 28 g of $Cu(OH)_2$ there is ?

(28 g × 1 mole) / 97.5 g = **0.287 mol** of $Cu(OH)_2$ units

In one unit of $Cu(OH)_2$ there are a total of three ions. One ion of Cu^{2+} and 2 ions of OH^-.

Therefore, if 1 mole of $Cu(OH)_2$ contains $3 \times 6.02 \times 10^{23}$ ions,

0.287 moles of $Cu(OH)_2$ contains ? ions

(0.287 moles × 3 × 6.02×10^{23} ions) / 1 mole = **5.18×10^{23} ions**

1. Calculate the mass of:
    a. 2.5 moles of sodium hydroxide.
    b. 0.2 moles of magnesium carbonate.
    c. 1.5 moles of zinc hydroxide.
    d. 1.2 moles of lead(II) nitrate.
2. Find the number of moles in:
    a. 1.6 g of methane
    b. 12.75 g of aluminium oxide.
    c. 0.5 g of calcium nitrate.
    d. 3.6 g of iron(III) sulfate.
3. Work out the number of:
    a. atoms in 3.8 g of copper.
    b. atoms in 5.5 g of sodium oxide.
    c. molecules in 80 g of nitrogen monoxide.
    d. ions present in 65 g of potassium sulfate.

Further reading on the mole.

RAM, RMM, RFM, mass number, isotopes, Avogadro's constant, and the mole by Mr D. Bugeja

Lesson 1: Lesson 2:

The mole

Amedeo Avogadro was an Italian scientist (1776-1856), who came up with the idea of a constant number of particles. In 1811, Avogadro published his work that explained that a volume of any gas was proportional to the number of gas atoms or molecules, at a given pressure and temperature. Therefore, for gases having the same volume, a relative mass could be found.

And this mass will have the same number of gas atoms/molecules. This law provided a way of calculating the number of atoms/molecules of the gas in a specific volume.

This idea wasn't accepted at the beginning, and it was later in 1909, when Jean Perrin (a French physicist) explained that the constant number of particles (which he termed Avogadro's constant), can be defined as the number of molecules in 16 grams of oxygen. This was later changed to be defined that one mole is the number of atoms found in exactly 12 grams of carbon-12.

3.2.2 Molecular and empirical formulae

There are **two** kinds of formulae for compounds:

(a) **molecular formula** (real formula) shows the actual numbers of atoms of each element present in a substance. For example:

- hydrogen peroxide (H_2O_2).
- ethene (C_2H_4).
- calcium chloride ($CaCl_2$).

(b) **empirical formula** (simplest formula) shows the simplest ratio of the atoms of each element present in a substance. For example:

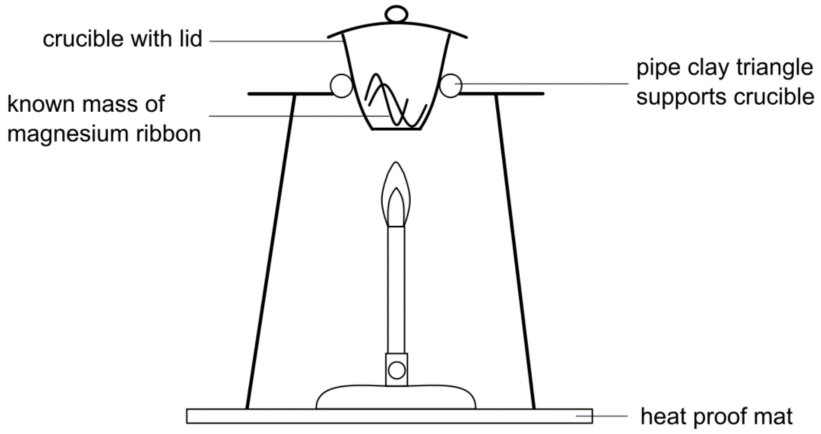
- hydrogen peroxide (H_2O_2) has an empirical formula of HO.
- ethene (C_2H_4) has an empirical formula of CH_2.
- calcium chloride ($CaCl_2$) has an empirical formula of $CaCl_2$. Note that in this case the molecular shows the simplest ratio between the elements.

3.2.3 Experimental determination of the empirical formula of a metal oxide

To work out the empirical formula of a compound, you need to know the masses of elements that combine together.

For example, magnesium combines with oxygen in the air to form magnesium oxide. The masses of elements that combine can be found by using the following procedure:

1. The apparatus is set up as shown in the following diagram:



2. The empty crucible and lid are weighed on an electronic balance.
3. A piece of magnesium ribbon is cleaned using emery paper to remove any impurities due to oxidation.
4. The magnesium ribbon is added to the crucible with lid and reweighed.
5. The crucible is heated over a blue Bunsen flame.
6. The lid is carefully raised at intervals to let oxygen in. The magnesium should burn brightly, leaving white magnesium oxide ash in the crucible.
7. When burning is complete, the crucible is left to cool (still with its lid on) and is reweighed.
8. Steps 5, 6, and 7 are repeated until constant mass is achieved.
9. An increase in mass should be noted from the start of the experiment, due to the oxygen in the air reacting with magnesium.

Sample Results

- Mass of crucible + lid = 19.24 g
- Mass of crucible + lid + magnesium = 20.68 g
- Mass of magnesium = 1.44 g
- Mass of crucible + lid + magnesium oxide = 21.64 g
- Mass of oxygen combined = 0.96 g
- So **1.44 g of magnesium** reacted with **0.96g of oxygen**.

The following table shows the steps taken to calculate the empirical formula:

	Magnesium	**Oxygen**
Mass	1.44 g	0.96 g
RAM	24	16
Quantity in moles ***(divide by molar mass)***	24 g is 1 mol 1.44 g is ? $\frac{1.44 \text{ g} \times 1}{24 \text{ g}}$ = 0.06 mol	16 g is 1 mol 0.96 g is ? $\frac{0.96 \text{ g} \times 1}{16 \text{ g}}$ = 0.06 mol
Convert to whole numbers. ***(Divide through by smallest number)***	$\frac{0.06}{0.06}$ = 1	$\frac{0.06}{0.06}$ = 1
Ratio	1	1
EMPIRICAL FORMULA	**MgO**	

Sources of error

The crucible lid was needed to prevent the magnesium from reacting all at once. The problem with allowing the magnesium to react like this is that a lot of MgO smoke would be produced and lost from the crucible. This would negatively affect the results obtained.

The crucible lid needs to be opened periodically to allow air (oxygen) to enter the crucible and react with the magnesium ribbon. Inevitably, some smoke is lost every time the lid is opened, and this affects the accuracy of the experiment.

To a lesser extent, magnesium also reacts with nitrogen, and this also includes a further source of error.

Worked example

Find the empirical formula of calcium chloride when 0.4 g of calcium reacts with 0.71 g of chlorine. (RAM: Ca = 40; Cl = 35.5)

	Ca	Cl
Mass	0.4 g	0.71 g
Quantity of moles (divide by molar mass)	40 g is 1 mol 0.4 g is ? $\frac{0.4 \times 1}{40}$ = 0.01 mol	35.5 g is 1 mol 0.71 g is ? $\frac{0.71 \times 1}{35.5}$ = 0.02 mol
Convert to whole numbers. (Divide through by smallest number)	$\frac{0.01}{0.01}$ = 1	$\frac{0.02}{0.01}$ = 2
Ratio	1	2
EMPIRICAL FORMULA	**$CaCl_2$**	

Calculating the formula of a compound

The concept of compounds that we have nowadays, is an idea that took time to develop and wasn't always as obvious and straight forward as we might see it. In the 18th century, this concept was still being debated and some scientists believed that elements could combine in any amount. In fact, at the time, scientists still didn't come up with a clear distinction between compounds and mixtures.

The first chemist to define this in his writing, was Joseph Proust in 1797, who explained that chemicals were made up of more than one type of atoms, combined in fixed proportions. This is now known as the 'Law of constant proportion'.

Calculating the formula of a substance from reacting masses by Mr D. Bugeja.

Lesson 5: Lesson 6:

3.2.4 Working out formulae using percentage composition

Percentage composition by mass of elements in a compound can be used to work out the molecular formula of that compound.

Worked Example

A hydrocarbon contains 85.7% carbon and 14.3% hydrogen by mass. This means that in 100 g of the compound, 85.7 g of carbon are combined with 14.3 g of hydrogen. The empirical formula of this hydrocarbon is worked out as follows:

(RAM: C = 12; H = 1)

	C	H
% by mass	85.7	14.3
Quantity in moles (divide by molar mass)	$\frac{85.7\text{ g}}{12\text{ g}} = 7.14$ mol	$\frac{14.3\text{ g}}{1\text{ g}} = 14.3$ mol
Divide by the smallest number	$\frac{7.14}{7.14} = 1$	$\frac{14.3}{7.14} = 2$
Ratio	1	2
Empirical formula	**CH_2**	

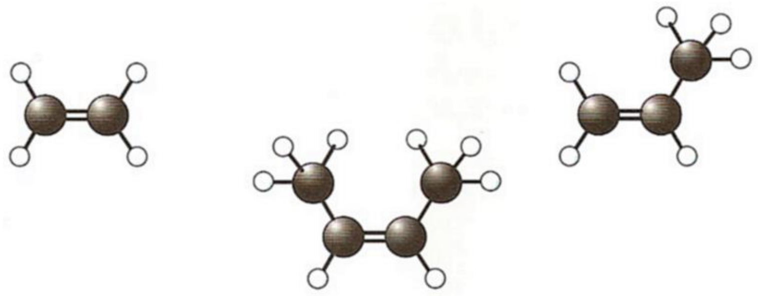
Working out percentage by mass calculations by Mr D. Bugeja

Lesson 3:

Lesson 4:

3.2.5 Converting empirical formulae into molecular formulae

In the previous worked example, CH_2 cannot be the real formula of the compound as carbon would have spare unbonded electrons. The **molecular formula** (the true formula) would have to be some multiple of CH_2, like C_2H_4 or C_3H_6 etc., as long as the ratio of C:H = 1:2.



All these compounds have an empirical formula of CH_2.

The molecular formula can be determined from the empirical formula if the RMM of the compound is known using the following statement:

Molecular mass = n (empirical mass)

where **n** is a multiple

Worked Example 1

What is the molecular formula of the compound having an empirical formula of CH_2 and RMM of 56. (RAM: C = 12; H = 1)

Empirical formula mass of CH_2 = 12 + (1 × 2) = 14

Molecular mass = n (empirical mass)

To find n, divide 56 by 14:

56/14 = 4, and so you need 4 units of $\mathbf{CH_2}$, in other words $\mathbf{4\ (CH_2) = C_4H_8}$

Therefore, the molecular formula is $\mathbf{C_4H_8}$.

Worked Example 2

(a) Determine the empirical formula of an organic compound which contains 66.7% carbon, 11.1% hydrogen, and 22.2% oxygen by mass.

	C	H	O
% by mass	66.7	11.1	22.2
Quantity in moles *(divide by molar mass)*	$\frac{66.7\text{ g}}{12\text{ g}} = 5.56$ mol	$\frac{11.1\text{ g}}{1\text{ g}} = 11.1$ mol	$\frac{22.2\text{ g}}{16\text{ g}} = 1.39$ mol
Divide by the smallest	$\frac{5.56}{1.39} = 4$	$\frac{11.1}{1.39} = 8$	$\frac{1.39}{1.39} = 1$
Ratio	4	8	1
Empirical formula	$\mathbf{C_4H_8O}$		

(b) If the RMM of the compound is 72, what is its molecular formula?

(RAM: C = 12; H = 1; O = 16)

The empirical formula mass = C_4H_8O = (12 × 4) + (1 × 8) + 16 = **72**

If the RMM of the organic compound is 72,

Therefore, $\frac{72}{72} = 1$

Then, there is one C_4H_8O unit in the molecular formula.

Therefore, the molecular formula is $\mathbf{C_4H_8O}$.

Relating empirical and molecular formulae of simple substances by Mr D. Bugeja.

3.2.6 Experimental determination of the empirical formula of a hydrated salt

A **hydrated salt** is an ionic compound in which a number of water molecules are attracted by the ions and therefore become embedded within its crystal lattice. The crystalline structure of these salts depends on the presence of this **water of crystallisation**.

The water of crystallisation is shown distinctly in the chemical formula of the hydrated salt. For example in hydrated copper(II) sulfate, the mole ratio of salt to water is 1 is to 5. So its formula is $CuSO_4.5H_2O$.

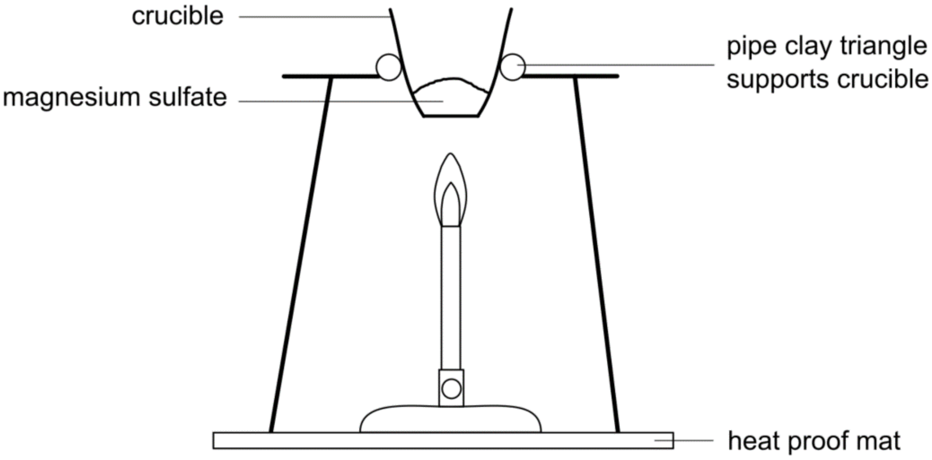
Water of crystallisation can be easily removed by heating to produce the **anhydrous** salt according to the following equation:

$$CuSO_4.5H_2O(s) \xrightarrow{heat} CuSO_4(s) + 5H_2O(l)$$

To work out the empirical formula of a hydrated salt, you need to know the masses of the anhydrous salt and the amount of water of crystallisation that combine together.

For example, hydrated magnesium sulfate can be heated to constant mass until all water of crystallisation is removed. No crucible lid is used in this case since water vapour needs to be lost.

1. The apparatus is set up as shown in the following diagram:



2. The empty crucible is weighed on a balance.
3. Some hydrated magnesium sulfate is added to the crucible and reweighed.
4. The crucible is heated over a blue Bunsen flame.
5. After heating for a while, the crucible is left to cool and is reweighed.
6. Steps 4 and 5 are repeated until constant mass is achieved.
7. A decrease in mass should be noted from the start of the experiment, due to the water of crystallisation lost to the surroundings.

Sample Results:

- Mass of crucible = 19.24 g
- Mass of crucible + hydrated magnesium sulfate = 24.16 g
- **Mass of hydrated magnesium sulfate = 4.92 g**
- Mass of crucible + anhydrous magnesium sulfate = 21.64 g
- **Mass of water removed = 2.52 g**
- **Mass of anhydrous magnesium sulfate** = 4.92 g – 2.52 g **= 2.40 g**

The following table shows the steps taken to calculate the empirical formula:

	Magnesium sulfate	Water of crystallisation
Mass	2.40 g	2.52 g
Molar masses	120 g	18 g
Quantity in moles *(divide by molar mass)*	120 g is 1 mol 2.40 g is ? $\frac{2.40\text{ g} \times 1}{120\text{ g}}$ = 0.02 mol	18 g is 1 mol 2.52 g is ? $\frac{2.52\text{ g} \times 1}{18\text{ g}}$ = 0.14 mol
Convert to whole numbers. *(Divide through by smallest number)*	$\frac{0.02}{0.02}$ = 1	$\frac{0.14}{0.02}$ = 7
Ratio	1	7
EMPIRICAL FORMULA	**$MgSO_4.7H_2O$**	

Worked Example

Hydrated iron(III) chloride contains 20.7% iron and 39.4% chlorine. Its relative formula mass is 270.5. Workout its formula.

% mass of salt = 20.7 + 39.4 = 60.1

% mass of water of crystallisation = 100 – 60.1 = 39.9

	Fe	Cl	H_2O
% by mass	20.7	39.4	39.9
Quantity in moles *(divide by molar mass)*	$\frac{20.7\text{ g}}{56\text{ g}}$ = 0.37 mol	$\frac{39.4\text{ g}}{35.5\text{ g}}$ = 1.11 mol	$\frac{39.9\text{ g}}{18\text{ g}}$ = 2.22 mol
Divide by the smallest number	$\frac{0.37}{0.37}$ = 1	$\frac{1.11}{0.37}$ = 3	$\frac{2.22}{0.37}$ = 6
Ratio	1	3	6
Empirical formula	$FeCl_3.6H_2O$		
Relative mass of empirical formula	56 + (3 x 35.5) + (6 x 18) = 270.5		
Formula of salt	Since empirical formula is equal to relative formula mass then the formula of the salt is equal to the empirical formula: **$FeCl_3.6H_2O$**		

Further reading on empirical and molecular formulae.

1. Find the empirical formula of a chloride of lead when 4.14 g of lead reacts with 1.42 g of chlorine.
2. An oxide of copper is heated in hydrogen to produce copper and water. In an experiment 12.7g of copper was produced from 16.3g of the oxide of copper.
    a. Write a word equation of the reaction taking place.
    b. Work out the mass of oxygen reacting.
    c. Find the formula of the oxide of copper.
3. Calculate the empirical formula of a compound made up of 2.18 g of carbon, 0.36 g of hydrogen and 1.46 g of oxygen.
4. 12.2 g of hydrated barium chloride crystals were placed in a crucible and heated strongly. After cooling the contents of the crucible were weighed. The process was repeated to achieve constant mass. At the end of the experiment the mass of anhydrous barium chloride obtained was 10.4 g.
    a. How many moles of anhydrous barium chloride were produced?
    b. Find the mass of water that evaporated.
    c. Calculate the number of moles of water.
    d. Find x in $BaCl_2.xH_2O$
    e. Explain why the contents were heated to constant mass.
5. Titanium chloride contains 25% titanium and 75% chlorine by mass. Work out the simplest formula of titanium chloride.
6. Work out the empirical formula of a compound made up of 27.4% sodium, 1.2% hydrogen, 14.3% carbon and 57.1% oxygen.
7. Butane has an empirical formula of C_2H_5. Its relative molecular mass is 58. What is the molecular formula of butane?
8. A hydrocarbon contains 92.3% carbon and 7.7% hydrogen.
    a. What is the empirical formula?
    b. Its relative molecular mass is 78. What is its molecular formula?
9. An organic liquid consists of carbon, hydrogen, and oxygen.
    a. Calculate the empirical formula of an organic liquid containing 26.67% of carbon and 2.22% of hydrogen, with the rest being oxygen.
    b. The M_r of the liquid is 90. What is its molecular formula?

3.2.7 Calculating the quantity of product formed from given quantity of reactant and vice-versa.

Apart from knowing what products are formed in a chemical reaction, chemists also use a balanced chemical equation to calculate the quantity of products formed from a given mass of starting material.

Consider the following worked example: Calculate (a) the mass of O_2 needed and (b) the mass of CO_2 produced, when 100 g of carbon are burned completely in air.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

According to the equation,

1 mole of carbon reacts with **1 mole of oxygen** to give **1 mole of carbon dioxide**.

The ratio of reacting particles is called **stoichiometry**.

There are **two** methods which can be used in these calculations.

Method 1

12 g of carbon react with **32 g** of oxygen to give **44 g** of carbon dioxide.

(a) Mass of O_2 needed:

12 g of carbon react with 32 g of oxygen
100 g of carbon react with ?

(100 g × 32 g) / 12 g = **266.67 g** of oxygen

(b) Mass of CO_2 produced:

12 g of carbon produce 44 g of carbon dioxide
100 g of carbon produce ?

(100 g × 44 g) / 12 g = **366.67 g** of carbon dioxide

Method 2

The quantity of moles of carbon in 100 g:
In 12 g of carbon there is 1 mole
In 100 g of carbon there is ? moles
(1 × 100) / 12 = 8.33 mol

(a) Mass of O_2 needed:

According to balanced equation:
1 mole of carbon reacts with 1 mole of oxygen
8.33 moles of carbon reacts with ? moles of oxygen
(1 × 8.33) / 1 = 8.33 moles of oxygen

1 mole of oxygen weighs 32 g
8.33 mole of oxygen weighs ?
(8.33 x 32) / 1 = **266.67 g** of oxygen gas

(b) Mass of CO_2 produced:

According to balanced equation:
1 mole of carbon produces 1 mole of carbon dioxide
8.33 moles of carbon produces ? moles of carbon dioxide
(1 × 8.33) / 1 = 8.33 moles of carbon dioxide

1 mole of carbon dioxide weighs 44 g
8.33 mole of carbon dioxide weighs ?
(8.33 x 44) / 1 = **366.67 g** of carbon dioxide gas

Calculating the amounts of products from a given amount and vice versa by Mr D. Bugeja
Lesson 8:
Lesson 9:

1. Iron reacts with sulfuric acid to give iron(II) sulfate and hydrogen gas.
   a. Write a balanced chemical equation for the reaction taking place.
   b. 5 g of iron are added to acid. Calculate the number of moles of iron reacting.
   c. From the chemical equation, work out the number of moles of iron(II) sulfate.
   d. Calculate the mass of iron(II) sulfate formed in the reaction.
2. When 18 g of zinc nitrate are heated, they thermally decompose as follows:

   $$2\,Zn(NO_3)_2\,(s) \rightarrow 2\,ZnO\,(s) + 4\,NO_2\,(g) + O_2\,(g)$$

   a. Calculate the number of moles of zinc nitrate.
   b. Determine the number of moles of zinc oxide produced.
   c. Work out the mass of zinc oxide.
3. Sodium reacts in air as in the following equation: $4Na(s) + O_2\,(g) \rightarrow 2Na_2O(s)$
   What mass of sodium oxide can be made by reacting 2.3 g of sodium in air?
4. Magnesium chloride solution reacts with silver nitrate solution to form a precipitate of silver chloride.
   a. Write a balanced chemical equation for the reaction.
   b. Work out the mass of silver chloride produced when 0.95 g of magnesium chloride react.
5. Hydrogen burns in chlorine to form hydrogen chloride gas. Calculate the mass of hydrogen chloride produced when 1 g of hydrogen is reacted with chlorine.
6. Calculate the mass of oxygen needed to burn 20 g of lead(II) sulfide using the following balanced chemical equation:

   $$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$$

3.2.8 Avogadro's Law

To find the quantity of a certain gas, it is more practical to measure its **volume** rather than its mass.

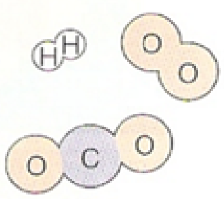
The volume of a certain mass of gas depends on its temperature and its pressure. We therefore need to state the temperature and the pressure at which a volume was measured. Standard temperature is 0 °C and standard pressure is 1 atmosphere (or 760 mm of mercury).

In this respect, Amadeo Avogadro formulated a law that states that:

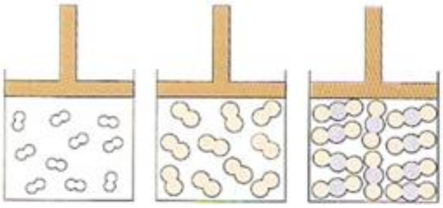
One mole of **ANY** gas occupies the same volume at the same temperature and pressure.

The volume of one mole of a gas is called the **molar gas volume** and occupies 22.4 dm^3 at standard temperature and pressure (STP). This value will be given in all exam papers.

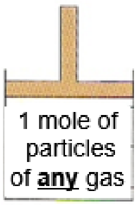
Note that 1 dm^3 is equivalent to 1000 cm^3, so 22.4 dm^3 is equivalent to 22400 cm^3.



You might expect one mole of carbon dioxide molecules to take up more space than a mole of hydrogen or oxygen molecules.



But this doesn't happen. At a given temperature and pressure, a mole of particles of **ANY** gas occupies the same volume.



At STP the volume of 1 mole of any gas is 22.4 dm^3.

Worked Examples

1. What volume does 0.25 moles of hydrogen gas occupy at STP?

1 mole of hydrogen gas occupies 22.4 dm^3

0.25 moles of hydrogen gas occupy ?

$$\frac{0.25 \text{ moles} \times 22.4 \text{ dm}^3}{1 \text{ mole}} = \mathbf{5.6\ dm^3\ of\ hydrogen}$$

2. Calculate the number of moles of ammonia gas (NH_3) in a volume of 72 dm^3 of the gas measured at STP.

In 22.4 dm^3 of ammonia gas there is 1 mole,

in 72 dm^3 of ammonia gas there is ?

$$\frac{72 \text{ dm}^3 \times 1 \text{ mole}}{22.4 \text{ dm}^3} = 3.21 \text{ moles of ammonia}$$

3. A reaction produces 100 cm^3 of hydrogen at STP. How many moles of hydrogen are present?

First convert cm^3 into dm^3: 100cm^3 / 1000cm^3 = 0.1 dm^3

22.4 dm^3 of hydrogen gas is occupied by 1 mole of hydrogen.

0.1 dm^3 of hydrogen gas is occupied by ?

$$\frac{0.1 \text{ dm}^3 \times 1 \text{ mole}}{22.4 \text{ dm}^3} = \mathbf{0.0045\ moles\ of\ hydrogen}$$

4. What volume does 22 g of carbon dioxide occupy at STP? (RAM of C = 12; O = 16)

Since this time the mass is given, you must first convert the mass given into moles.

mass of 1 mole of CO_2 = 12 + (16 × 2) = 44 g

44 g of carbon dioxide is 1 mole.

22 g of carbon dioxide is ?

(22 g × 1 mole) / 44 g = 0.5 moles of carbon dioxide

1 mole occupies 22.4 dm^3

0.5 moles occupy ?

$$\frac{0.5 \text{ moles} \times 22.4 \text{ dm}^3}{1 \text{ mole}} = \mathbf{11.2\ dm^3}$$

5. Calculate the mass of oxygen gas (O_2) in a volume of 6 dm^3 of the gas measured at STP?

In 22.4 dm^3 of oxygen gas there is 1 mole,

in 6 dm^3 of oxygen gas there is ?

$$\frac{6\ dm^3 \times 1\ mole}{22.4\ dm^3} = \mathbf{0.27\ moles}$$

Since this time the mass is needed, now you must convert the moles found into mass.

Therefore, 1 mole of oxygen gas (O_2) weighs 32 g

0.27 moles of oxygen gas (O_2) weigh ?

$$\frac{0.27\ moles \times 32\ g}{1\ mole} = \mathbf{8.64\ g}$$

6. Sulfur burns in air to form sulfur dioxide. What volume of this polluting gas (SO_2) is produced when 1 g of sulfur burns at STP? (RAM of S = 32; the molar gas volume at STP is 22.4 dm^3)

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

32 g of sulfur is 1 mole

1 g of sulfur is ?

$$\frac{1g \times 1\ mole}{32\ g} = 0.031\ \text{moles of sulfur}$$

According to the equation, 1 mole of sulfur gives 1 mole of sulfur dioxide,

0.031 moles of sulfur gives 0.031 moles of sulfur dioxide

Therefore, if 1 mole of sulfur dioxide occupies 22.4 dm^3 at STP,

0.031 moles of sulfur dioxide occupies ?

$$\frac{0.031\ moles \times 22.4\ dm^3}{1\ mole} = \mathbf{0.69\ dm^3}$$

7. Phosphorus burns in air to form phosphorus(V) oxide, P_4O_{10}, as shown below

$$4P(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$$

What volume of oxygen is needed to react with 3.1 g of phosphorus at STP?

4 × 31 g = 124 g of phosphorus reacts with 5 × 22400 cm^3 = 112000 cm^3 of oxygen gas

3.1 g of phosphorus reacts with ? oxygen gas

(3.1 × 112000) / 124 = **2800 cm^3 of oxygen gas**

Further reading on Avogadro's law.

1. Calculate the volume of 0.3 mole of hydrogen chloride (HCl) gas at STP
2. Calculate the volume of 4.4 g of carbon dioxide at STP
3. When 12.4 g of copper(II) carbonate is heated, it thermally decomposes to give copper(II) oxide and carbon dioxide.
    a. Write a balanced chemical equation for the reaction.
    b. Work out the number of moles of copper(II) carbonate being heated.
    c. Determine the number of moles of copper(II) oxide produced.
    d. Work out the mass of copper(II) oxide.
    e. Work out the volume of carbon dioxide gas produced at STP.
4. 120 cm^3 of hydrogen gas at STP are released when zinc reacts with hydrochloric acid.
    a. Write a balanced chemical equation for the reaction.
    b. Work out the number of moles of hydrogen gas produced.
    c. Determine the number of moles of zinc required.
    d. Work out the mass of zinc used in this reaction.
5. Ammonium sulfate reacts with potassium hydroxide to release a pungent smelling gas that turns damp red litmus paper blue.

    $$(NH_4)_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2NH_3(g) + 2H_2O(l)$$

    a. Calculate the mass of ammonium sulfate required to react with 42 g of potassium hydroxide.
    b. Calculate the volume of ammonia gas released from this reaction at STP.
6. Calculate the volume of ammonia produced at STP when 26 g of nitrogen reacts with hydrogen.
7. Iron(III) oxide is reduced to iron by carbon monoxide as in the following chemical equation:

    $$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

    a. Calculate the mass of iron that could be produced when 800 kg of iron(III) oxide are used.
    b. Work out the volume of carbon dioxide gas at STP that is released from this reaction.
8. When calcium carbonate is heated, it decomposes into solid calcium oxide and carbon dioxide gas as shown in the equation below.

    $$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

    Calculate the mass of calcium carbonate needed to give 200 cm^3 of CO_2 at STP.

3.2.9 Calculating the theoretical and percentage yield of a product

The yield is the amount of product one obtains from a reaction. Consider a factory that makes paint or fertilisers. The owners will want the highest yield possible, for the lowest cost! In this section you'll learn how to calculate the percentage (%) yield from a reaction.

The % yield can be calculated as follows:

$$\textbf{\% yield} = \frac{\text{actual mass obtained}}{\text{theoretical (calculated) mass}} \times 100$$

Worked Example

The medical drug aspirin is made from salicylic acid. 1 mole of salicylic acid gives 1 mole of aspirin as shown below.

$$\underset{\text{salicylic acid}}{C_7H_6O_3} \rightarrow \underset{\text{aspirin}}{C_9H_8O_4}$$

In a trial, 100.0 grams of salicylic acid gave 121.2 grams of aspirin. What was the % yield? (RAM of C = 12; H = 1; O = 16)

(First find the RMM of a salicylic acid and aspirin)

RMM of salicylic acid, $C_7H_6O_3$: $(12 \times 7) + (1 \times 6) + (16 \times 3) = 138$

RMM of aspirin, $C_9H_8O_4$: $(12 \times 9) + (1 \times 8) + (16 \times 4) = 180$

(Now find the number of moles of salicylic acid used in the reaction)

Therefore, if 138 g of salicylic acid is 1 mole,

100.0 g of salicylic acid is ?

$$\frac{100\text{ g} \times 1\text{ mole}}{138\text{ g}} = 0.725\text{ moles}$$

(Now find the number of moles of aspirin produced in the reaction using ratios)

If 1 mole of salicylic acid gives 1 mole of aspirin,

0.725 moles of salicylic acid give 0.725 moles of aspirin.

(Now find the mass of aspirin formed in the reaction)

Therefore, if 1 mole of aspirin weighs 180 g,

0.725 moles of aspirin weighs ?

$$\frac{0.725\text{ moles} \times 180\text{ g}}{1\text{ mole}} = 130.5\text{ g}$$

(Finally calculate the % yield of aspirin in the reaction)

So 130.5 g is the **theoretical (calculated) mass/yield** for the reaction.

But the **actual mass/yield** obtained in the trial was 121.2 g.

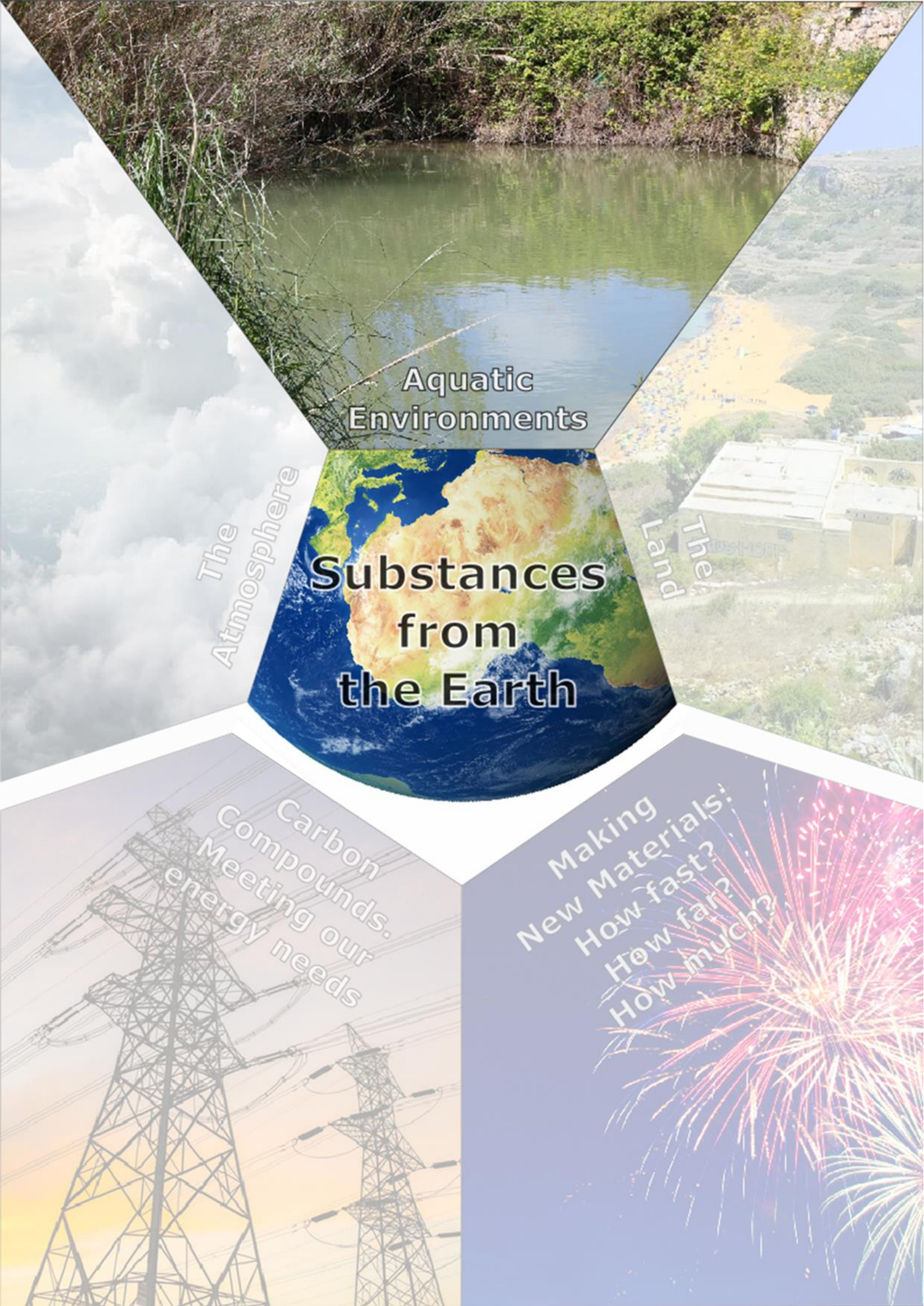
So % yield: $\frac{121.2\text{ g} \times 100}{130.5\text{ g}} = \mathbf{92.9\%}$

1. In an experiment to make calcium oxide the predicted yield was 2.8 g. The actual yield was 2.0 g. Calculate the percentage yield.
2. Copper(II) oxide is reduced to copper by passing hydrogen gas over the oxide:

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

   a. Calculate the volume of H_2 gas at STP needed to react with 8 g of copper(II) oxide.
   b. Calculate the mass of copper metal obtained from this reaction.
   c. The students obtained 5.8 g of copper. Work out the percentage yield.

Calculating the theoretical and empirical yield of products by Mr D. Bugeja



Aquatic Environments

The Atmosphere

Substances from the Earth

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

4 Groups in the Periodic Table – LO 6

4.1 Some groups in the Periodic Table. Distinguishing between metals and non-metals

By the end of this section, you should be able to:

- Name the groups of the Periodic Table. *(Limited to alkali metals, alkaline earth metals, transition metals, halogens, and noble gases.)* (6.1a)
- Distinguish between metals and non-metals in terms of their physical properties. (6.2a)

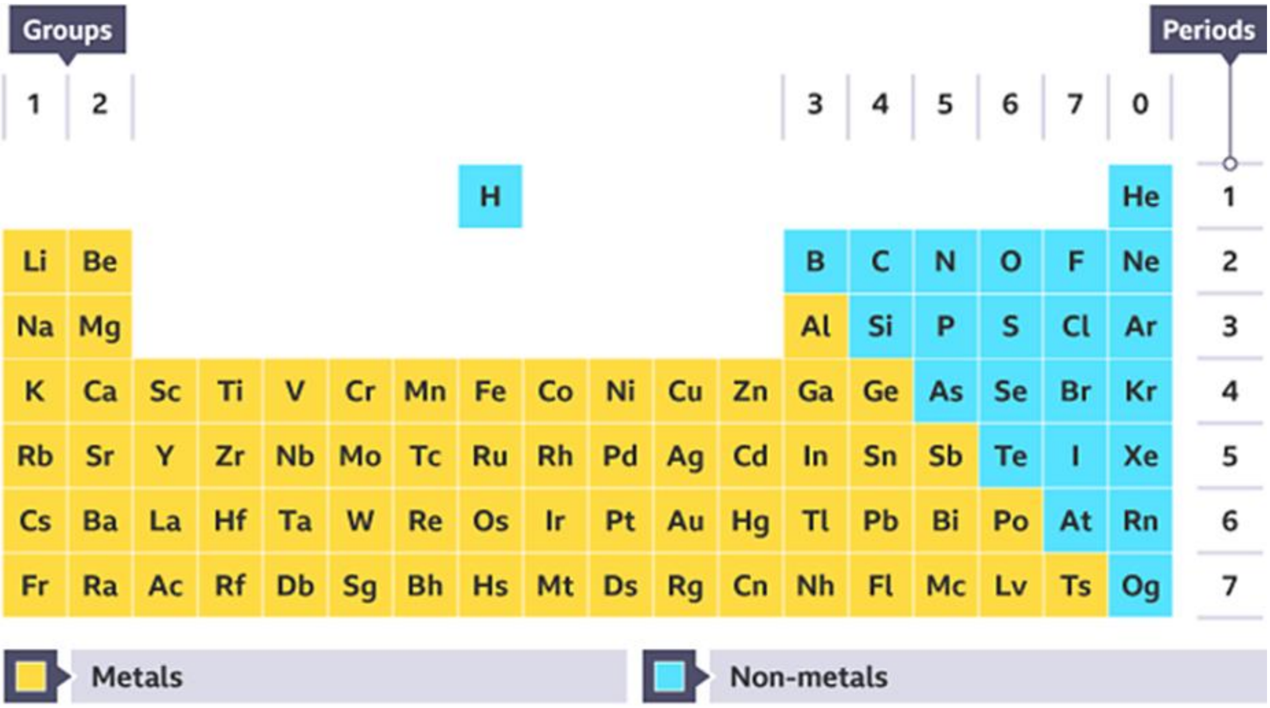
2019: The International Year of the Periodic Table

The Periodic Table is such an essential tool for anyone who studies or works in science, that UNESCO designated 2019 as the International Year of the Periodic Table.

This marks the 150th anniversary from the Mendeleev Periodic Table. UNESCO is the United Nations Educational, Scientific and Cultural Organization, which aims at promoting world peace and security through international cooperation in education, arts, sciences, and culture.

4.1.1 Metals and non-metals

The Periodic Table is a tabular display containing information about the elements. They have been organised in groups of elements that have similar physical and chemical properties. About 80% of elements are metals. The Periodic Table below shows that metals are found on the left-hand side of the Periodic Table. The non-metals are found on the right-hand side.



Groups

1	2											3	4	5	6	7	0	Periods
						H											He	1
Li	Be											B	C	N	O	F	Ne	2
Na	Mg											Al	Si	P	S	Cl	Ar	3
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	4
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	5
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	6
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og	7

🟨 Metals 🟦 Non-metals

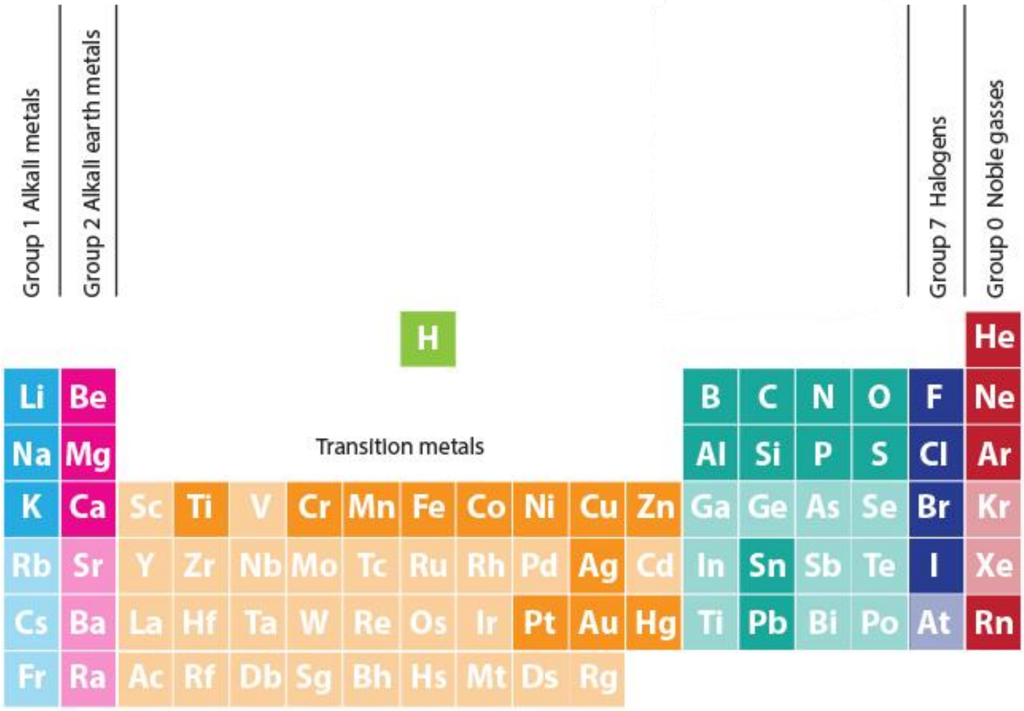
Source: https://www.bbc.co.uk/bitesize

The table below compares the typical properties of metals and non-metals.

Metals	**Non-Metals**
E.g. Gold, silver, aluminium, copper, magnesium, sodium	E.g. Hydrogen, oxygen, chlorine, sulfur
Usually metals: • are solid at room temperature. • they have relatively high melting and boiling points. • are good conductors of heat and electricity. • are shiny. • are malleable, that is, they can be hammered into new shapes without breaking. • are hard and strong. • have a high density. • are sonorous, that is, they make a ringing sound when hit. • Are ductile that is they can be drawn into wires.	Many non-metals: • are gases or liquids at room temperature because they have low melting and boiling points, however sulfur is a solid at room temperature. • are poor conductors of heat and electricity. • are dull in their appearance. • are weak and brittle, that is, they break easily. • are not sonorous.

4.1.2 Names of some groups in the Periodic Table

Metals and non-metals are further divided into other groups: alkali metals (group 1), alkaline earth metals (group 2), transition metals, halogens (group 7), and noble gases (group 0).



Adapted from: All elements of the periodic table are arranged into 9 group (mammothmemory.net)

4.2 Group 1 and group 7 elements

4.2.1 Group 1 elements

By the end of this section, you should be able to:

- List common uses of group 1 metal compounds. *(Limited to sodium chloride, potassium nitrate, and sodium hydrogencarbonate.)* (6.1d)
- Describe trends in physical and chemical properties of group 1 metals. *(Limited to: Physical properties: melting/boiling points and hardness. Chemical properties: reactions of metals with water to form alkalis and with oxygen to form simple oxides.)* (6.2d)

Some common uses of group 1 metal compounds include:

- Sodium compounds are used in table salt, baking soda, soap, glass, and paper. For example, sodium chloride is the main source of dietary sodium and sodium hydrogencarbonate is used as a leavening agent in baking.
- Potassium compounds are used in fertilizers, matches, fireworks, and explosives. For example, potassium nitrate is a key ingredient in gunpowder.

Source: chem.libretexts.org

Group 1 metals all have one electron in the outer shell. They are called alkali metals because when they react with water, they form alkaline solutions.

Lithium

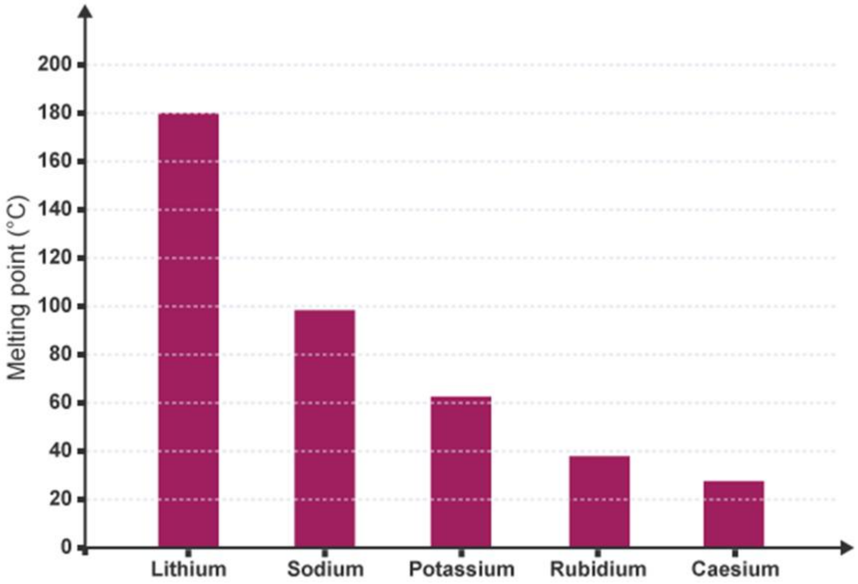
Lithium was created during the Big Bang together with hydrogen and helium. It is still a rare element since it is later used as the stars' fuel. On Earth, it is primarily used in batteries for electronic equipment, pacemakers, electric car batteries and as a medication for the treatment of bipolar disorder. It is also used in glass to make it more resistant to thermal shock, for light and strong aluminium-lithium alloys, and in grease to make it temperature resistant even at –60 °C. Lithium is extracted by electrolysis of concentrated brine originating from Chile and Bolivia.

4.2.1.1 Physical properties of group 1 metals

The typical physical properties of alkali metals are listed below.

- They are soft. In fact, they can be cut easily with a knife.
- Their melting and boiling points are relatively low when compared to other metals.

The graph below shows the melting points of the metals in the alkali metal group. One can observe that the melting point decreases on going down the group.



Source: https://www.bbc.co.uk/bitesize

4.2.1.2 Chemical properties of group 1 metals

Alkali metals are kept under oil to prevent them from reacting with water vapour and gases in the air.

Adapted from: https://www.sciencephoto.com/media/1890/view

4.2.1.2.1 Reactions of alkali metals with water

Group 1 metals react violently with cold water, floating on the surface. Hydrogen gas and a clear alkaline solution of the hydroxide are formed according to the following general word equation:

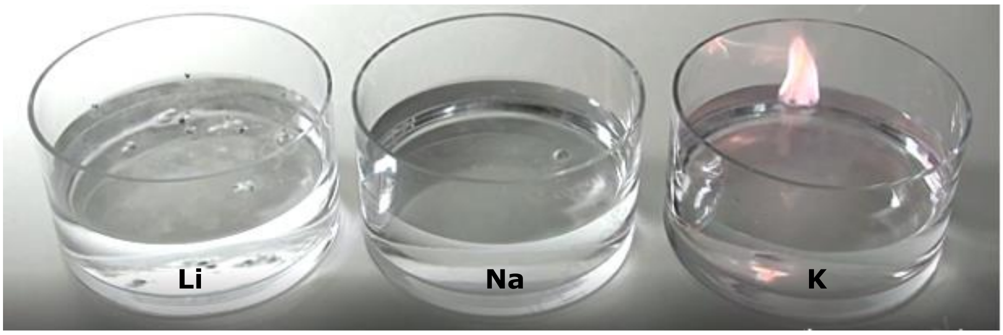
metal + water ⟶ metal hydroxide + hydrogen

Lithium reacts gradually with cold water by fizzing steadily. Sodium reacts more vigorously, melting on contact with cold water to form a silvery ball and fizzing rapidly while potassium burns with a lilac flame, fizzing violently upon touching the water. The chemical equations show the similarity with which these metals react with water.

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$$



The reactions of alkali metals with water show that their reactivity increases down the group.

Reactivity of alkali metals with water.

4.2.1.2.2 Reactions of alkali metals with oxygen

Alkali metals react with oxygen very quickly to form the oxide as shown by the following general word equation.

metal + oxygen ⟶ metal oxide

When an alkali metal is cut, it reacts readily with air to form an oxide on the surface. This oxide tarnishes the metal. To prevent this from happening, alkali metals are stored under oil.

They react readily in air as shown by the equations below. Note the similarity between the two equations.

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

$$4K(s) + O_2(g) \longrightarrow 2K_2O(s)$$

However, the reaction of sodium with the oxygen present in air is slower than that of potassium.

The reactions of alkali metals with oxygen go to show that reactivity increases down a group of metals.

Further information on group 1 elements.

4.2.2 Group 7 elements

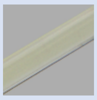
By the end of this section, you should be able to:

- List common uses of halogens. *(E.g. Bleaching and antibacterial action of chlorine in water and antiseptic properties of iodine.)* (6.1b)
- Describe the trends in physical and chemical properties of group 7 elements. *(Limited to state and colours of halogens at room temperature and reactions of halogens with hydrogen.)* (6.2b)
- Investigate displacement reactions of halogen/halide mixtures to construct a reactivity series of non-metals. *(Limited to chlorine, bromine, and iodine. Represent reactions using balanced chemical equations.)* (6.3b part)

Group 7 elements are known as halogens.

4.2.2.1 Physical properties of the halogens

The physical properties of the halogens are summarised in the table below.

Halogen	State at room temperature	Colour	Image
Fluorine	Gas	Pale yellow	
Chlorine	Gas	Greenish yellow	
Bromine	Liquid	Reddish brown	
Iodine	Solid	Dark Violet	

Source of images: https://images-of-elements.com/

Note that on going down the group the halogens change from gases to liquids to solids since the melting and boiling point of group 7 elements increases down the group. Their colour becomes darker as well.

4.2.2.2 Chemical properties of the halogens

The halogens react with hydrogen to produce compounds called **hydrogen halides**.

hydrogen + halogen ⟶ hydrogen halide

$$H_2(g) + X_2(g) \longrightarrow 2HX(g)$$

hydrogen + chlorine ⟶ hydrogen chloride

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

The table below describes what is observed when halogens react with hydrogen.

Halogen	Reaction with hydrogen
Fluorine	Explodes in the cold and dark.
Chlorine	Explodes in sunlight.
Bromine	Reaction occurs at high temperatures.
Iodine	Very slow reaction when heated strongly.

The hydrogen halides are gases at room temperature. They dissolve in water to produce acidic solutions.

Hydrogen chloride dissolves in water to produce hydrochloric acid, $HCl(aq)$.

As can be seen from the above reactions the reactivity of the halogens decreases down the group.

4.2.2.3 Uses of halogens

The following are some common uses:

- Chlorine compounds dissolved in water are used for bleaching, so they are used to make fabrics and paper whiter.
- Chlorine compounds act as an antibacterial agent, so they are used in water sanitation (drinking water, pools, etc.)
- Chlorine is used to make plastics like polyvinyl chloride (PVC) and pharmaceutical products.
- Iodine dissolved in potassium iodide solution is an antiseptic, so it is used to clean and sanitise wounds.
- Fluorine compounds are found in toothpaste as they prevent tooth decay.

Astatine

Astatine is incredibly scarce on earth. There are no more than 50 milligrams in the earth's crust at a given time. Its name is derived from the Greek 'astatos' which means unstable as its most stable isotope has a half-life of just over 8 hours.

Based on the trends down the halogen group, it's likely to be a black solid. It is also the only halogen that is monoatomic.

4.2.2.4 Displacement reactions

When chlorine (as a gas or dissolved in water) is added to sodium bromide solution, the chlorine replaces the bromide ion. The clear solution becomes dark orange. This colour change is due to bromine as it changes from Br^- ions to Br_2 and simultaneously chlorine changes from Cl_2 to Cl^- ions in sodium chloride. Chlorine is more reactive than bromine, so it displaces bromine from sodium bromide.

In this equation, the Cl and Br swap places:

$$\text{chlorine} + \text{sodium bromide} \longrightarrow \text{sodium chloride} + \text{bromine}$$

$$Cl_2(aq) + 2NaBr(aq) \longrightarrow 2NaCl(aq) + Br_2(aq)$$

This type of reaction happens with all the halogens.

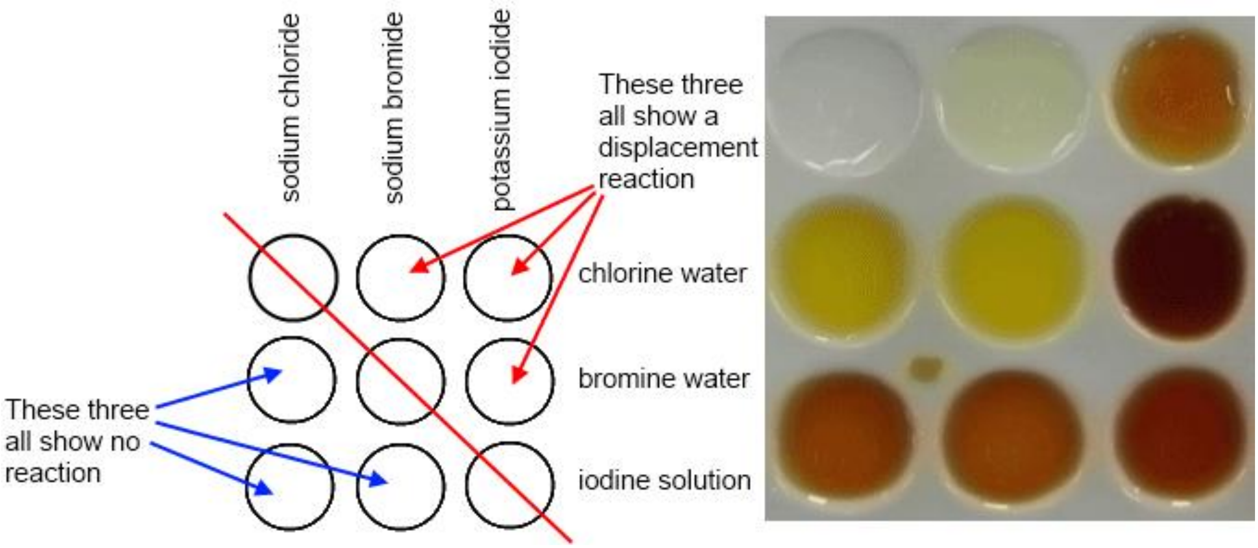
A more reactive halogen displaces a less reactive halogen
from a solution of one of its salts.

If you test different combinations of the halogens and their salts, you can work out a reactivity series for Group 7.

- The most reactive halogen displaces all of the other halogens from solutions of their salts and is itself displaced by none of the others.
- The least reactive halogen displaces none of the others and is itself displaced by all of the others.

It doesn't matter whether you use sodium salts or potassium salts – it works the same for both types.

	Sodium chloride ($NaCl$)	**Sodium bromide ($NaBr$)**	**Potassium iodide (KI)**
Chlorine (Cl_2)	X	Chlorine displaces the bromide ions. Yellow–orange colour of bromine observed.	Chlorine displaces the iodide ions. Brown colour of iodine observed.
Bromine (Br_2)	No reaction	X	Bromine displaces the iodide ions. Brown colour of iodine is observed.
Iodine (I_2)	No reaction	No reaction	X



Source: https://chemlegin.wordpress.com/2013/07/

Watch an experiment on the displacement reactions of the halogens.

Further information about group 7 elements.

4.2.3 Comparison of trends in reactivity in groups 1 and 7

By the end of this section, you should be able to:

- Compare trends in reactivity found in groups 1 and 7 using atomic structures to explain the variation of reactivity within a group. (6.3d)

An atom is made up of neutrons, protons, and electrons. Protons in the nucleus attract electrons in the electron shells.

When metals such as group 1 elements react, they lose one electron to form the ion.

Going down the group, the atom gets larger.

The outer electron to be lost is further away from the nucleus and there is increased shielding.

So there is less attraction between the nucleus and the outermost electron.

Therefore, less energy is needed for the atom to lose the outer electron and is more reactive.

Non-metals, such as group 7 elements, form ions by attracting electrons which are added to their outer shell. The attractive capability of the atoms decreases going down the group as the atoms get larger.

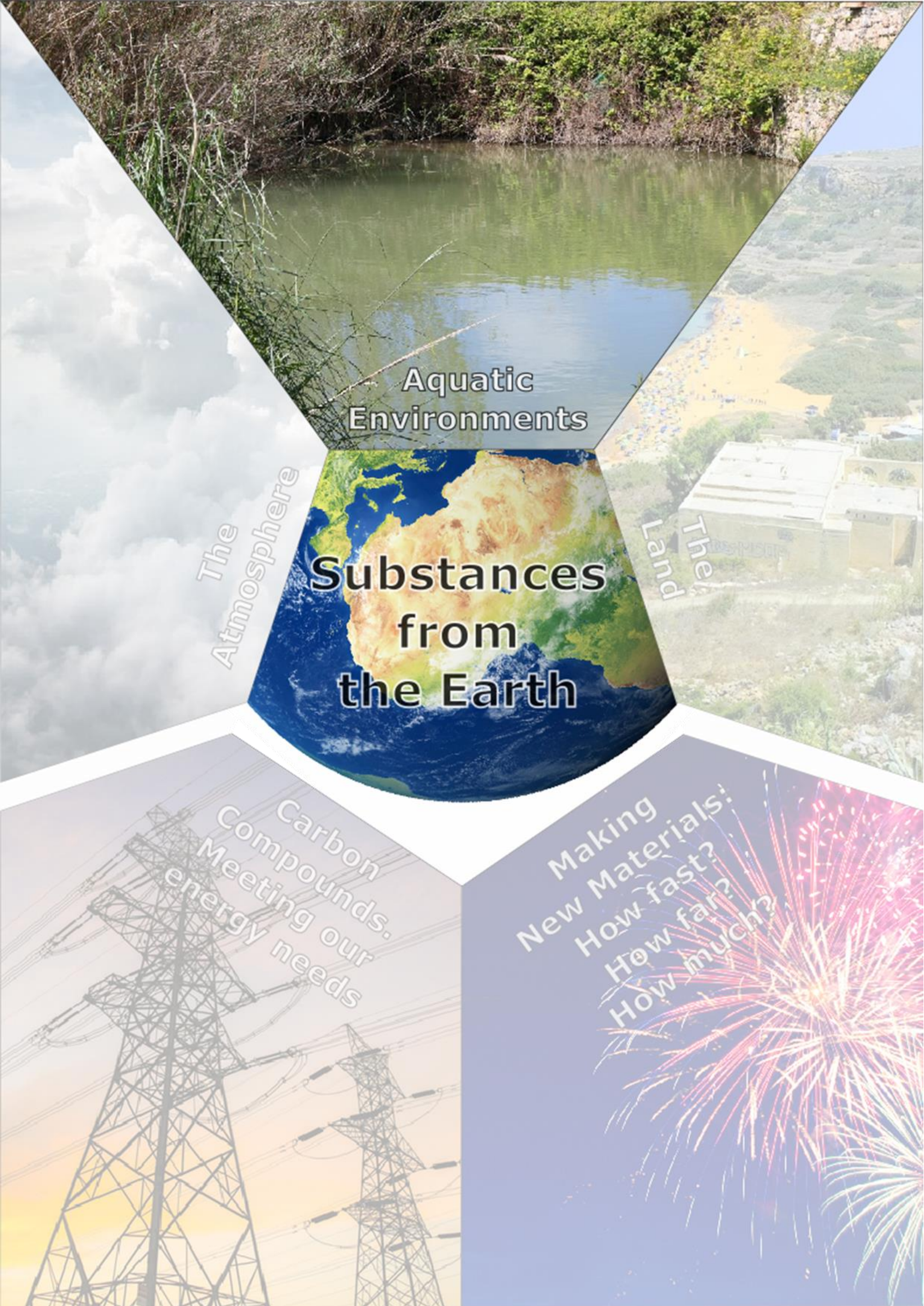
Going down the group, the atom gets larger.

The outer shell receiving an incoming electron is further away from the nucleus and has more shielding.

So, there is less attraction between the nucleus and the incoming electron.

As a result much more energy is required for the atom to gain the incoming electron and is less reactive.

Substances from the Earth



5 The conduction of electricity through solutions and molten salts – LO 5

5.1 Classifying materials as conductors, non-conductors, electrolytes, or non-electrolytes

By the end of this section, you should be able to:

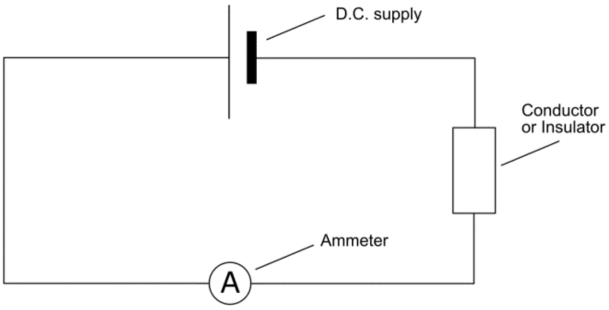
- Give examples of conductors and non-conductors (insulators), electrolytes and non-electrolytes. (5.1*a*)
- Define conductors and non-conductors (insulators), electrolytes and non-electrolytes of electricity. (5.2*a*)
- State whether solid/molten ionic and covalent substances conduct electricity when connected to *a* DC circuit. (5.1b)
- Perform an experiment to show what happens when an electric current passes through solids, molten ionic salts, graphite, and covalent substances. (5.2b)
- Explain why conductive solids, molten ionic salts, and graphite conduct electricity but solid ionic and covalent substances do not. (5.3b)

This chapter deals with electrolysis. This process involves passing current electricity through various substances in the molten and/or aqueous states in order to bring about chemical changes. Electrolysis has many uses such as the extraction of useful metals, electroplating, etc.

5.1.1 Conductors and non-conductors

Testing the conduction of electricity in solids.

Circuit 1 can be set up to test whether a solid material is a conductor or an insulator.



Circuit 1: Testing the conduction of electricity in solids.

Different solids, as shown in the table below, are connected to circuit 1 to test whether they conduct electricity or not. The ammeter reads a current when the solid conducts electricity.

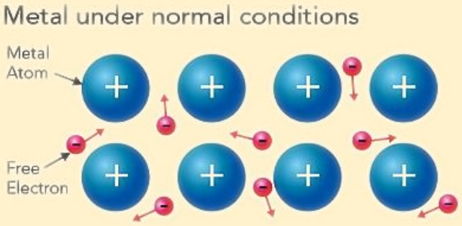
Material	Does the ammeter read a current?	Conductor/ insulator
iron nail	yes	conductor
plastic spoon	no	insulator
aluminium foil	yes	conductor
carbon rod (graphite)	yes	conductor
dry wooden block	no	insulator

From the above experiment it can be concluded that:

- Aluminium and iron (metals) as well as graphite are conductors of electricity.
- A **conductor** allows an electric current to pass through it.
- An **insulator (non-conductor)** does not allow an electric current to pass through it such as plastic, wood, and rubber.
- No chemical changes occur when electricity passes through a conductor.

Why do metals conduct electricity?

The electrons in the outer shell of the atoms of a metal move freely throughout the metal. They are known as delocalised electrons. Metals are conductors of electricity due to the delocalised electrons within the metal. These are responsible for the flow of current electricity throughout the metal.



Why does graphite conduct electricity?

Graphite is the only non-metal that conducts electricity. Just like metals graphite has delocalised electrons. These electrons are free to move between layers in the structure of graphite, so it can conduct electricity. This makes graphite useful in batteries and electrolysis.

Semi-conductors – Silicon

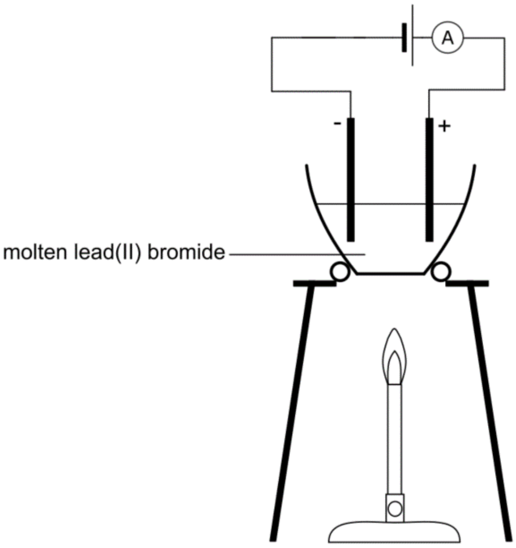
A semiconductor is a material which has an electrical conductivity value falling between that of a conductor, such as copper, and an insulator, such as glass. Its properties can be modified to form diodes, transistors, and most modern electronics.

Some examples of semiconductors are silicon, germanium, gallium arsenide, and elements near the "metalloid staircase" on the Periodic Table.

5.1.2 Electrolytes and non-electrolytes

5.1.2.1 Testing the conduction of electricity in molten materials

Circuit 2 shows the circuit setup used to test the conductivity of a molten substance.



Circuit 2 is used to test whether molten materials conduct electricity.

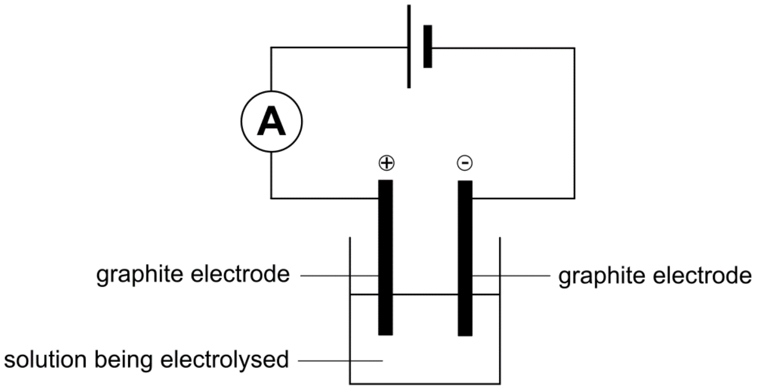
Material	Does the ammeter register a current?	Observations	Electrolyte?
molten lead(II) bromide	yes	shiny substance and a reddish-brown gas are produced	yes
molten wax	no	nothing happens	no
molten sodium chloride	yes	shiny substance and a yellowish-green gas are produced	yes
sodium chloride crystals	no	nothing happens	no
petroleum jelly	no	nothing happens	no
solid wax	no	nothing happens	no

From the above experiment it can be concluded that:

- Ionic solids in the molten state conduct electricity. At the same time a chemical change takes place. These are called **electrolytes**.
- When an ionic substance such as lead(II) bromide and sodium chloride is in the solid form it does not conduct electricity.
- Covalent substances such as petroleum jelly and wax, whether in the solid or molten state do not conduct electricity.

5.1.2.2 Testing the conduction of electricity in liquids/solutions

Circuit 3 shows the circuit setup to test the conductivity of solutions or liquids.



Circuit 3 is used to test whether solutions or liquids conduct electricity.

Material	Observations	Does the ammeter register a current?	Electrolyte?
sodium chloride solution	gases produced	yes	yes
paraffin oil	nothing happens	no	no
sulfuric acid	gases produced	yes	yes
ethanol	nothing happens	no	no
ethanoic acid	gases produced	yes	yes
distilled water	nothing happens	no	no

From the above experiment it can be concluded that:

- Ionic compounds in solution conduct electricity. At the same time a chemical change takes place. Aqueous ionic compounds are **electrolytes**.
- A pure covalent liquid such as paraffin oil, ethanol, and distilled water does not conduct electricity.
- The ammeter registers a larger current when a strong acid such as sulfuric acid is tested compared to a weak acid such as ethanoic acid.

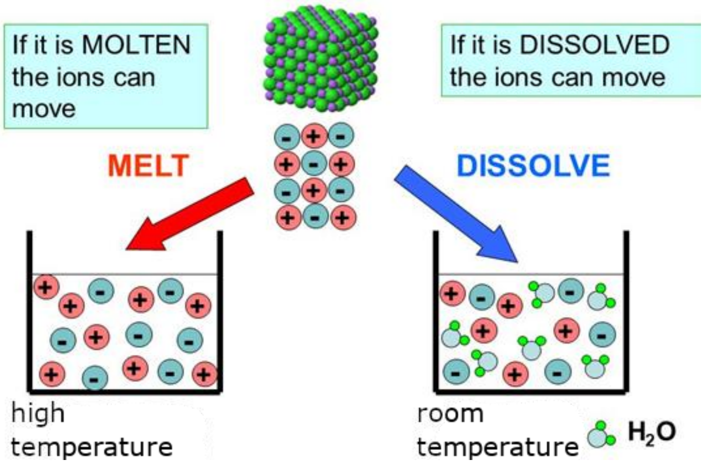
Some definitions:

- An **electrolyte** is a compound containing ions which when molten or in solution conducts an electric current and is decomposed by it, such as acids, alkalis, and soluble salts.
- A **strong electrolyte** conducts electricity easily because it has a large number of mobile ions. Strong acids and alkalis are strong electrolytes since they completely ionise in solution. Salts that are moderately to highly soluble in water are also strong electrolytes.
- A **weak electrolyte** poorly conducts electricity. A weak acid or alkali are weak electrolytes since they partially ionise in solution. Salts that are weakly soluble in water are also weak electrolytes.
- A **non-electrolyte** is a substance that does not conduct electricity either when molten or in solution. Molecular or covalent substances such as ethanol, petrol, paraffin oil, and sugar, do not conduct electricity.
- **Electrolysis** is the decomposition of an ionic compound either when molten or in aqueous solution by the passage of an electric current through it.

5.1.2.3 The conduction of electricity through molten or aqueous solutions of ionic compounds

Ionic compounds are made of a giant lattice of oppositely charged ions.

Ionic substances do not conduct electricity when solid since their ions cannot move.



When ionic substances are molten or dissolved in water, they conduct electricity and are decomposed at the same time. When molten or dissolved, the ions become free to move around and conduct electricity. These liquids or solutions are called electrolytes.

5.1.2.4 Comparison between a conductor and an electrolyte

The following table compares conductors with electrolytes.

Conductor	Electrolyte
No chemical change takes place during conduction of electricity.	A chemical change occurs at each electrode as the electrolyte conducts current.
Conductors conduct electricity due to the presence of **delocalised electrons**.	Electrolytes conduct electricity due to the presence of **mobile ions**.
Metals in the solid or liquid phase, or graphite.	Always in the liquid or aqueous phase.

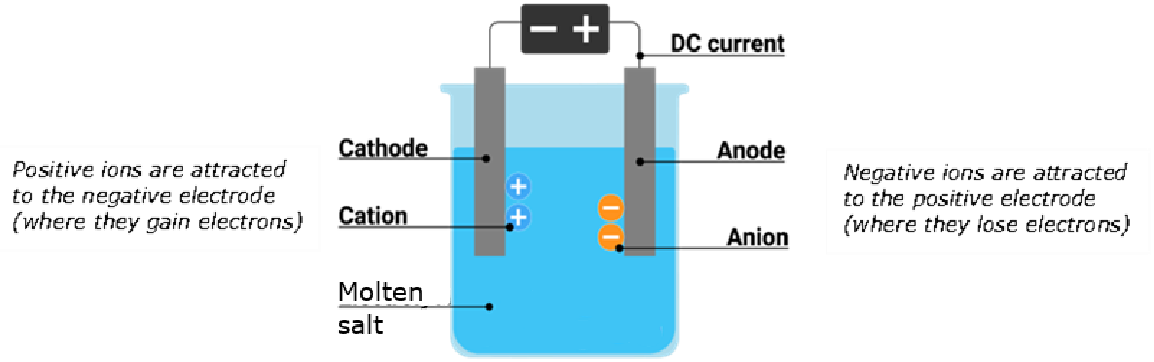
5.2 The electrolysis of solutions and molten salts

By the end of this section, you should be able to:

- Describe what happens when electricity is applied to molten ionic salts. (5.2c)
- Explain what happens when electricity is applied to molten ionic salts. *(E.g. lead(II) bromide)* (5.3c)
- Explain what happens when electricity is applied to solutions of salts. *(E.g. Electrolysis of dilute sulfuric acid, electrolysis of copper(II) sulfate solution using inert and active electrodes and electrolysis of concentrated sodium chloride solution.)* (5.3d)
- Describe electrolysis using half equations. (5.3e)
- Interpret electrolytic half equations in terms of oxidation and reduction. (5.3f)

5.2.1 Electrolysis of molten (fused) salts

Electrolysis is the decomposition of an ionic compound using electricity. Two **electrodes** are connected to a DC (direct current) supply and immersed into an electrolyte. The electrodes are usually made of graphite or platinum because they are inert, that is they do not react with the electrolyte. The electrode connected to the negative terminal of the DC supply is called the **cathode** and the electrode connected to the positive terminal of the DC supply is called the **anode**.

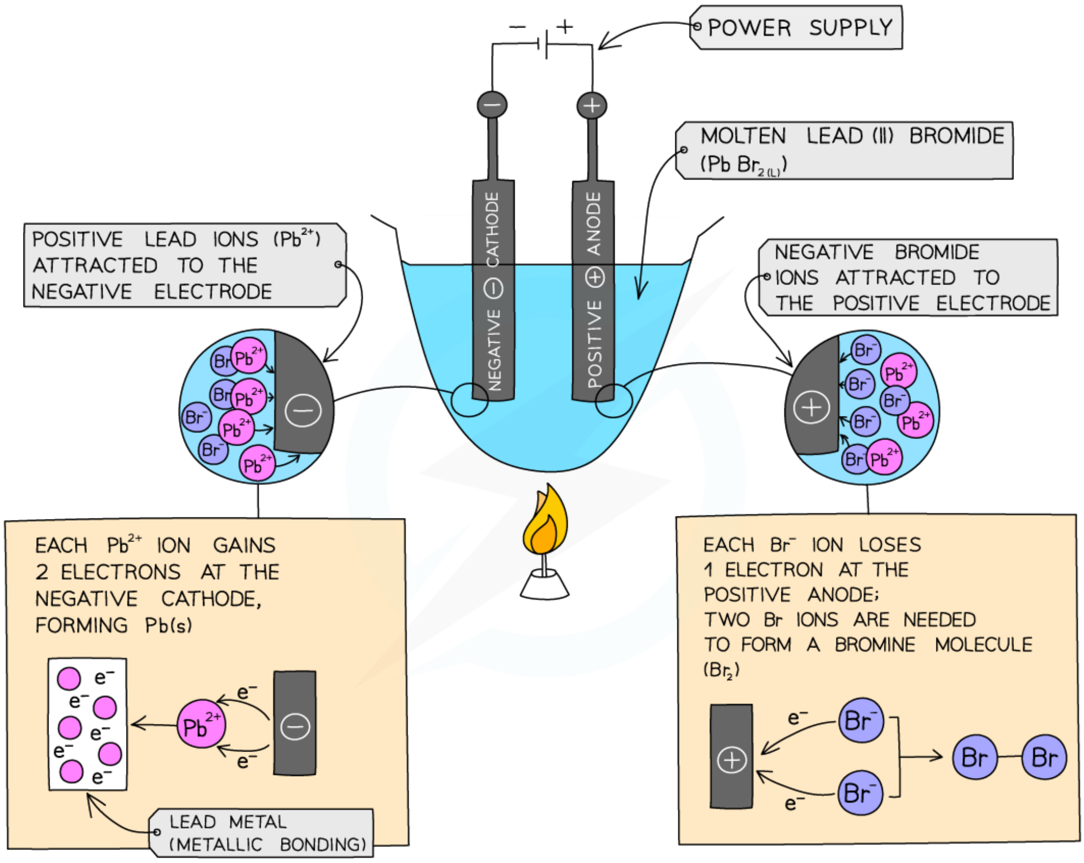


When the circuit is switched on, the **anions** (negative ions) are attracted to the positive electrode (anode) and the **cations** (positive ions) are attracted to the negative electrode (cathode). When the ions reach the electrodes, they are discharged, that is they lose their charge by gaining or losing electrons.

At cathode (-)	At anode (+)
The positive ions **gain** electrons at the cathode and become neutral atoms.	The negative ions **lose** electrons at the anode and become neutral atoms.
Reduction takes place at the cathode due to gain of electrons by the cation. $M^+ + e^- \rightarrow M$ The metal is formed as the product.	**Oxidation** takes place at the anode due to loss of electrons by the anion. $X^- \rightarrow X + e^-$ The non-metal is formed as the product.

5.2.1.1 Electrolysis of molten lead(II) bromide

During the electrolysis of molten lead(II) bromide, bromine vapour (reddish-brown) starts to bubble around the anode, while a bead of molten lead forms below the cathode.



When lead(II) bromide melts, lead(II) ions (Pb^{2+}) and bromide ions (Br^-) become free to move. Since opposite charges attract, the Pb^{2+} ions are attracted and move towards the cathode while the Br^- ions are attracted and move towards the anode.

At the cathode	At the anode
Lead ions are attracted to the cathode. Lead ions (Pb^{2+}) each gain 2 electrons and become lead atoms (Pb). We say that the lead ions are discharged. $\mathbf{Pb^{2+}(l) + 2e^- \rightarrow Pb(l)}$ Lead (Pb) atoms collect on the cathode and fall to the bottom of the crucible.	Bromide ions are attracted to the anode. Bromide ions (Br^-) each lose 1 electron to become bromine atoms (Br). We say that the bromide ions are discharged. $Br^- \rightarrow Br + e^-$ These pair together forming molecules: $\mathbf{2Br^-(l) \rightarrow Br_2\,(g) + 2e^-}$ Bromine (Br_2) bubbles off as a reddish-brown gas.
This is a **reduction** process where Pb^{2+} ions are reduced to Pb atoms due to **gain of electrons**.	This is an **oxidation** process where Br^- ions are oxidised to bromine molecules due to **loss of electrons**.

During electrolysis, each Pb^{2+} ion gains two electrons from the cathode while two Br^- ions each lose an electron to the anode. The number of electrons gained by the cations at the cathode **is the same** as the number of electrons lost by the anions at the anode.

The two equations representing the chemical changes at the electrodes are known as **half equations**. Together they give the complete chemical change that is occurring.

Half equation at the cathode. (Reduction)	$Pb^{2+}(l) + 2e^- \rightarrow Pb(l)$
Half equation at the anode. (Oxidation)	$2Br^-(l) \rightarrow Br_2(g) + 2e^-$
Adding both half equations.	$Pb^{2+}(l) + 2Br^-(l) \rightarrow Pb(l) + Br_2(g)$
Balanced chemical equation occurring in the electrolytic cell.	$PbBr_2(l) \rightarrow Pb(l) + Br_2(g)$

During electrolysis reduction (gain of electrons) always occurs at the cathode. Oxidation (loss of electrons) always occurs at the anode. Since **red**uction and **ox**idation occur simultaneously, this process is known as a **redox** reaction.

You could remember it as:

OIL RIG – Oxidation **I**s **L**oss of electrons, **R**eduction **I**s **G**ain of electrons.

One must note that, in conductors, electricity flows by the movement of electrons through a conductor while in electrolytes electricity flows by the movement of ions in an electrolyte resulting in discharge of the ions at the electrodes.

Summary

- When molten ionic compounds undergo electrolysis, the metal is produced at the cathode and the non-metal is produced at the anode.
- The passage of an electric current in the process of electrolysis results in the decomposition of the ionic compound into its constituent elements.

a) Explain why molten zinc chloride conducts electricity but zinc chloride crystals do not.

b) Complete the following table to explain what happens during the electrolysis of molten zinc chloride.

Substance	**Molten zinc chloride**
Ions present	
Observations at cathode	
Half equation for the reaction at cathode	
Product at cathode	
Observations at anode	
Half equations for the reaction at anode	
Product at anode	

Watch this video which explains the electrolysis of molten lead(II) bromide and zinc chloride.

5.2.1.2 Metal extraction

Reactive metals (such as K, Na, Ca, Mg and Al) are extracted from their molten compounds through electrolysis. Refer to the extraction of aluminium from molten alumina in section 6.3.2 on the extraction of aluminium.

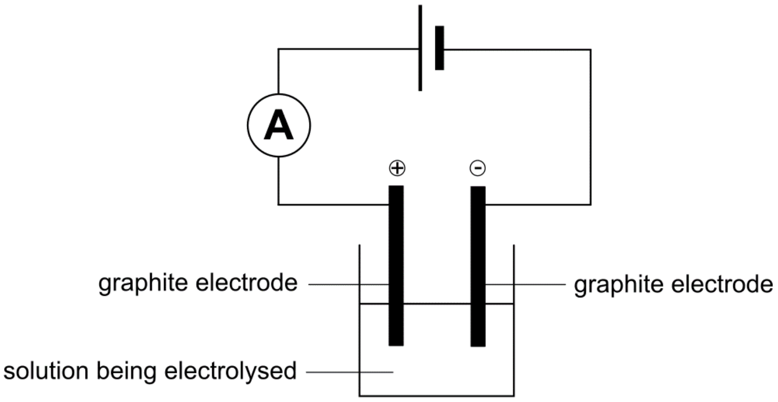
5.2.2 Electrolysis of salt solutions

When salts are dissolved in water, their ions move freely. Solutions of ionic compounds are electrolytes. The products of electrolysis of solutions of salts are often different from that of the molten salts.

In reality, the hydrogen ion (H^+) does not exist on its own in aqueous solutions. Since a **H^+ ion** is a **proton**, it bonds to water molecules to form the hydronium ion (H_3O^+). This ion is also known as the hydroxonium or the oxonium ion.

However, for the sake of simplicity, we still refer to the hydrogen ion and represent it in equations as H^+.

The following diagram shows a typical simple electrolytic cell (voltameter) that is used in a school laboratory to electrolyse solutions of salts.



As the solution is electrolysed, H^+ and OH^- ions from water compete with the ions from the dissolved salt for discharge at the electrodes. Three factors determine which ions are selectively discharged at the electrodes. These are:

1. The position of the ion in the electrochemical series (ECS).
2. The concentration of the ions in solution.
3. The nature of the electrodes.

5.2.2.1 Position of ions in the Electrochemical Series (ECS)

When two or more ions of similar charge are present under similar conditions e.g.: H^+ and Na^+ in NaCl(aq) one ion is **preferentially selected** for discharge. The selection of the ion discharged depends on its position in the **electrochemical series**.

Order of discharge at cathode	
Increasing Ease of Discharge ↓	Na^+
	Mg^{2+}
	Al^{3+}
	Zn^{2+}
	Fe^{2+}
	Pb^{2+}
	H^+
	Cu^{2+}
	Ag^+

Order of discharge at anode
1. For aqueous very dilute solutions OH^- is discharged.
2. For aqueous concentrated solutions containing halide ions (Cl^-, Br^- and I^-), these are discharged in preference to OH^-.
3. SO_4^{2-}, NO_3^- and CO_3^{2-} are never discharged from aqueous solutions.

Source: SEC 06 Syllabus (2025) Chemistry

At the cathode: The positive ions (cations) from the dissolved salt will compete with the H^+ ions from the water. The cation which is at the lowest position in the ECS will be selectively discharged.

At the anode: The negative ions (anions) from the dissolved salt will compete with the OH^- ions from the water. The anion which is at the lowest position in the ECS will be selectively discharged.

Electrolyte	Ions present	Cation discharged	Anion discharged
sodium sulfate solution	Na^+ SO_4^{2-} H^+ OH^-	H^+	OH^-
silver nitrate solution	Ag^+ NO_3^- H^+ OH^-	Ag^+	OH^-
dilute copper(II) chloride solution	Cu^{2+} Cl^- H^+ OH^-	Cu^{2+}	OH^-

Electrolysis of acidified water

Pure water is a very weak electrolyte due to a small number of water molecules that split into hydrogen (H^+) and hydroxide (OH^-) ions.

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

Water can be made to conduct electricity better if a strong electrolyte is added to it such as dilute sulfuric acid.

When passing an electric current through a solution of dilute sulfuric acid using inert electrodes (platinum or graphite), hydrogen and oxygen are obtained at the electrodes.

In a dilute solution of sulfuric acid, the following ions are present:

H^+ and SO_4^{2-} ions from sulfuric acid

H^+ and OH^- ions from water

As the cations and anions are discharged, more H^+ and OH^- ions are formed by the ionisation of water molecules. The total number of hydrogen and hydroxide ions present in solution at one time is more or less constant because the water molecules are continuously ionising to replace those which are discharged. H_2SO_4 is completely always ionised into ions.

At the cathode	At the anode
H^+ ions migrate to the cathode where they gain electrons to become hydrogen. $2H^+(aq) + 2e^- \rightarrow H_2(g)$ The number of electrons lost at the cathode and gained at the anode must be the equal. Therefore: $4H^+(aq) + \mathbf{4e^-} \rightarrow 2H_2(g)$	Both OH^- and SO_4^{2-} migrate to the anode. At the anode OH^- is selectively discharged in preference to SO_4^{2-}. (OH^- is lower in the ECS than SO_4^{2-}) $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + \mathbf{4e^-}$

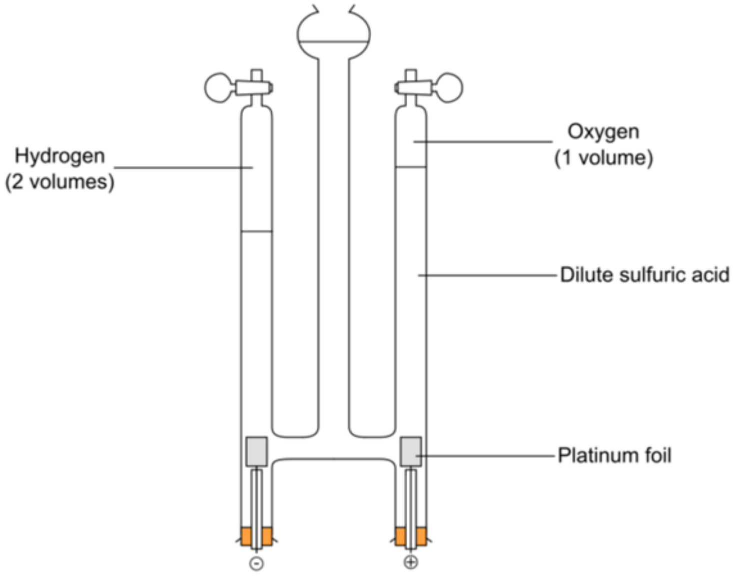
Overall result:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

Electrolysis of water can be carried out using a simple setup as shown earlier. However, a Hofmann voltameter shown here can be used in order to collect the gases produced.

This equipment offers the following advantages:

- Gases produced at the cathode and anode are collected separately.
- It allows for the volume of the gases produced to be measured.



Whilst using the Hofmann voltameter, one notes that 2 volumes of hydrogen are produced at the cathode while 1 volume of oxygen is produced at the anode. This observation reflects the mole ratio of hydrogen and oxygen in the balanced chemical equation. Also note that for the same number of electrons, the volume of hydrogen collected is double that of oxygen.

Note that oxygen is more soluble in water (0.0043 g/100 g of water at 20 °C, 1 atm) than hydrogen (0.00016 g/100 g of water at 20 °C, 1 atm) hence less oxygen may be collected at first.

Watch this video about the electrolysis of acidified water.

5.2.2.2 Concentration of ions in solution

Electrolysis of dilute sodium chloride solution

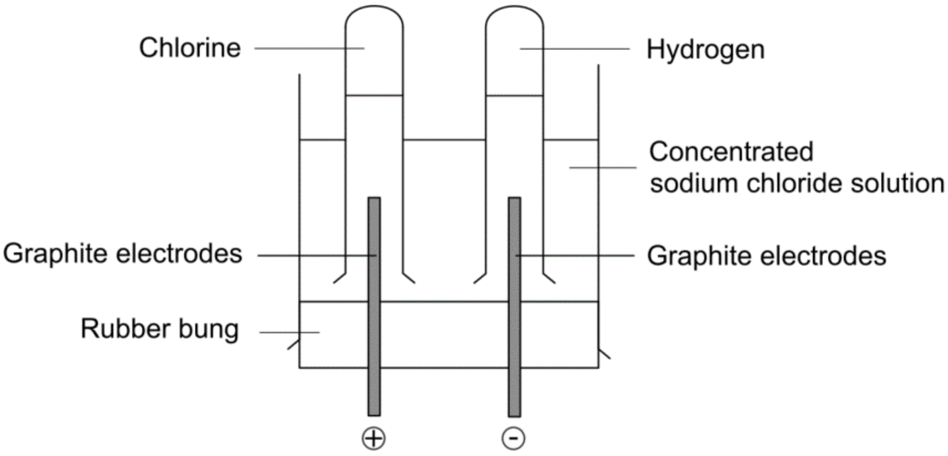
In a **very dilute** solution of sodium chloride there are four ions present. These are Na^+, Cl^-, H^+, and OH^-.

	At the cathode	At the anode
Ion discharge	Both Na^+ and H^+ migrate to the cathode. **H^+** ions readily accept electrons being lower in the ECS).	Both Cl^- and OH^- ions migrate to the anode. **OH^-** ions readily give electrons being lower in ECS.
Half equation	$4H^+(aq) + 4e^- \rightarrow 2H_2$	$4\ OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$
Observation	Colourless gas bubbles at cathode	Colourless gas bubbles at the anode

What would happen if a **concentrated solution** of sodium chloride undergoes electrolysis?

Electrolysis of concentrated sodium chloride solution

A concentrated solution of sodium chloride is electrolysed using graphite or platinum electrodes.



In concentrated sodium chloride solution, there are four ions present: Na^+, Cl^-, H^+, and OH^-.

Equal volumes of hydrogen and chlorine gas are formed. Why does this happen?

	At the cathode	At the anode
Ion discharge	Both Na$^+$ and H$^+$ migrate to the cathode. H$^+$ ions readily accept electrons than Na$^+$ ions (being lower in the ECS).	Both Cl$^-$ and OH$^-$ ions migrate to the anode. Cl$^-$ ions are discharged because they are present in a higher concentration than OH$^-$.
Half equation	$2H^+(aq) + 2e^- \rightarrow H_2$	$2Cl^-(aq) \rightarrow Cl_2 + 2e^-$
Observation	Colourless gas bubbles at cathode.	Greenish yellow gas bubbles* which turn damp blue litmus red and then bleached.

*N.B. The solubility of chlorine in water is 0.729 g/100 g of water at 20 °C and 1 atm. Therefore, the volume of chlorine collected during the first stages of the experiment is smaller than that of hydrogen.

Na$^+$ and OH$^-$ ions are left in solution so a sodium hydroxide solution forms. Therefore, the solution becomes increasingly alkaline as electrolysis proceeds.

Three important products are produced from the electrolysis of concentrated sodium chloride which include hydrogen and chlorine gases, and sodium hydroxide solution.

In the electrolysis of a concentrated halide solution the halide ions (Cl$^-$, Br$^-$ and I$^-$) are present in **greater concentrations** and they are discharged producing the halogen at the anode.

The sulfate, nitrate, and carbonate ions are never discharged from aqueous solutions even at very high concentrations.

Watch this video of the electrolysis of concentrated sodium chloride in the lab.

5.2.2.3 Nature of the electrodes

During electrolysis the nature of the electrode may determine which ions are discharged.

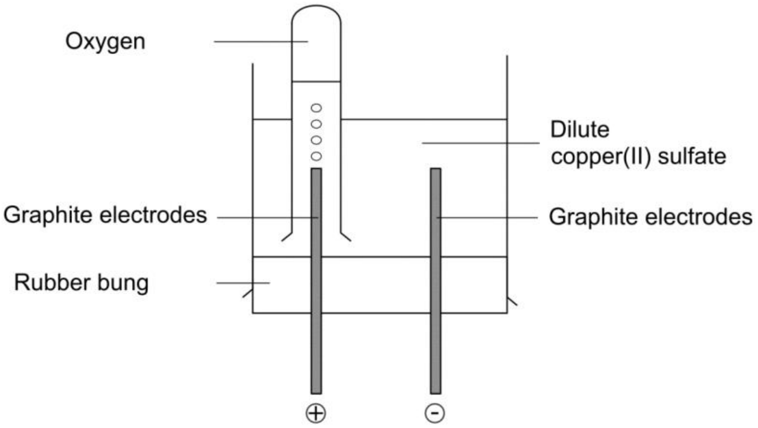
Usually, electrodes are made from materials that do not interfere chemically with the process. Graphite or platinum are used in such cases as they are quite unreactive. Such electrodes are called **inert** electrodes.

Sometimes, other materials are used which may affect the outcome of electrolysis. Electrodes in such cases are known as **active** electrodes.

Electrolysis of copper(II) sulfate solution – using graphite electrodes

A solution of copper(II) sulfate contains the following ions: Cu^{2+}, SO_4^{2-}, H^+, and OH^- ions.

	At the cathode	At the anode
Ion discharge	Both Cu^{2+} and H^+ ions migrate to cathode but Cu^{2+} is preferentially discharged being lower in the ECS.	Both SO_4^{2-} and OH^- ions migrate to anode but OH^- is preferentially discharged being lower in the ECS.
Half equation	$2Cu^{2+}(aq) + 4e^- \rightarrow 2Cu\ (s)$	$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$
Observation	A pink solid deposits at the cathode which oxidises rapidly and turns reddish brown.	A colourless gas which relights a glowing splint is evolved.

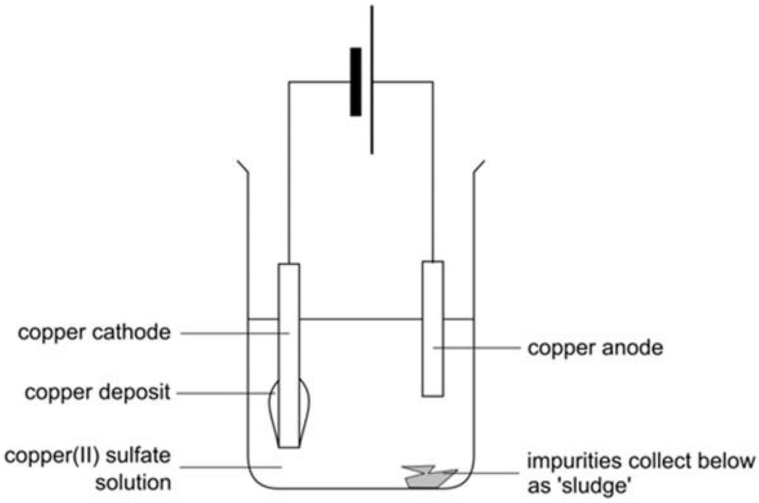


Note that:

- When copper(II) sulfate solution is electrolysed using graphite electrodes, copper and oxygen are obtained.
- The blue solution fades due to the decrease of the concentration of Cu^{2+} ions in solution.
- The solution turns acidic due to the increased concentration of H^+ ions in solution. This comes about due to the discharge of OH^- ions from the solution to produce oxygen gas.

Electrolysis of copper(II) sulfate solution – using copper electrodes

Copper(II) sulfate solution is electrolysed using copper electrodes. The copper anode is known as the **active** electrode because it is made of the **same metal as the metal ion in solution**. The table below shows what happens in this case.



	At the cathode	At the anode
Ion discharge	Copper ions migrate to the cathode, pick up electrons and become copper atoms.	The copper anode loses electrons, to form copper ions in solution.
Half equation	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Observation	A pink solid deposits at the cathode. Cathode grows thicker, as copper is deposited.	Anode wears away and loses mass.

Note that:

- The mass of the copper from the anode decreases as it goes into the copper(II) sulfate solution as Cu^{2+} ions. These ions are discharged and deposited on the cathode.
- $CuSO_4$ solution remains blue as the concentration of Cu^{2+} (aq) does not change. For every Cu^{2+} ion that is discharged at cathode, another Cu^{2+} ion goes into solution from the anode.
- Mass of copper lost by anode = mass of copper gained by cathode.

Watch this video about the electrolysis of copper(II) sulfate using inert and copper electrodes.

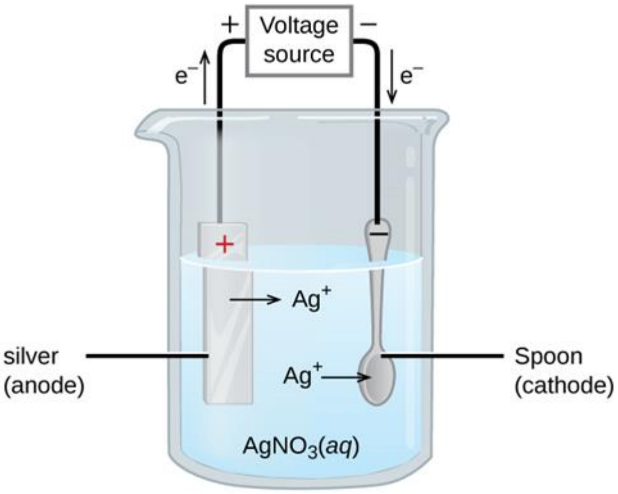
5.2.3 Electroplating

Electroplating is the process by which a metal is deposited on a conductive surface. Electroplating has been used to decorate objects by giving them a gold or silver finish. It is also used in modern technology e.g., corrosion inhibition, production of electronic circuitry and nanotechnology.

Electroplating is often used to coat one metal with another, usually to make the metal look more attractive or to prevent it from rusting.

The diagram shows how an object can be electroplated with silver (silver plating). During this process:

- The **object** to be electroplated is put as the **cathode**.
- The **anode** is made of pure **silver**.
- The electrolyte solution must contain soluble **silver ions**.



At the anode:
Silver loses electrons to form silver ions in solution:

$$Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$$

At the cathode:
Silver ions gain electrons at the cathode to form a layer of silver on the object:

$$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$$

When the layer of silver is thick enough the object is removed.

Uses of electroplating:

1. Electroplating is used in industry e.g., sink faucets are electroplated with chromium to make them shiny and protect from rusting.

2. Steel is electroplated with tin to make food cans.

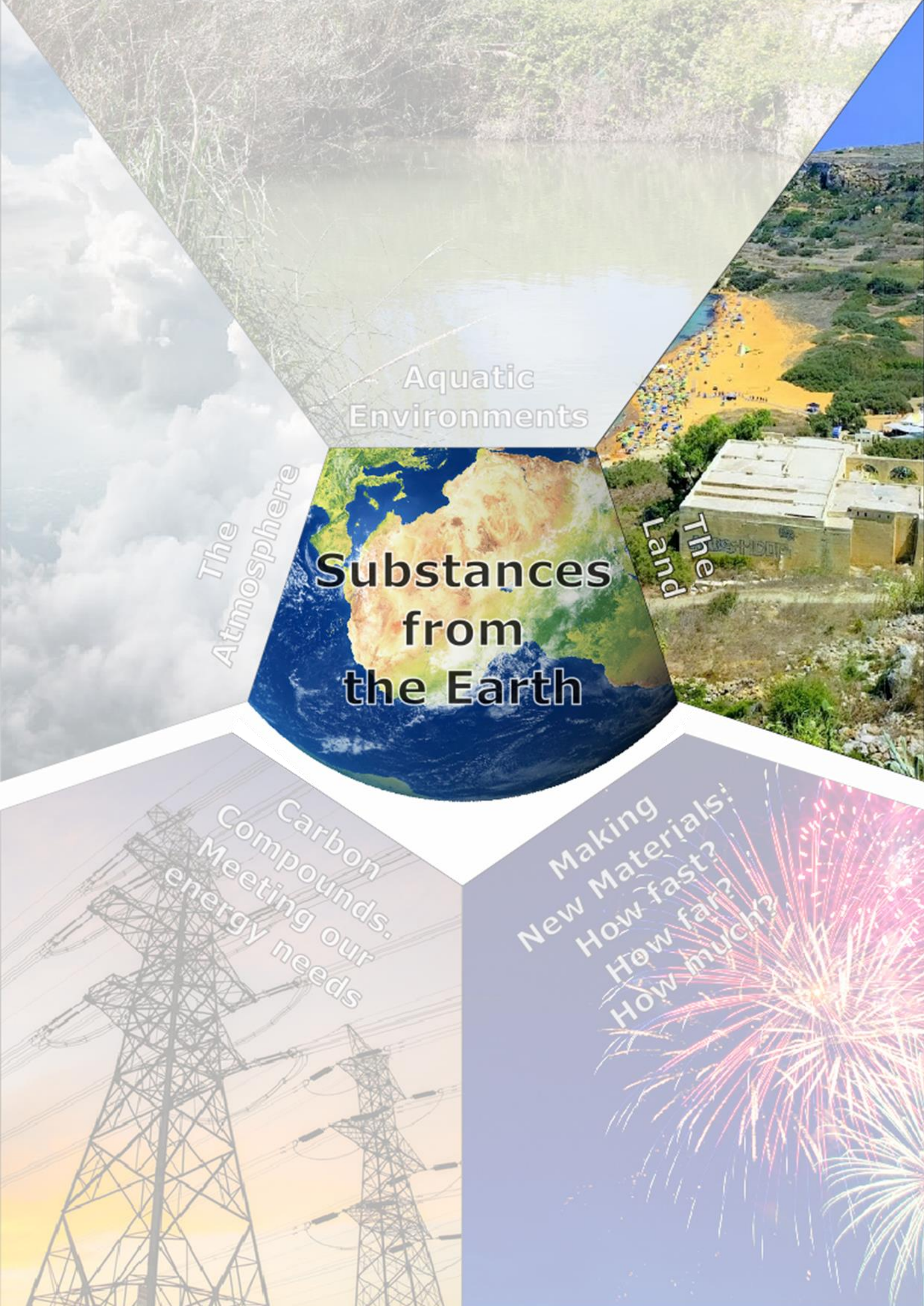
Source: www.euroboxpackaging.com

The future of electrochemistry

Electrochemistry holds huge potential for the technological advances in various areas such as hydrogen fuel cells and the hydrogen economy. Some researchers are looking into producing hydrogen from air moisture.

Anne McNeil is developing ways of extracting chlorine from PVC. Other scientists are looking into producing more environmentally friendly batteries, better ways of recycling catalysts, etc. The opportunities are endless.

Further reading on the industrial applications of electroplating.



Substances from the Earth

Aquatic Environments

The Atmosphere

The Land

Carbon Compounds. Meeting our energy needs

Making New Materials: How fast? How far? How much?

6 Substances found in rocks, their extraction, chemical nature, responsible use, and environmental impact – LO 8

This chapter explores the main materials extracted from the land around us such as limestone and other building materials, metals and their compounds, and various carbon substances.

6.1 Limestone and other building materials

6.1.1 Limestone and its uses

By the end of this section, you should be able to:

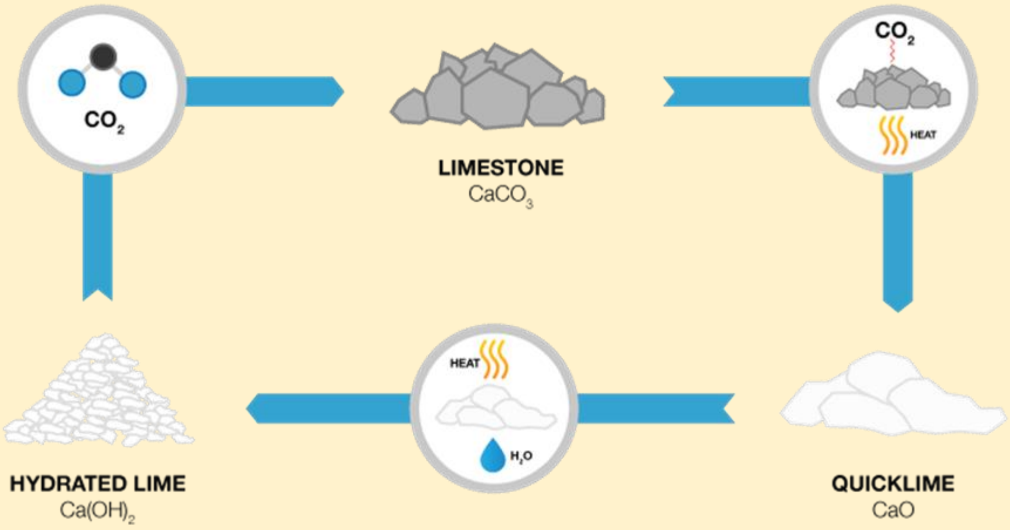
- State uses of limestone. (8.1a)
- Describe the use of limestone in industry. *(Including the manufacture of quicklime and slaked lime. As an aggregate in construction.)* (8.2a)

Limestone is a sedimentary rock which mainly consists of calcium carbonate.

Limestone is used:

- to make cement, in the foundation of roads, and in buildings.
- in the manufacture of quicklime (calcium oxide) and slaked lime (calcium hydroxide).
- in the construction industry for example as an aggregate (a mixture of sand, gravel, and crushed rock) in concrete.

The Lime Cycle



Limestone turns to quicklime (calcium oxide), when heated to 1000 °C in a lime kiln (a high temperature oven). This is a type of reaction called **thermal decomposition.**

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

Quicklime can be hydrated, that is, it combines with water. Hydrated lime known as slaked lime is produced according to the reaction below which is very exothermic.

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(s)$$

Slaked lime reacts with carbon dioxide to produce calcium carbonate as shown below:

$$Ca(OH)_2(s) + CO_2(g) \longrightarrow CaCO_3(s)$$

6.1.2 Investigating the physical and chemical properties of building materials

By the end of this section, you should be able to:

- Investigate simple physical properties of substances used in buildings and relate them to their use. *(E.g. density, heat and electrical conductivity of limestone, concrete, wood, steel, and aluminium.)* (8.2b)
- Investigate the chemical properties of substances used in buildings and relate them to their use. *(Limited to action of acids and water on limestone, concrete, wood, steel, and aluminium)* (8.3b)

In conjunction with limestone in construction, other materials are widely used such as concrete, wood, steel, and aluminium.

6.1.2.1 Physical properties of building materials

6.1.2.1.1 Limestone

- Limestone is made up of tiny fossils, shell fragments and other fossilized debris.
- It is present in rock and can have different colorations. In Malta limestone is mainly yellowish white but other colorations may be observed.
- It is a relatively soft and dense rock when compared to marble and granite (which are both harder and denser than limestone).
- Limestone is a porous material and allows liquids such as water to flow through it.
- It is susceptible to deterioration when it comes in contact with acidic sources of water.
- Limestone is a poor conductor of heat and electricity.
- It is fire resistant. However, exposure to very high temperatures makes it decompose and lose its structural integrity.

6.1.2.1.2 Concrete

Concrete is the major building material in the construction industry. Concrete is used extensively in dams, residential and commercial buildings, and road building.

Sand, gravel, and crushed stones are used in concrete. The properties of concrete depend on the mix proportion of fine and coarse aggregate of sand and gravel (a form of limestone) bonded together with cement and water. Cement acts as a binding material which forms a paste with water and on hardening, this paste binds together with the aggregate to form a solid mass.

To increase its strength concrete is reinforced with steel structures which are embedded within it.


By Joe Kirkwood and Holly Kirkwood

The picture shows concrete being poured into a steel mesh which strenghtens the end product.

Concrete is a:

- building material that gains strength over time as it dries. It is durable and resists weathering.
- non-conductor of electricity if the steel structure within is not exposed.
- better insulator of heat when compared to metals.
- fire resistant.

6.1.2.1.3 Wood

Wood is:

- a natural resource that is obtained from trees.
- a better insulator of heat than limestone and concrete.
- a flammable material that burns easily when exposed to flames.
- an electrical insulator but as moisture content increases electrical conductivity increases.

6.1.2.1.4 Steel

Steel is made from iron by removing impurities and adding substances such as manganese, chromium, nickel, tungsten, molybdenum, and vanadium to produce alloys with properties that make the material suitable for specific uses.

Most steels also contain small but definite percentages of carbon (0.04%–2.5%). Steel has a density of 7,850 kg/m^3, making it 7.85 times as dense as water.

Its melting point may vary from 1300 to 1540 °C and is higher than that of most metals. The physical properties of steel include high strength, relatively low weight, durability, ductility, and resistance to corrosion.

The American Iron and Steel Institute (AISI) puts steel into 4 major types (i) carbon steel, (ii) alloy steel, (iii) stainless steel; (iv) tool steel.
The above four steel types are further subdivided into a number of categories based on varying chemical composition.

6.1.2.1.5 Aluminium

Aluminium compounds such as bauxite are very abundant in the Earth's crust. Aluminium is extracted from its ore by the electrolysis of alumina (Al_2O_3).

Aluminium is a:

- relatively soft, durable, lightweight, ductile, and malleable metal.
- good thermal and electrical conductor.
- fairly corrosion resistant material due to the oxide layer that it forms.
- recyclable material.
- relatively light material because it has a relatively low density.

As early as the 1900s aluminium has been used in the construction industry. However, it was not until the 1920s that the electrolytic process reduced the cost of production and aluminium became widely used in the construction industry. The Empire State building utilized aluminium in all its basic structures and even on its famous Spire. Today, nearly a quarter of all aluminium produced in the world is used in the construction industry.

6.1.2.2 Chemical properties of building materials

6.1.2.2.1 Limestone

Limestone reacts with acids.

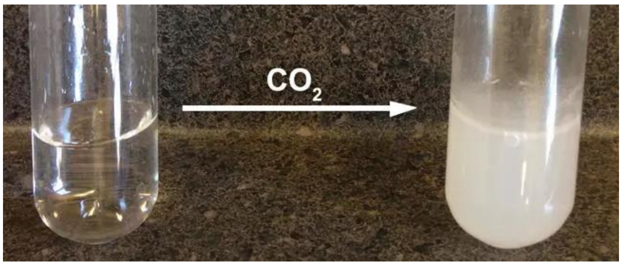
To investigate this chemical property of limestone you can carry out the following simple experiment.

(a) Put some calcium carbonate (marble chips) in a dilute hydrochloric acid.

(b) What do you observe?

(c) Test the gas evolved with damp blue litmus paper.

(d) Bubble the gas through freshly prepared lime water (calcium hydroxide solution).



The picture shows the effect carbon dioxide has on bubbling it through lime water.
Source: https://bit.ly/41loasl

Observations and results

Effervescence is observed, meaning that a gas is evolved. The damp blue litmus paper turns red showing that the gas is acidic. Lime water turns cloudy which shows that carbon dioxide is given off.

When carbonates react with acids, a salt, carbon dioxide, and water are given off.

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

Since rain is slightly acidic by nature, weathering of limestone happens as a natural process. Acid rain speeds up this process since the pH of acid rain is below 4.5.

6.1.2.2.2 Concrete

Concrete is a composite material made up of a filler (the aggregate) and a binder (the cement). As water is added to the mixture, the cement “glues” the aggregate together to form concrete. Cement is made by burning limestone and clay together at very high temperatures ranging from 1400 to 1600 °C.

Concrete is impermeable to pure water. However, as soon as water becomes acidic, the alkaline component of cement in concrete reacts with it. Concrete deterioration increases as the pH of the acid decreases.

6.1.2.2.3 Aluminium

Aluminium reacts with air forming a layer of aluminium oxide. This layer acts as a protective layer reducing the chemical reactivity of aluminium.

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

When aluminium is placed in a cold and dilute, strong / weak acid, the oxide layer must be first removed for the reaction to proceed. If the reacting acid is moderately concentrated and the reaction mixture heated to several hundred degrees Celsius (using Bunsen flame heating) no prior removal of the oxide layer is required since the oxide surface coated aluminium reacts directly with the acid under these more aggressive reaction conditions.

$$2Al(s) + 6HCl(g) \longrightarrow 2AlCl_3(s) + 3H_2(g)$$

Aluminium does not react with water at room temperature but will react with steam.

$$2Al(s) + 3H_2O(g) \longrightarrow Al_2O_3(s) + 3H_2(g)$$

6.1.2.2.4 Steel

Steel is mostly an alloy of iron and carbon. An alloy is made by combining two or more elements, usually metals, but also non-metals such as carbon in the case of steel. Alloys are formed to enhance the metal's properties such as strength and increased resistance to corrosion.

Both oxygen and water vapour present in the atmosphere react with steel to form rust which is a hydrated form of iron(III) oxide. Steel can also be corroded by reaction with acids.

There are many variations of steel depending on the mixture of elements that are added to the iron. One example is stainless steel which is generally more resistant to corrosion.

6.1.2.2.5 Wood

Wood shrinks leading to cracks developing during dry seasons due to a loss in moisture. On the other hand, wood expands during hot and humid summer conditions. Heavier, denser wood will expand and contract less.

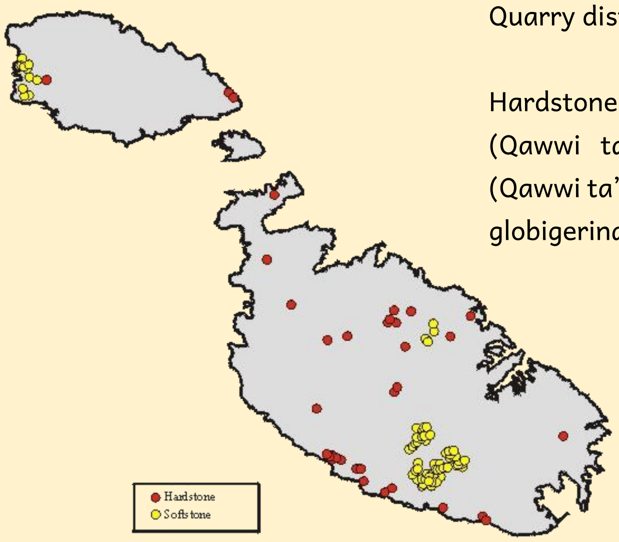
Acid deposition, for example from acid rain, increases the rate of deterioration of unpainted wood. The acid reacts with the cellulose which makes up wood and which gives structural integrity to the material. Wood should be painted as soon as possible after it is installed outdoors. Acid deposition does not affect wood that has been painted before exposure.

6.1.3 The economic and environmental impact of open quarrying of stone

By the end of this section, you should be able to:

- Describe the economic and environmental impact of open quarrying of stone. (8.2c)
- Debate the economic and environmental impact of open quarrying of stone. (8.3c)

Stone quarries in Malta



Quarry distribution in Malta and Gozo, 2012.

Hardstone refers to upper coralline limestone (Qawwi ta' fuq) and lower coralline limestone (Qawwi ta' taħt / Żonqor) whilst soft stone refers to globigerina limestone (Tal-Franka).

As time goes by the use of globigerina limestone is being swapped for concrete blocks.

6.1.3.1 Advantages of limestone quarrying

- Readily available cheap resource.
- Creates job opportunities.
- Decreases the use of wood and so saves on cutting trees.
- Quarrying provides materials (in Malta these include stone slabs, aggregate, and sand) for building new homes which are in huge demand.
- Quarrying provides income to local councils in some countries.

6.1.3.2 Disadvantages of limestone quarrying

- Agricultural land and wildlife habitats are destroyed.
- Quarrying creates pollution from noise and dust.
- Heavy trucks and other vehicles cause pollution and congestion on narrow country roads. The vibrations from heavy traffic can cause damage to nearby buildings.
- Quarries create visual pollution to the landscape.
- Limestone is a non-renewable resource - so it can be argued that quarrying is unsustainable.
- Land may subside and rock strata can be destabilised due to the effects of the quarry digging/blasting activities.

ReStone – Recycling Stone

Currently limestone is a non-renewable resource, however researchers at the University of Malta, Prof Spiridione Buhagiar and architect Franco Montesin have been working on recycling limestone. In 2021 the first room made of crushed recycled quarry waste was constructed near Ta' Qali Crafts Village.

6.2 Metals and their reactivity

6.2.1 Metal sources in the Earth's crust

By the end of this section, you should be able to:

- Identify metals that are found free in nature or that are extracted from certain minerals found in rocks. (Limited to iron from haematite and aluminium from bauxite as well as the very few metals found as elements in the ground e.g. gold and platinum.) (8.1d)

Metals are found as ores which are rocks found in the Earth's crust which contain free metals and/or metal compounds.

Metals which are chemically unreactive are found in their free (native) state in the Earth's crust. Examples include gold and platinum.

More reactive metals are naturally found combined with other elements. They have to be chemically extracted from compounds found in ores. Examples include iron from haematite and aluminium from bauxite.

Source: www.lincroftvillagejewelers.com

The method of extraction of metals depends on their reactivity as discussed in section 6.3.

Further reading about metals which are found free in nature.

6.2.2 The reactivity series

The following is a list of metals in decreasing order of reactivity. It is called the **Reactivity Series**.

Reactivity series	
Decreasing Reactivity ↓	Potassium
	Sodium
	Calcium
	Magnesium
	Aluminium
	Carbon
	Zinc
	Iron
	Lead
	Hydrogen
	Copper
	Silver
	Gold
	Platinum

Source: SEC 2025 syllabus

***** Carbon and hydrogen are not metals and are included as reference points to determine whether particular chemical reactions can take place.

The order of reactivity of the metals can be determined by reacting the metals with other substances such as oxygen, water/steam, and hydrochloric acid as well as by means of displacement reactions as shown in the following sections.

6.2.3 The reaction of metals with oxygen

By the end of this section, you should be able to:

- Relate metals' position in the reactivity series to their ease of corrosion and extraction. *(Metals limited to potassium, sodium, calcium, magnesium, aluminium, zinc, iron, lead, copper, silver, gold, and platinum.)* (8.2i)

The higher the metal is in the reactivity series the more easily it reacts with oxygen when it is exposed to air at room temperature.

Metal	**Reaction with oxygen**	**Observation**
Potassium	$4K(s) + O_2(g) \rightarrow 2K_2O(s)$	Instantaneous reaction to produce the oxide, rapidly tarnishes when exposed to air. Potassium tarnishes faster than sodium.
Sodium	$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$	
Calcium	$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$	Slower reaction with oxygen to form the oxide when exposed to air. Calcium reacts faster with oxygen than magnesium.
Magnesium	$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$	
Aluminium	$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$	Forms an oxide layer which prevents further reaction.
Zinc	$2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$	Forms an oxide slowly.
Iron	$2Fe(s) + 3O_2(g) + nH_20(l) \rightarrow Fe_2O_3.nH_2O(s)$	At room temperature, iron reacts with both oxygen and water to form rust.
Lead	$2Pb(s) + O_2(g) \rightarrow 2PbO(s)$	Forms an oxide layer slowly.
Copper	$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$	Forms an oxide layer very slowly.
Silver	Reacts with oxygen in air when heated.	Found free in nature.
Gold	Does not react with oxygen.	
Platinum		

Note that as one goes down the reactivity series of metals their **ease of corrosion** decreases.

6.2.4 The reaction of metals with water and steam

By the end of this section, you should be able to:

- Determine metals' position in the reactivity series from their reactions with water/steam. *(Metals limited to calcium, magnesium, aluminium, zinc, iron, lead, and copper. Represent reactions using balanced chemical equations.)* (8.2j part)

Highly reactive metals such as potassium and sodium react violently with cold water to form the metal hydroxide and hydrogen gas. A more reactive metal will react more vigorously with cold water.

Some metals such as zinc and iron, do not react with cold water but react with steam to form the metal oxide and hydrogen gas.

The following table lists the metals in order of decreasing reactivity as shown by their decrease in reactivity with water/steam.

Metal	Reaction with water or steam	Observation
Calcium	$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$	Reacts vigorously.
Magnesium	$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$ $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$	Very slow reaction with water. Violent reaction with steam.
Aluminium	$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$	Reacts readily with steam but not with cold water. Aluminium reacts faster than zinc.
Zinc	$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$	
Iron	$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$	Reacts slowly with steam.
Lead	No reaction with water or steam.	
Copper		
Silver		
Gold		
Platinum		

6.2.5 The reaction of metals with hydrochloric acid

By the end of this section, you should be able to:

- Determine metals' position in the reactivity series from their reactions with hydrochloric acid. *(Metals limited to calcium, magnesium, aluminium, zinc, iron, lead, and copper. Represent reactions using balanced chemical equations.)* (8.2j part)

Most metals react with a dilute, non-oxidising acid to form a salt and hydrogen gas.

Potassium and sodium react explosively with dilute acids. Potassium is more reactive than sodium. **Never** attempt these reactions in the laboratory.

Metal	**Reaction with hydrochloric acid**	**Observation**
Calcium	$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$	Reacts violently.
Magnesium	$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$	Reacts rapidly. Magnesium reacts faster than aluminium.
Aluminium	$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$	
Zinc	$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	Reacts moderately.
Iron	$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$	Reacts slowly.
Lead	$Pb(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2(g)$	Reacts very slowly and forms a protective layer that stops the reaction.
Copper	No reaction.	
Silver		
Gold		
Platinum		

Note that lead should react with dilute acids since it is higher than hydrogen in the Reactivity Series. The initial reaction between lead and hydrochloric acid will form an insoluble layer of lead(II) chloride. This prevents further reaction of the hydrochloric acid with the underlying lead metal. The reaction slows down and eventually stops. Hence, lead does not appear to react with hydrochloric acid.

A similar reaction occurs when sulfuric acid is used. Lead does not appear to react with sulfuric acid because of the insoluble layer of lead(II) sulfate coated on the underlying lead metal.

The chemical equation for the reaction is:

$$Pb(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + H_2(g)$$

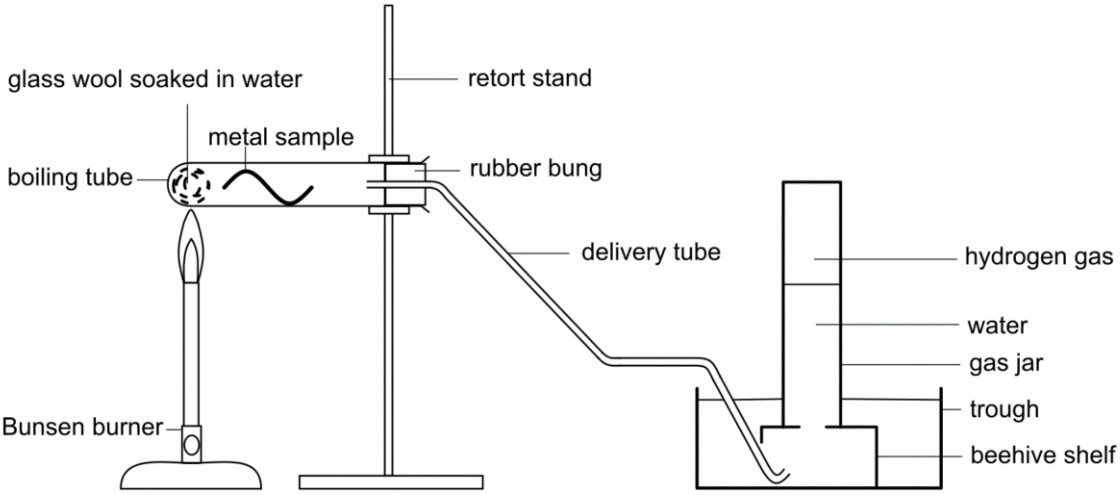
Reactivity Series of Metals

6.2.6 Determining the position of an unknown metal with respect to other metals in the reactivity series

By the end of this section, you should be able to:

- Determine the position of an unknown metal (e.g. tin) with respect to other metals in the reactivity series from their reactions with water/steam and hydrochloric acid. *(Other metals limited to magnesium, aluminium, zinc, iron, lead, and copper. Represent reactions using balanced chemical equations.)* (8.3j)

The position of tin (Sn) in the reactivity series is determined by observing its reaction with water and hydrochloric acid.



1. Some pieces of tin are placed in a test tube with water and any reaction is observed.
2. When no reaction with water is observed, the apparatus above is used to observe its reaction with steam.
3. Then, tin is placed in dilute hydrochloric acid and its reaction is observed.

No reaction is observed when tin is placed in water, but tin reacts readily with steam.

$$Sn(s) + H_2O(l) \rightarrow \text{no reaction}$$

$$Sn(s) + H_2O(g) \rightarrow SnO(s) + H_2(g)$$

Tin reacts slowly with dilute hydrochloric acid.

$$Sn(s) + 2HCl(aq) \rightarrow SnCl_2(aq) + H_2(g)$$

These results put tin between iron and lead in the reactivity series.

6.2.7 Displacement reactions

By the end of this section, you should be able to:

- Determine metals' position in the reactivity series from displacement reactions. *(Metals limited to calcium, magnesium, aluminium, zinc, iron, lead, and copper. Represent reactions using balanced chemical equations.)* (8.3k part)

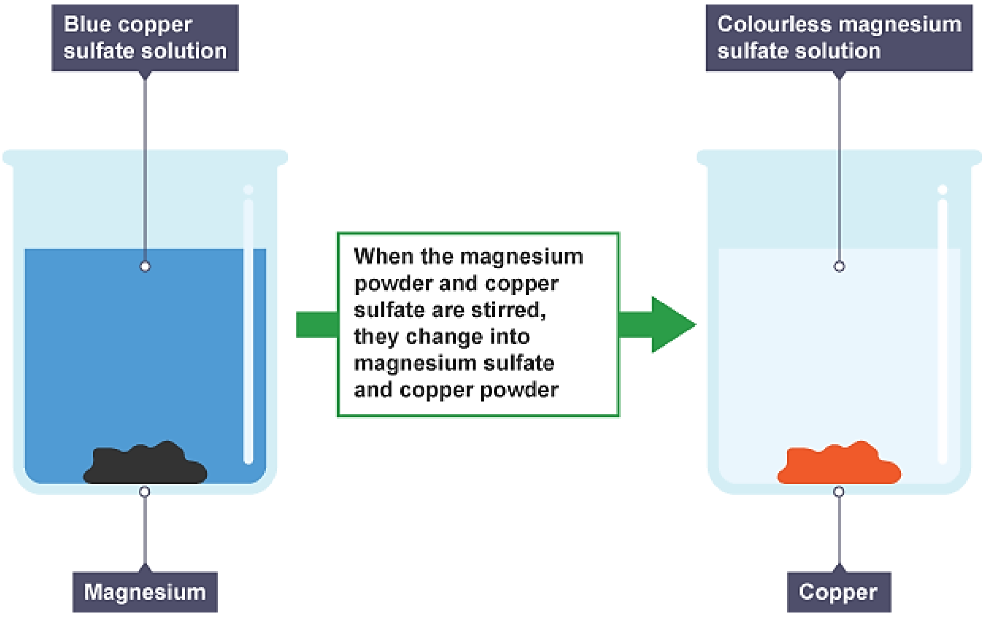
Displacement reactions involve a metal and the compound of a different metal.

A more reactive metal will displace a less reactive metal from its compound in a displacement reaction.

The less reactive metal is left uncombined after the reaction.

For example, magnesium is more reactive than copper. When a piece of magnesium is dipped into blue copper(II) sulfate solution, a displacement reaction occurs.

$$Mg(s) + CuSO_4 (aq) \rightarrow Cu(s) + MgSO_4 (aq)$$



Source: BBC Bitesize

This reaction cannot happen in reverse. Copper metal cannot displace magnesium ions from a compound because copper is less reactive than magnesium.

When observing displacement reactions, we look for a colour change or change in temperature as evidence that a reaction is happening. A metal deposit is produced.

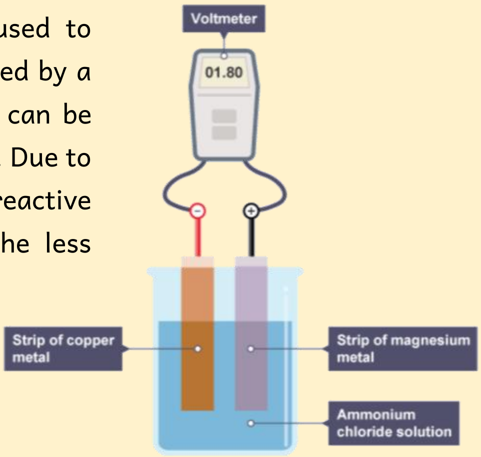
A simulation on investigating the displacement of metals.

Infographic about the **Tin Hedgehog** experiment.

Batteries

The difference in reactivity of two metals can be used to produce batteries. If two different metals are connected by a wire and are inserted in an electrolyte, a simple cell can be made. Several cells together can in turn form a battery. Due to the difference in reactivity, electrons from the more reactive metal (e.g. magnesium) flow through the wires to the less reactive metal (e.g. copper).

Positively charged ions flow from the copper through the electrolyte to the magnesium. Electricity will continue to flow until there is no more magnesium left to react.



6.3 Metals and their extraction

By the end of this section, you should be able to:

- Use the reactivity series of metals to predict the best method of metal extraction by reduction with carbon or electrolysis. (8.3m)

The method of extraction of metals depends on the position of the metal with respect to carbon in the reactivity series as shown in the diagram below.

Potassium **Most reactive**
Sodium
Calcium
Magnesium
Aluminium
Carbon *
Zinc
Iron
Lead
Copper
Silver
Gold
Platinum **Least reactive**

* Carbon is not a metal. It is being included in this list because it indicates a reference point of the method of extraction that can be used.

Metals, such as aluminium, that are above carbon in the reactivity series are extracted by electrolysis.

Metals that are lower than carbon in the reactivity series are extracted by a process called reduction using carbon or carbon monoxide.

Metals like gold and platinum, are unreactive and, as mentioned previously, are found free in nature.

6.3.1 Iron

By the end of this section, you should be able to:

- Describe the essential chemical reactions and conditions in the industrial extraction of iron in the blast furnace. *(Drawing of diagrams and technical details are not required.)* (8.3d part)

Source: www.europeanheritageawards.eu

Iron is not found free in nature. It is found combined with many other elements to form many substances. The most common ore from which iron is extracted is called haematite (iron(III) oxide, Fe_2O_3). The industrial plant which is used to extract iron from haematite is known as a blast furnace.

Coke, limestone, and haematite are fed into the top part of the blast furnace while blasts of hot air are shot into the sides of the furnace.

A series of chemical reactions occur in the blast furnace.

1. The coke (carbon) burns with oxygen to produce carbon dioxide. This reaction is exothermic (gives out heat).

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

2. The CO_2 then reacts with more coke to give the gas carbon monoxide.

$$CO_2(g) + C(s) \rightarrow 2CO(g)$$

3. Carbon monoxide reacts with the iron ore (Fe_2O_3) to produce iron.

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$$

Molten iron trickles to the bottom of the furnace where it is collected. Silicon dioxide is the main impurity in the ore which is removed in the blast furnace as follows:

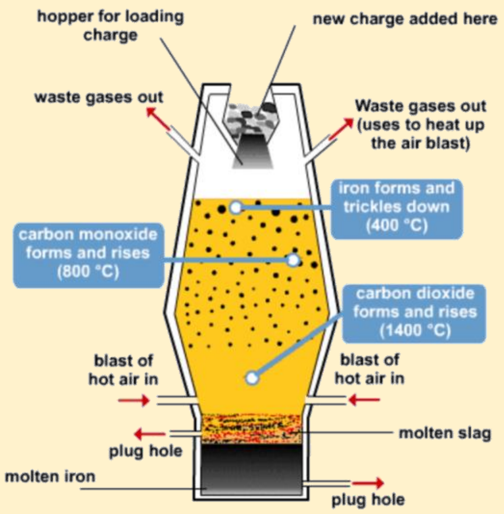
4. Limestone is added to the furnace which decomposes to form calcium oxide.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

5. Calcium oxide combines with silicon dioxide to form slag, which floats on top of the molten iron and can be removed physically.

$$CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$$

The blast furnace



6.3.1.1 Typical properties of transition metals

By the end of this section, you should be able to:

- Describe typical properties of transition elements/compounds. (8.2f)

Iron is a typical transition metal. The following properties are usually attributed to transition metals.

- They can exist in different oxidation states (E.g. Iron may exist as Fe^{2+} and Fe^{3+} in its compounds.)
- They can act as catalysts in both their elemental state and as compounds.
- They can form coloured compounds.
- They have relatively high densities.
- They have relatively high melting and boiling points.

6.3.1.2 Rusting and its prevention

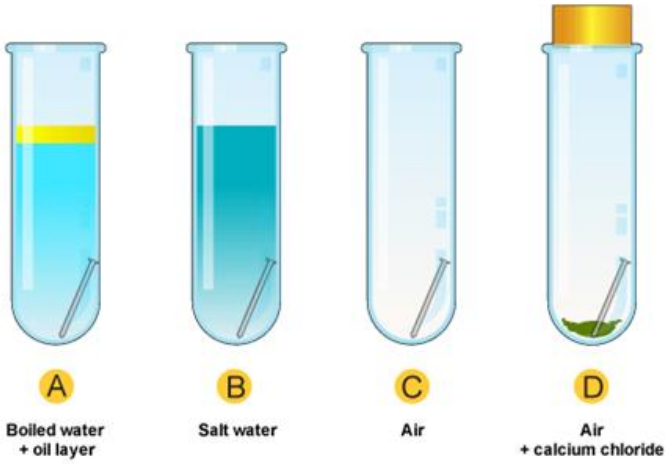
By the end of this section, you should be able to:

- Describe methods that prevent rusting. (8.1h)
- Investigate the conditions needed for iron to rust. (8.2h)
- Investigate the effectiveness of various rust prevention techniques in different situations. (8.3h)

During rusting, iron combines with oxygen in the air in the presence of water to form a hydrated form of iron(III) oxide which is called rust. Rust is an orange-red powder. Both water and oxygen are needed for iron to rust, and if one of these two substances is missing then rusting will not take place.

The rate of rusting of iron is accelerated in the presence of salt.

Most methods of rust prevention prevent the iron from coming into contact with air and/or water.



Source: BBC Bitesize

The following are different methods for rust prevention:

- **Painting** of cars, ships, and bridges. If the paint is scratched, the exposed area starts to rust. Rust will then spread under the paint which is still intact.
- **Using oil or grease**. This can be used in moving parts of machinery which have to be lubricated and it also provides a protective coating. The layer of oil and grease must be renewed frequently.
- **Coating with plastic**: Fence netting and garden furniture are examples of iron objects that are covered with a layer of plastic to prevent them from rusting. Similar to painting

if the layer of plastic is broken/cracked the iron will rust due to the exposure to air which contains both water vapour and oxygen.

- **Electroplating**: Tin plating is used for food cans. Water faucets and bicycle parts can be chromium plated which both protects the steel and provides an attractive finish. Where the metal used is less reactive than iron, once it is scratched, the iron rusts in preference. This can be seen when food cans are scratched and start to rust. The iron then rusts quickly.
- **Sacrificial protection**: Magnesium and zinc are often used as sacrificial metals. They are more reactive than iron and so oxygen reacts with them in preference to iron. Magnesium or zinc blocks are screwed to ship hulls to mitigate rusting.
- **Galvanising**: This process involves coating iron with zinc. This stops oxygen and water reaching the metal underneath, but the zinc also acts as a sacrificial metal. Zinc is more reactive than iron, so oxygen reacts with zinc in preference to the iron object. Galvanising protects the iron from rusting for many years even in coastal areas.

6.3.1.3 Economic and environmental considerations

By the end of this section, you should be able to:

- Discuss the environmental issues surrounding the mining of metals. (8.2o)
- Evaluate the economic and environmental impact of the extraction of iron. (8.3o part)

The environmental impact of iron ore mining in all its phases from excavation to processing to transportation may have detrimental effects on air quality, water quality, and living organisms. About 98% of iron ore on the global market is used in iron and steel production.

The high demand for iron, necessitates continuous mining and processing, which generates a large amount of solid and liquid waste.

The main sources of emissions during both the construction and operation phases include the products of combustion such as oxides of nitrogen, carbon dioxide, carbon monoxide, and sulfur dioxide from the operation of equipment.

Dust emissions are proportional to the disturbed land area and the level of activity which vary substantially from day to day with varying meteorological conditions.

Some animals are more susceptible to change than others. Large mammals such as black bears and wolves show behavioural changes and are sensitive to noise levels caused by iron ore mining. These disturbances increase the distances moved by the animals and may effectively result in starvation and a decrease in their population.

Water is one of the major natural resources that is being polluted by iron ore mining operations. Metals are leached out, and acidic water carries them downstream to the sea. Water bodies are polluted during iron ore mining operations.

Communities that live near quarries and mining sites suffer from noise, air, land, and even water pollution. In the long run, and most especially if the company is not following safety and security protocols, it will pose detrimental and hazardous risks to the health of the people living in the community.

Although there are added jobs and increased salaries for the people as well as increased revenue to the community, this comes at a high cost.

6.3.1.4 Uses of iron – advantages and disadvantages

By the end of this section, you should be able to:

- List uses of iron in everyday life. (8.1e part)
- State advantages and disadvantages of using iron. (8.2e part)

Iron is an important element for living things that have red blood. It is a main component of haemoglobin, a protein in red blood cells that carries oxygen and carbon dioxide.

Iron is one of the three ferromagnetic elements. The other two are cobalt and nickel.

Iron is less brittle than stone yet, compared to wood or copper, it is extremely strong. If properly heated it is relatively easy to shape into various forms because it is malleable. Such properties make iron versatile for various uses.

Cast iron is very brittle (it cracks easily) but it has a greater resistance to corrosion than either pure iron or steel. Cast iron is used for manhole covers on roads and pavements and as engine blocks for petrol and diesel engines.

Pure iron is called wrought iron. Wrought iron is malleable and is mainly used in ornamental work for gates.

Iron is also the catalyst in the Haber Process. The large majority of iron from the blast furnace is converted to steel.

A disadvantage of using iron is that it rusts when exposed to air. The rate of rusting is accelerated when it is used near the sea because of the salt present. Therefore to protect iron from rusting some rusting prevention must be employed which increases both the initial cost, as well as the cost of ownership, of iron made objects.

6.3.2 Aluminium

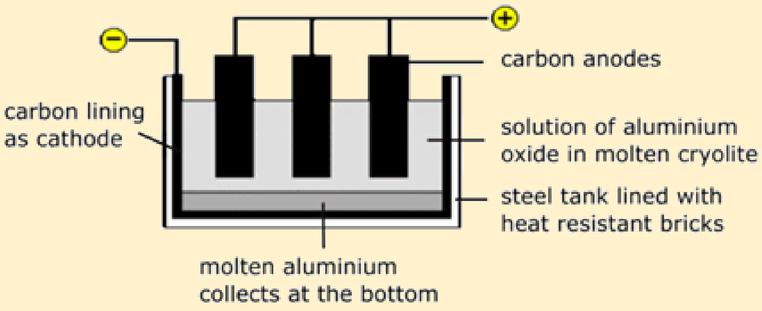
By the end of this section, you should be able to:

- Describe the essential chemical reactions and conditions in the industrial extraction of aluminium from bauxite. *(Drawing of diagrams and technical details are not required.)* (8.3d part)
- Interpret the extraction of metals as examples of redox reactions. *(In terms of loss or gain of electrons.)* (8.3g part)

Aluminium is the most abundant metallic element in the Earth's crust. It is extracted by electrolysis from **bauxite**, an ore which mostly consists of aluminium oxide (Al_2O_3). The process involves the following steps:

1. Bauxite is first treated with sodium hydroxide to obtain pure aluminium oxide (known as alumina) to remove impurities such as iron(III) oxide and sand.
2. The purified aluminium oxide is then dissolved in molten cryolite. Cryolite is used to reduce the melting point of the aluminium oxide from 2072 °C to around 900 °C.
3. The molten mixture is then electrolysed in a cell as shown below.

Diagram of the industrial electrolytic cell used to produce aluminium



Adapted from: chemguide.co.uk

The anodes are blocks of graphite which are lowered into the molten mixture. The cathode is the graphite lining of the cell.

During electrolysis the negatively charged oxide ions are attracted to the anode where they lose electrons (loss of electrons is oxidation) to form oxygen gas.

$$6O^{2-}(l) \rightarrow 3O_2(g) + 12e^-$$

The positive aluminium ions are attracted to the cathode. They gain electrons (gain of electrons is reduction) to form molten aluminium metal.

$$4Al^{3+}(l) + 12e^- \rightarrow 4Al(l)$$

The overall reaction which takes place is:

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

The molten aluminium collects at the bottom of the cell and is collected at regular intervals.

At the working temperature (900 °C) of the electrolytic cell, the oxygen liberated reacts with the graphite anodes and they are oxidized producing carbon dioxide.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

The anodes wear away and must be replaced on a regular basis.

6.3.2.1 Environmental and economic considerations

By the end of this section, you should be able to:

- Discuss the environmental issues surrounding the mining of metals. (8.2o)
- Evaluate the economic and environmental impact of the extraction of aluminium. (8.3o part)


Source: thesca.org

Aluminium is an abundant element in the Earth’s crust but getting the bauxite ore out of the ground is not an easy process, requiring heavy machinery to blow up acres of land. Fortunately, the environmental impact of that process can be balanced out by rehabilitation of the mined land, recycling, and reducing consumption.

The electrolytic process requires huge amounts of electricity, so it is an expensive process. Normally these industries are located near geothermal and hydroelectricity power stations so that electricity would be cheaper.

The graphite anodes which are regularly replaced are very expensive.

Some of the cryolite is lost in the electrolytic process and needs to be replaced.

Turning bauxite into aluminium requires a lot of electrical energy and has a variety of environmental impacts.

Mines affect plant and animal life immediately surrounding an area and beyond for multiple generations. Cutting trees and grasslands contributes to biodiversity loss, habitat loss, carbon emissions, and erosion. Mining operations have taken this damage into account more and more in recent years, incorporating plans to restore and regrow natural habitats once mining is completed.

Refining processes use electricity and water, which may cause higher carbon emissions, create air, water, noise, and heat pollution.


Source: thirtytonetzero.com

As explained above, mining takes its toll on the environment. However, aluminium can be recycled! There is still a lot of aluminium lost to landfills, which can be improved by consumer awareness.

Compared to producing new aluminium from raw bauxite, recycling consumes just 5% of the energy and releases 5% of the greenhouse gases. Infinitely recyclable, aluminium loses none of its integrity even when it is melted down repeatedly.

6.3.2.2 Uses of aluminium – advantages and disadvantages

By the end of this section, you should be able to:

- List uses of aluminium in everyday life. (8.1e part)
- State advantages and disadvantages of using aluminium. (8.2e part)

Aluminium is usually alloyed with other elements such as silicon, copper, or magnesium. Pure aluminium isn't very strong, and alloying it adds to its strength.

Aluminium is especially useful because it has the following properties:

- has a low density.
- is strong when turned into an alloy.
- is an excellent conductor of electricity.
- has a good appearance.
- resists corrosion because of the thin layer of aluminium oxide on its surface.

These properties allow for aluminium to have the following uses:

- As an alloy it is used in aircraft because high tensile strength is combined with lightness.
- It is used in cooking utensils because it is relatively light, is a good conductor of heat and resistant to chemical attack.
- In overhead electric cables because of its low density.
- In aluminium paint.

6.3.3 The Three Rs: Reduce, Reuse, Recycle

By the end of this section, you should be able to:

- Describe the best course of action when considering the finite nature of many metals. *(Reduce, reuse, recycle.)* (8.2p)
- Evaluate the best course of action when considering the finite nature of many metals. *(Reduce, reuse, recycle)* (8.3p)

Reducing the amount you buy by purchasing only the goods that we need, is the most important of all the options to manage waste.

If every household reduces its waste this would be a way to reduce the problem. You can begin with analysing – what you throw away and what goods you need at home.

Today many materials we use such as aluminium and iron which we discussed in this chapter, glass, building materials like limestone, steel and wood, clay ceramics and most plastics are made from limited natural resources.

Some items made from these materials can be **reused**, and this saves the most energy and reduces the impact on the environment. For example, glass bottles only need to be washed and sterilised before they can be filled again. Other products cannot be reused in this way, but they can be **recycled.**

6.3.3.1 Advantages of recycling

The advantages of recycling compared to producing materials and objects from natural resources include:

- Fewer quarries and mines are needed to extract finite reserves of metal ores.
- Less crude oil (petroleum) needs to be extracted to make plastics.
- Less energy is needed for recycling, so the emission of greenhouse gases is reduced.
- The amount of waste that ends in landfills is reduced.
- Recycling has a positive economic impact since it allows for more jobs to be created.
- Manufacturing facilities will develop alternative uses for recycled materials.

6.3.3.2 Disadvantages of recycling

The following are disadvantages of recycling:

- The collection and transport of used items needs organisation, workers, vehicles, and fuel.
- It can be difficult to sort different metals and other material from one another.
- The sorted metal may need to be transported to where it can be turned into ingots.

The amount of sorting required depends on the purity of the mixture of metals and on how pure you need the final recycled metal to be. For example, copper for electrical wiring requires a very high purity, so if wiring is made from recycled copper, the waste copper will need processing before being melted again into copper for wiring.

However, steel used in construction doesn't need such a high purity. This means that scrap iron can easily be added to the steel furnace when steel is being made. This reduces the cost of making steel from iron ore.

6.4 Redox reactions

By the end of this section, you should be able to:

- Interpret the extraction of metals as examples of redox reactions. *(In terms of loss or gain of oxygen/hydrogen, loss or gain of electrons and change in oxidation numbers. Oxidation numbers limited to binary compounds. Oxidising and reducing agents.)* (8.3g)

Oxidation and reduction take place simultaneously.

When one substance in the reaction is **oxidised**, the other is **reduced**. This reaction is called an oxidation-reduction reaction or **redox** reaction.

Not all chemical reactions are redox reactions. The terms oxidation and reduction have several definitions that are related to particular situations as shown in the following subsections.

6.4.1 Redox in terms of gain or loss of oxygen

When a substance **combines with oxygen**, it becomes **oxidised**. When a substance **loses oxygen**, it is **reduced**.

This can be summarised as follows:

Oxidation is the gain of oxygen.

Reduction is the loss of oxygen.

The following examples show different kinds of reaction where gain/loss of oxygen happens:

Example 1: Heating a metal in oxygen.

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

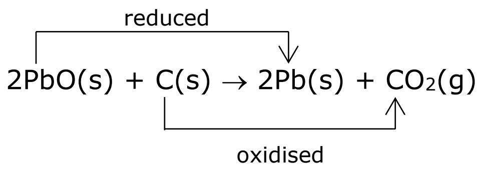
- Magnesium gains oxygen during combustion and so magnesium is oxidised.
- Oxygen is the **oxidising agent** while magnesium is the **reducing agent**.

An **oxidising agent** gives oxygen to the other substance, and it is reduced.

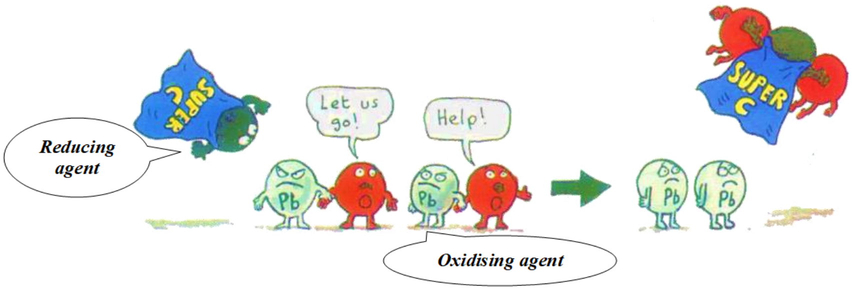
A **reducing agent** takes oxygen from the other substance, and it is oxidised.

All reactions involving the reaction of an element with oxygen are redox reactions.
Examples include all combustion reactions and rusting of iron.

Example 2: Heating a metal oxide with carbon.



$$2PbO(s) + C(s) \rightarrow 2Pb(s) + CO_2(g)$$



Adapted from "Chemistry For You" by Lawrie Ryan

- Lead(II) oxide is **reduced** to lead because oxygen is removed, so it is the **oxidising agent**.
- Carbon is **oxidised** to carbon dioxide because it gains oxygen, so it is the **reducing agent**.

Similar redox reactions can be used to extract metals from their ores.

6.4.2 Redox in terms of gain or loss of hydrogen

When a substance **combines with hydrogen**, it is **reduced**. When a substance **loses hydrogen**, it is **oxidised**.

This can be summarised as follows:

Oxidation is the **loss of hydrogen**.

Reduction is the **gain of hydrogen**.

The following examples show different kinds of reaction where gain/loss of hydrogen happens:

Example 1: Addition of hydrogen to a substance.

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

Hydrogen has been added to ethene (C_2H_4) to form ethane (C_2H_6).

- Ethene is **reduced** to ethane due to **addition of hydrogen**.
- Ethene is the **oxidising agent** while hydrogen is the **reducing agent**.

An **oxidising agent** takes hydrogen from another substance, and it is **reduced.**

A **reducing agent** gives hydrogen to another substance, and it is **oxidised**.

Example 2: Removal of hydrogen from a substance.

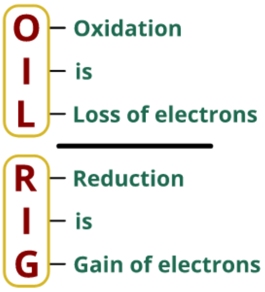
oxidised

$$H_2S(g) + Cl_2(g) \rightarrow S(s) + 2HCl(g)$$

reduced

- Hydrogen sulfide is **oxidised** to sulfur due to the **removal of hydrogen** and is the **reducing agent**.
- Chlorine is **reduced** to hydrogen chloride as it **gains hydrogen,** and it is the **oxidising agent**.

6.4.3 Redox in terms of gain or loss of electrons

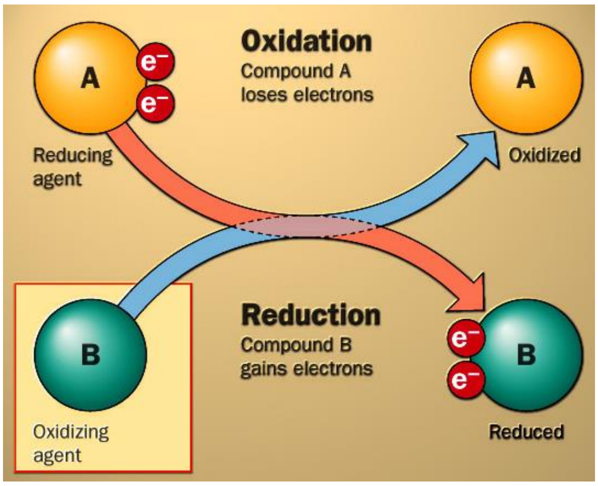


Redox reactions can be interpreted in terms of the electrons substances gain or lose. This way of looking at redox reactions goes further from reactions involving oxygen and hydrogen. The following statements are important to remember:

- **Oxidation** is the process of **losing electrons**.
- **Reduction** is the process of **gaining electrons**.

Consequently:

- An **oxidising agent** is an **acceptor of electrons**.
- A **reducing agent** is a **donor of electrons**.

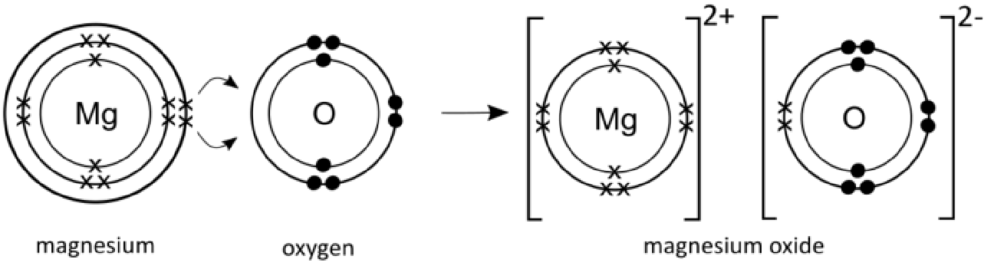


Source: www.the-mad-scientist.net

The following examples illustrate how redox reactions can be interpreted.

Example 1: Reaction of magnesium with oxygen.

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



Magnesium loses electrons to form Mg^{2+} ions:

$$2Mg - 4e^- \rightarrow 2Mg^{2+} \quad \textbf{or} \quad 2Mg \rightarrow 2Mg^{2+} + 4e^-$$

Oxygen gains electrons to form O^{2-} ions:

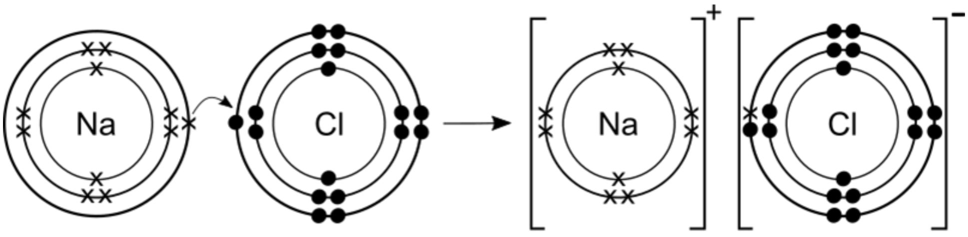
$$O_2 + 4e^- \rightarrow 2O^{2-}$$

Therefore:

- Magnesium is **oxidised** to Mg^{2+} since magnesium **lost electrons.**
- Oxygen is **reduced** to O^{2-} since it **gained electrons.**
- Magnesium is the **reducing agent** because it **gives electrons.**
- Oxygen is the **oxidising agent** since it **accepts electrons.**

Example 2: Reaction of sodium with chlorine.

$$2Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$$



sodium chlorine sodium chloride

Sodium loses electrons to form Na^+ ions:

$$2Na - 2e^- \rightarrow 2Na^+ \quad \textbf{or} \quad 2Na \rightarrow 2Na^+ + 2e^-$$

Chlorine gains electrons to form Cl^- ions:

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

Therefore:

- Sodium is **oxidised** to Na^+ since sodium **lost electrons.**
- Chlorine is **reduced** to Cl^- since it **gained electrons.**
- Sodium is the **reducing agent** because it **gives electrons.**
- Chlorine is the **oxidising agent** since it **accepts electrons.**

6.4.4 Redox in terms of change in oxidation numbers

Redox reactions can also be explained in terms of oxidation numbers. An atom uses electrons to bond with atoms from other elements. In a chemical formula each atom has an oxidation number.

The oxidation number is the number of electrons that an atom uses to bond with atoms of another element. The oxidation number indicates the state of oxidation of an element in a compound.

Oxidation numbers are worked out by a set of rules.

1. All **elements in the free state** (i.e., uncombined with any other elements) have an oxidation number of **zero**.
   E.g.: Fe(s), Na(s), S(s), C(s) and $Cl_2(g)$ all have an oxidation state of 0.

2. In the case of **simple ions**, the element has an oxidation number with the **same size and sign of the charge** on the ion.
   - **Group 1 metal ions** have an oxidation number of **+1**.
     E.g. Na^+, K^+ have an oxidation state of +1.

   - **Group 2 metal ions** have an oxidation number of **+2**.
     E.g. Ca^{2+}, Mg^{2+} have an oxidation state of +2.

   - **Group 3 metal ions** have an oxidation number of **+3**.
     E.g. Al^{3+} has an oxidation state of +3.

   - **Group 7 non-metal ions** have an oxidation number of **-1**.
     E.g. Cl^- and Br^- have an oxidation state of -1.

   - Hydrogen has an oxidation number of **+1** when combined with non-metals (e.g. hydrogen chloride) and **−1** when combined with metals (e.g. in sodium hydride).

3. The **oxide ion** always has an oxidation number of **-2**.

4. The oxidation number of polyatomic ions is the **same size** and **sign of the charge** as that of the polyatomic ion.
   E.g. The sulfate ion has an oxidation number of **-2**.

5. The **sum of all the oxidation numbers** of the elements in a compound is **zero**.

To identify reduction and oxidation in a chemical reaction the rules are used to assign an oxidation number to each element in every reactant and product.

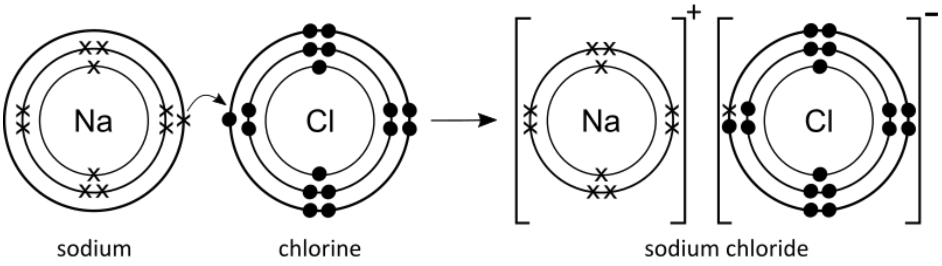
The change in oxidation number for each element during the reaction is noted such that:

Oxidation occurs when the **oxidation number** of the element **increases.**

Reduction occurs when the **oxidation number** of the element **decreases.**

Let's consider a synthesis reaction where sodium reacts with chlorine.

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$



In terms of electron transfer:

$2Na \rightarrow 2Na^+ + 2e^-$ Sodium atoms are **oxidised** to sodium ions due to **loss of electrons**.

$Cl_2 + 2e^- \rightarrow 2Cl^-$ Chlorine atoms are **reduced** to chloride ions due to **gain of electrons**.

In terms of oxidation numbers:

Write the oxidation number underneath each substance.

$$\overset{0}{2Na(s)} + \underset{0}{Cl_2(g)} \rightarrow \overset{+1}{2Na}\,\underset{-1}{Cl(s)}$$

Determine whether the oxidation number has increased or decreased.

Sodium is oxidised to Na^+ due to an increase in oxidation number [0 to +1]

Chlorine is reduced to Cl^- due to a decrease in oxidation number [0 to -1]

Oxidation:
Sodium **loses electrons** to become Na^+: the **oxidation number increases** (more positive)

Reduction:
Chlorine **gains electrons** to become Cl^-: the **oxidation number decreases** (more negative)

6.4.5 Summary on redox reaction

Oxidation is	Reduction is
the gain of oxygen.	the loss of oxygen.
the loss of hydrogen.	the gain of hydrogen.
the loss of electrons.	the gain of electrons.
an increase in oxidation number.	a decrease in oxidation number.

6.4.6 The blast furnace as an example of redox reactions

A series of chemical reactions occur in the blast furnace:

1. The coke (carbon) burns with oxygen to produce carbon dioxide. This reaction is exothermic (gives out heat). In this **oxidation** reaction carbon is **oxidised** to carbon dioxide because it is **gaining oxygen**.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

2. The CO_2 then reacts with more coke to produce carbon monoxide gas. In this **reduction** reaction carbon dioxide is **reduced** to carbon monoxide since it is **losing oxygen**.

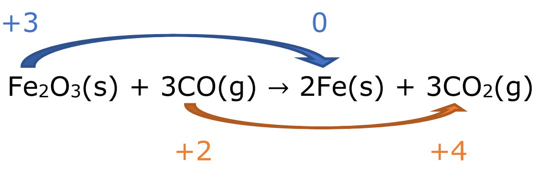
$$CO_2(g) + C(s) \rightarrow 2CO(g)$$

3. Carbon monoxide acts as a **reducing agent** and reacts with the iron ore (Fe_2O_3) to reduce it to iron. Since Fe_2O_3 is reduced as it loses oxygen, it is acting as the **oxidising agent**.

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

Since all the above reactions include reduction and oxidation processes occurring simultaneously, they are known as **redox** reactions.

The above equation can be interpreted in terms of oxidation numbers.



The oxidation state of iron decreased from +3 to 0 so it was reduced, while the oxidation state of carbon increased from +2 to +4 so it was oxidised.

6.5 Net ionic equations

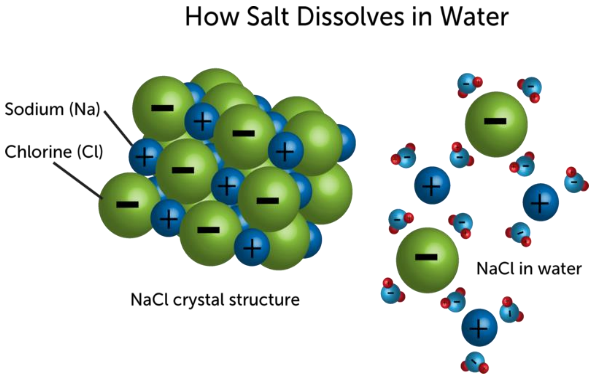
By the end of this section, you should be able to:

- Represent reactions of non-oxidising acids with bases/alkalis, carbonates/hydrogencarbonates, fairly reactive metals, and sulfites, using net ionic equations. (4.3c)
- Represent the reaction of an alkali with an ammonium salt using net ionic equations. (4.3d)
- Represent the precipitation of an insoluble salt using net ionic equations. (4.3e)
- Represent displacement reactions using net ionic equations. (8.3k part)

In a solid ionic compound, the positive and negative ions are strongly bonded to each other in a crystal structure known as a lattice. When dissolved in water, the ions separate and become surrounded with water molecules.

An ionic equation shows the reaction that takes places as ions react.

Some of the ions do not react and remain unchanged. These are called **spectator ions** as they are present but do not take part in the reaction.



The following steps should be followed when writing a net ionic equation:

1. Write a balanced chemical equation including state symbols.
2. Rewrite the equation splitting soluble **ionic compounds** into ions.
   (Covalent compounds, gases, and metals do **not** have ions. Precipitates are not separated into ions.)
3. Cross out the spectator ions. These are ions which remain the same during the reaction as they do not take part in the reaction.
4. Copy the ionic equation including state symbols.
   (Check that the number of elements / ions on the left-hand side is equal to the number of elements / ions on the right-hand side and the net charge on each side of the equation is equal.)

The following sections delve deeper into the different kinds of reactions that can be represented using a net ionic equation.

6.5.1 Net ionic equations of non-oxidising acids with bases / alkalis

Not all acids are the same.

- Some acids, such as hydrochloric acid, never act as an oxidising agent not even when they are concentrated.
- Other acids, such as sulfuric acid, are not an oxidising agent when they are dilute but act as oxidising agents when they are concentrated.
- Other acids, such as nitric acid, act as an oxidising agent both when they are dilute and concentrated.

The following steps show how you can write a net ionic equation for the reaction between dilute hydrochloric acid and copper(II) oxide (a base).

Step 1: Write the balanced chemical equation in full.

base	**+**	**acid**	**→**	**salt**	**+**	**water**
$CuO(s)$	+	$2HCl(aq)$	→	$CuCl_2(aq)$	+	$H_2O(l)$

Step 2: Split soluble ionic compounds into their ions.

$CuO(s)$	+	$H^+ \quad Cl^-$ $H^+ \quad Cl^-$	→	$Cu^{2+} \quad Cl^-$ Cl^-	+	H_2O

Step 3: Identify and cross out the spectator ions.

$CuO(s)$	+	$H^+ \; H^+ \; \cancel{Cl^- \; Cl^-}$	→	$Cu^{2+} \; \cancel{Cl^- \; Cl^-}$	+	H_2O

Step 4: Write the net ionic equation in its simplest ratio.

$CuO(s)$	**+**	**$2H^+(aq)$**	→	**$Cu^{2+}(aq)$**	+	**$H_2O(l)$**

The following steps show how you can write a net ionic equation for the reaction between dilute sulfuric acid and potassium hydroxide solution (an alkali).

Step 1: Write the balanced chemical equation in full.

alkali	**+**	**acid**	**→**	**salt**	**+**	**water**
$2KOH\ (aq)$	+	$H_2SO_4\ (aq)$	→	$K_2SO_4\ (aq)$	+	$2H_2O\ (l)$

Step 2: Split soluble ionic compounds into their ions.

$K^+ \quad OH^-$ $K^+ \quad OH^-$	+	$H^+ \quad SO_4^{2-}$ H^+	→	$K^+ \quad SO_4^{2-}$ K^+	+	H_2O H_2O

Step 3: Identify and cross out the spectator ions.

$\cancel{K^+ \; K^+} \; OH^- \; OH^-$	+	$H^+ \; H^+ \; \cancel{SO_4^{2-}}$	→	$\cancel{K^+ \; K^+} \; \cancel{SO_4^{2-}}$	+	$H_2O \quad H_2O$

Step 4: Write the net ionic equation in its simplest ratio.

$OH^-(aq)$	**+**	**$H^+(aq)$**	→	**$H_2O(l)$**		

6.5.2 Net ionic equations of non-oxidising acids with carbonates / hydrogencarbonates

The following steps show how you can write a net ionic equation for the reaction between dilute sulfuric acid and potassium carbonate.

Step 1: Write the balanced chemical equation in full.

carbonate + acid → salt + water + carbon dioxide

$K_2CO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)$

Step 2: Split soluble ionic compounds into their ions.

$K^+ \; K^+ \; CO_3^{2-} + H^+ \; H^+ \; SO_4^{2-} \rightarrow K^+ \; K^+ \; SO_4^{2-} + H_2O + CO_2$

Step 3: Identify and cross out the spectator ions.

$\cancel{K^+ \; K^+} \; CO_3^{2-} + H^+ \; H^+ \; \cancel{SO_4^{2-}} \rightarrow \cancel{K^+ \; K^+} \; \cancel{SO_4^{2-}} + H_2O + CO_2$

Step 4: Write the net ionic equation in its simplest ratio.

$\mathbf{CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)}$

The following steps show how you can write a net ionic equation for the reaction between dilute hydrochloric acid and potassium hydrogencarbonate.

Step 1: Write the balanced chemical equation in full.

hydrogencarbonate + acid → salt + water + carbon dioxide

$KHCO_3(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l) + CO_2(g)$

Step 2: Split soluble ionic compounds into their ions.

$K^+ \; HCO_3^- + H^+ \; Cl^- \rightarrow K^+ \; Cl^- + H_2O + CO_2$

Step 3: Identify and cross out the spectator ions.

$\cancel{K^+} \; HCO_3^- + H^+ \; \cancel{Cl^-} \rightarrow \cancel{K^+ \; Cl^-} + H_2O + CO_2$

Step 4: Write the net ionic equation in its simplest ratio.

$\mathbf{HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)}$

6.5.3 Net ionic equations of non-oxidising acids with fairly reactive metals

The following steps show how you can write a net ionic equation for the reaction between dilute sulfuric acid and magnesium.

Step 1: Write the balanced chemical equation in full.

metal	+	**acid**	→	**salt**	+	**hydrogen**
$Mg(s)$	+	$H_2SO_4(aq)$	→	$MgSO_4(aq)$	+	$H_2(g)$

Step 2: Split soluble ionic compounds into their ions.

Mg	+	H^+ H^+ SO_4^{2-}	→	Mg^{2+} SO_4^{2-}	+	H_2

Step 3: Identify and cross out the spectator ions.

Mg	+	$H^+\ H^+\ \cancel{SO_4^{2-}}$	→	$Mg^{2+}\ \cancel{SO_4^{2-}}$	+	H_2

Step 4: Write the net ionic equation in its simplest ratio.

$Mg\ (s)$	**+**	**$2H^+\ (aq)$**	→	**$Mg^{2+}\ (aq)$**		**$H_2\ (g)$**

6.5.4 Net ionic equations of non-oxidising acids with sulfites

The following steps show how you can write a net ionic equation for the reaction between dilute hydrochloric acid and sodium sulfite.

Step 1: Write the balanced chemical equation in full.

sulfite	+	**acid**	→	**salt**	+	**water**	+	**sulfur dioxide**
$Na_2SO_3(aq)$	+	$2HCl(aq)$	→	$2NaCl(aq)$	+	$H_2O(l)$	+	$SO_2(g)$

Step 2: Split soluble ionic compounds into their ions.

Na^+ Na^+ SO_3^{2-}	+	H^+ Cl^- H^+ Cl^-	→	Na^+ Cl^- Na^+ Cl^-	+	H_2O	+	SO_2

Step 3: Identify and cross out the spectator ions.

$\cancel{Na^+\ Na^+}\ SO_3^{2-}$	+	$H^+\ H^+\ \cancel{Cl^-\ Cl^-}$	→	$\cancel{Na^+\ Na^+}\ \cancel{Cl^-\ Cl^-}$	+	H_2O	+	SO_2

Step 4: Write the net ionic equation in its simplest ratio.

$SO_3^{2-}(aq)$	**+**	**$2H^+(aq)$**	→			**$H_2O(l)$**	**+**	**$SO_2(g)$**

6.5.5 Net ionic equations of an alkali with an ammonium salt

The following steps show how you can write a net ionic equation for the reaction between sodium hydroxide (an alkali) and ammonium nitrate (an ammonium salt).

Step 1: Write the balanced chemical equation in full.

ammonium salt	+	**alkali**	→	**salt**	+	**water**	+	**ammonia**
$NH_4NO_3(aq)$	+	$NaOH(aq)$	→	$NaNO_3(aq)$	+	$H_2O(l)$	+	$NH_3(g)$

Step 2: Split soluble ionic compounds into their ions.

$$NH_4^+ \quad NO_3^- \quad + \quad Na^+ \quad OH^- \quad \rightarrow \quad Na^+ \quad NO_3^- \quad + \quad H_2O \quad + \quad NH_3$$

Step 3: Identify and cross out the spectator ions.

$$NH_4^+ \quad \cancel{NO_3^-} \quad + \quad \cancel{Na^+} \quad OH^- \quad \rightarrow \quad \cancel{Na^+} \quad \cancel{NO_3^-} \quad + \quad H_2O \quad + \quad NH_3$$

Step 4: Write the net ionic equation in its simplest ratio.

$$\mathbf{NH_4^+(aq) \quad + \quad OH^-(aq) \quad \rightarrow \quad H_2O(l) \quad + \quad NH_3(g)}$$

6.5.6 Net ionic equations of precipitation reactions

The following steps show how you can write a net ionic equation for the reaction between iron(II) chloride and sodium hydroxide.

Step 1: Write the balanced chemical equation in full.

soluble salt	+	**soluble salt**	→	**insoluble salt**	+	**soluble salt**
$FeCl_2(aq)$	+	$2NaOH(aq)$	→	$Fe(OH)_2(s)$	+	$2NaCl(aq)$

Step 2: Split soluble ionic compounds into their ions except for precipitates.

$$Fe^{2+} \begin{matrix} Cl^- \\ Cl^- \end{matrix} \quad + \quad \begin{matrix} Na^+ & OH^- \\ Na^+ & OH^- \end{matrix} \quad \rightarrow \quad Fe(OH)_2(s) \quad + \quad \begin{matrix} Na^+ & Cl^- \\ Na^+ & Cl^- \end{matrix}$$

Step 3: Identify and cross out the ions that do **not** form the precipitate.

$$Fe^{2+} \quad \cancel{Cl^- \; Cl^-} \quad + \quad \cancel{Na^+ \; Na^+} \; OH^- \; OH^- \quad \rightarrow \quad Fe(OH)_2(s) \quad + \quad \cancel{Na^+ \; Na^+} \; \cancel{Cl^- \; Cl^-}$$

Step 4: Write the net ionic equation in its simplest ratio.

$$\mathbf{Fe^{2+}(aq) \quad + \quad 2OH^-(aq) \quad \rightarrow \quad Fe(OH)_2(s)}$$

Note that precipitation reactions are **NOT** redox reactions as shown by the unchanged oxidation numbers below:

$$\begin{array}{llll} \mathbf{Fe^{2+}(aq)} \; + & \mathbf{2OH^-(aq)} & \rightarrow & \mathbf{Fe(OH)_2(s)} \\ +2 & \quad -1 & & +2 \quad -1 \end{array}$$

6.5.7 Net ionic equations of displacement reactions

A more reactive metal displaces a less reactive metal from a salt solution.

The following steps show how you can write a net ionic equation for the reaction between magnesium and copper(II) sulfate solution.

$$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$$

$$Mg + Cu^{2+}\ \cancel{SO_4^{2-}} \rightarrow Mg^{2+}\ \cancel{SO_4^{2-}} + Cu$$

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

A more reactive halogen displaces a less reactive halogen from solution of its halide.

The following steps show how you can write a net ionic equation for the reaction between chlorine gas and potassium bromide solution.

$$Cl_2(g) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$$

$$Cl_2 + \cancel{2K^+}\ 2Br^- \rightarrow \cancel{2K^+}\ 2Cl^- + Br_2$$

$$\mathbf{Cl_2(g) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)}$$

6.6 Displacement reactions as redox reactions

By the end of this section, you should be able to:

- Interpret displacement reactions of metal/metal ion mixtures in terms of oxidation and reduction. (8.3l)
- Interpret displacement reactions of halogen/halide mixtures in terms of oxidation and reduction. (6.3c)

Displacement reactions involve redox reactions whereby electrons are gained and lost by reacting particles.

You could remember it as:

OIL RIG – Oxidation **I**s **L**oss of electrons, **R**eduction **I**s **G**ain of electrons.

6.6.1 Metal displacement reactions

In a displacement reaction, iron is more reactive than copper, so iron displaces copper ions from solution. The blue solution of copper(II) sulfate turns green due to the formation of iron(II) sulfate. A red brown deposit of copper covers the nail as shown in the adjacent picture.

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

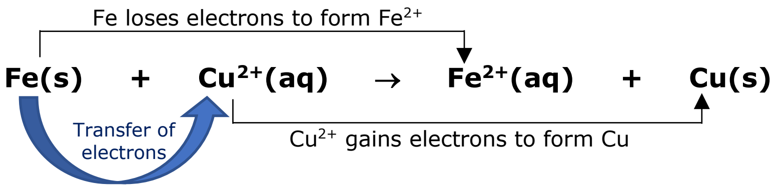
First work out the **ionic equation**:

$$Fe + Cu^{2+}\ \cancel{SO_4^{2-}} \rightarrow Fe^{2+}\ \cancel{SO_4^{2-}} + Cu$$

$$\mathbf{Fe(s) + Cu^{2+}}(aq) \rightarrow \mathbf{Fe^{2+}}(aq) + \mathbf{Cu}(s)$$

Fe loses electrons and becomes Fe^{2+} ions. Therefore, iron supplies electrons to Cu^{2+} ions.

Cu^{2+} ions accept electrons to become copper atoms.



Then work the **half equations:**

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ Fe is oxidised to Fe^{2+} due to loss of electrons.

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ Cu^{2+} is reduced to Cu due to gain of electrons.

Cu^{2+} ions are the **oxidising agent** as they accepted electrons.

Fe is the **reducing agent** since it donated electrons.

Note that metals are reducing agents as they react by losing electrons to form positive ions.

6.6.2 Reactions of metals with acids

When zinc is added to sulfuric acid, zinc sulfate and hydrogen are produced.

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

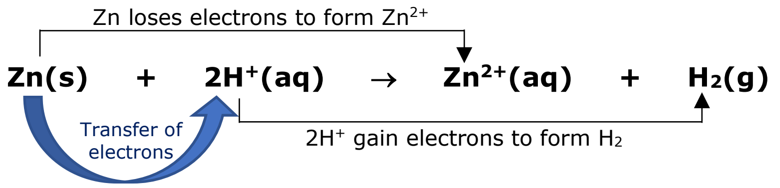
First work out the **ionic equation**:

$$Zn + H^+\ H^+\ \cancel{SO_4^{2-}} \rightarrow Zn^{2+}\ \cancel{SO_4^{2-}} + H_2$$

$$\mathbf{Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)}$$

Zn loses electrons and becomes Zn^{2+} ions. Therefore, zinc supplies electrons to the H^+ ions.

The H^+ ions accept electrons to become hydrogen molecules.



Then work the **half equations**:

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ Zn is oxidised to Zn^{2+} due to loss of electrons.

$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $2H^+$ is reduced to H_2 due to gain of electrons.

H^+ ions are the **oxidising agent** as they accepted electrons.

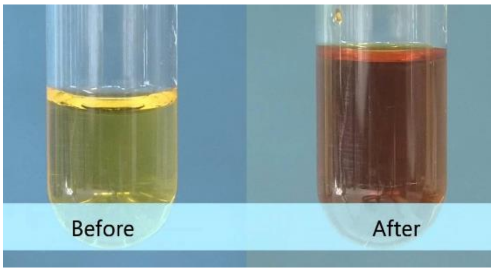
Zn is the **reducing agent** since it donated electrons.

6.6.3 Non-metal displacement reactions

Halogen displacement reactions are redox reactions. You may recall that redox reactions involve both oxidation (loss of electrons) and reduction (gain of electrons).

The reactivity of halogens decreases down the group. So, chlorine is more reactive than bromine which is more reactive than iodine.

In a displacement reaction bromine is more reactive than iodine and so bromine displaces iodide ions from solution. Bromine solution is an orange solution and after it reacts with sodium iodide it forms a brown solution.



bromine + sodium iodide → sodium bromide + iodine

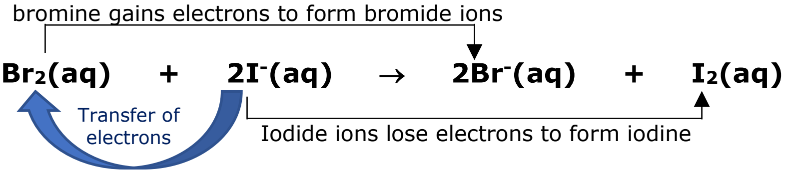
$$\mathbf{Br_2(aq) + 2NaI(aq) \rightarrow 2NaBr(aq) + I_2(aq)}$$

First work out the **ionic equation**:

$$Br_2 + \cancel{2Na^+}\ 2I^- \rightarrow \cancel{2Na^+}\ 2Br^- + I_2$$

$$\mathbf{Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)}$$

Bromine removes electrons from iodide ions (I^-). By gaining electrons bromine becomes bromide ions. As the iodide ions lose electrons they turn to iodine.



Then work the **half equations:**

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$

Bromine is reduced to bromide ions (Br^-) due to gain of electrons.

$2I^-(aq) \rightarrow I_2(aq) + 2e^-$

Iodide ions (I^-) are oxidised to iodine due to loss of electrons.

Bromine is the **oxidising agent** as it accepted electrons.

Iodide ions are the **reducing agent** since they donated electrons.

6.7 Allotropes of carbon

By the end of this section, you should be able to:

- Identify diamond, graphite, graphene, and carbon nanotubes from given molecular diagrams. (8.1n)
- Explain that diamond, graphite, and carbon nanotubes are allotropes. (8.2n)
- Relate the structure of diamond, graphite, and graphene (as examples of giant molecular structures), and carbon nanotubes to their properties and uses. (8.3n)

Carbon may exist as different **allotropes**.

Allotropes are different forms of the same element in the same state of matter.

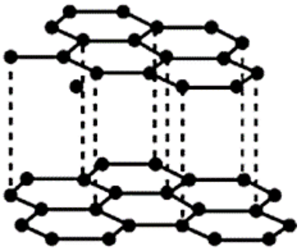
An element that has allotropes is said to show **allotropy**.

Diamond, graphite, and graphene are allotropes of carbon with different giant covalent structures. Giant covalent elemental substances have many atoms joined together by covalent bonds.

6.7.1 Graphite

Graphite is a giant covalent substance in which:

- each carbon atom is joined to three other carbon atoms by strong covalent bonds.
- the carbon atoms form a hexagonal layered network structure.
- the layers have weak forces between them and so can slide over each other.
- each carbon atom has one outer electron not used in bonding.
- these unbound electrons are delocalised and are free to move.


graphite

Delocalised electrons are free to move through the structure of graphite, so graphite can conduct electricity. This makes it useful for electrodes in batteries and electrolysis, and in electric motor brushes.

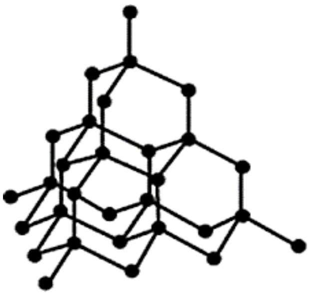
The layers in graphite can slide over each other because the forces between them are weak. This makes graphite slippery, so it is useful as a lubricant in electric motor brushes.

Due to the giant covalent structure graphite has a high melting point.

6.7.2 Diamond

Diamond is a giant covalent substance in which:

- each carbon atom is joined to four other carbon atoms by covalent bonds.
- the carbon atoms form a regular tetrahedral network structure.
- there are no free electrons.


diamond

The rigid network of carbon atoms, held together by strong covalent bonds, makes diamond very hard and contributes to a very high melting point (even higher than that of graphite). This makes it useful for cutting tools, such as diamond-tipped glass cutters and oil rig drills.

6.7.3 Comparison between diamond and graphite

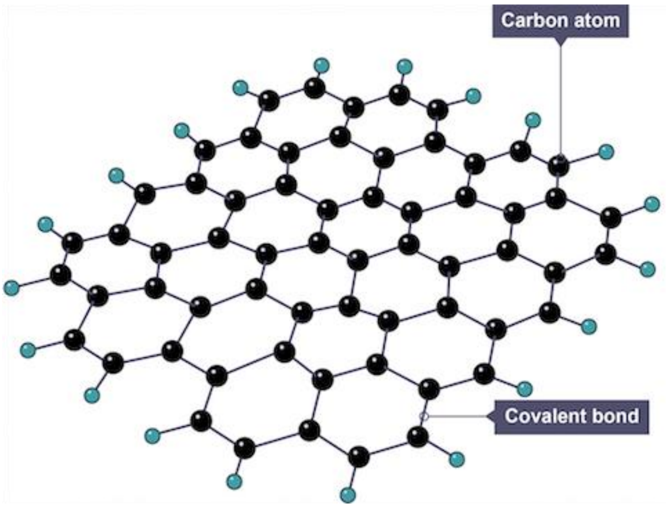
The following table compares the properties of diamond and graphite.

Diamond	Graphite
Colourless, transparent	Greyish black, opaque
Sparkling	Shiny
Density 3.5 g cm^{-3}	2.3 g cm^{-3}
Hardest natural substance	Soft and slippery
Burns at 900 °C	Burns at 700 °C
Does not conduct electricity	Good conductor

6.7.4 Graphene

Graphene is a single-atom thick layer of graphite with strong covalent bonds between each carbon atom. The atoms are arranged in hexagons. Its properties include:

- high melting and boiling points. Graphene's many covalent bonds are strong and substantial energy is needed to break them.
- good electrical conductivity. Each carbon atom has an unbonded electron. The unbonded electrons are delocalised electrons that are free to move and carry charge.
- very strong. Graphene's strong covalent bonds makes it 100 times stronger than steel. It is also the thinnest material possible – one atom thick – and very lightweight and transparent.



Source of text and picture: https://www.bbc.co.uk/bitesize

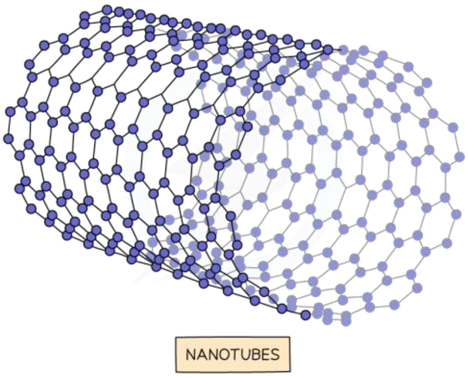
Infographic about graphene.

6.7.5 Carbon nanotubes

Graphene can also be rolled into a cylinder to produce an interesting type of fullerene called a nanotube. Fullerenes are a group of carbon allotropes which consist of molecules that form hollow tubes or spheres.

Carbon nanotubes have high tensile strength and are resistant to breaking or stretching.

As in graphene, nanotubes can also conduct electricity which makes them useful in composites and specialised materials, electronics, and nanotechnology.



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Nobel prize in chemistry

The Nobel Prize in Chemistry 1996 was awarded jointly to Robert F. Curl Jr., Sir Harold W. Kroto and Richard E. Smalley "for their discovery of fullerenes".

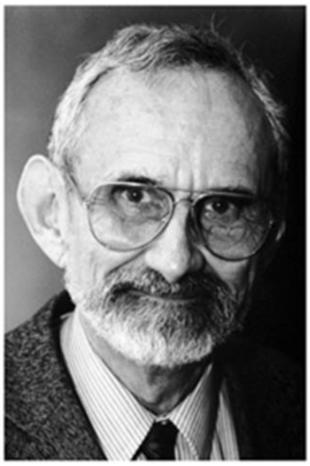
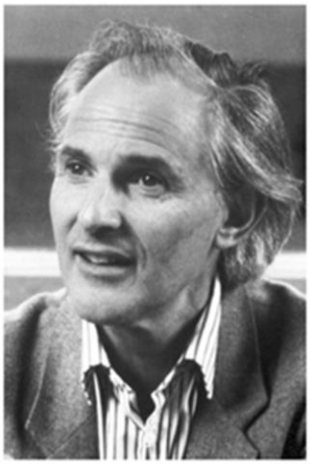


Photo from the Nobel Foundation archive.

Robert F. Curl Jr.

Prize share: 1/3



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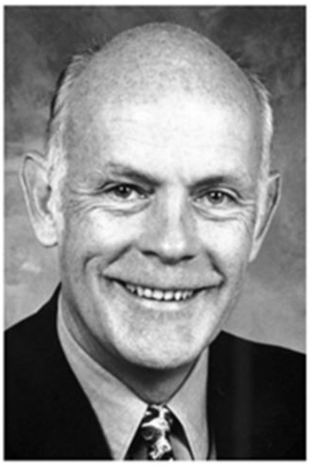
Sir Harold W. Kroto

Prize share: 1/3



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Richard E. Smalley

Prize share: 1/3

7 Periodic Table of the Elements

PERIODIC TABLE OF THE ELEMENTS

1	**2**											**3**	**4**	**5**	**6**	**7**	**0**
							1 **H** Hydrogen 1										4 **He** Helium 2
7 **Li** Lithium 3	9 **Be** Beryllium 4											11 **B** Boron 5	12 **C** Carbon 6	14 **N** Nitrogen 7	16 **O** Oxygen 8	19 **F** Fluorine 9	20 **Ne** Neon 10
23 **Na** Sodium 11	24 **Mg** Magnesium 12											27 **Al** Aluminium 13	28 **Si** Silicon 14	31 **P** Phosphorus 15	32 **S** Sulfur 16	35.5 **Cl** Chlorine 17	40 **Ar** Argon 18
39 **K** Potassium 19	40 **Ca** Calcium 20	45 **Sc** Scandium 21	48 **Ti** Titanium 22	51 **V** Vanadium 23	52 **Cr** Chromium 24	55 **Mn** Manganese 25	56 **Fe** Iron 26	59 **Co** Cobalt 27	59 **Ni** Nickel 28	63.5 **Cu** Copper 29	65 **Zn** Zinc 30	70 **Ga** Gallium 31	73 **Ge** Germanium 32	75 **As** Arsenic 33	79 **Se** Selenium 34	80 **Br** Bromine 35	84 **Kr** Krypton 36
85 **Rb** Rubidium 37	88 **Sr** Strontium 38	89 **Y** Yttrium 39	91 **Zr** Zirconium 40	93 **Nb** Niobium 41	96 **Mo** Molybdenum 42	99 **Tc** Technetium 43	101 **Ru** Ruthenium 44	103 **Rh** Rhodium 45	106 **Pd** Palladium 46	108 **Ag** Silver 47	112 **Cd** Cadmium 48	115 **In** Indium 49	119 **Sn** Tin 50	122 **Sb** Antimony 51	128 **Te** Tellurium 52	127 **I** Iodine 53	131 **Xe** Xenon 54
133 **Cs** Caesium 55	137 **Ba** Barium 56	139 **La** Lanthanum 57	178 **Hf** Hafnium 72	181 **Ta** Tantalum 73	184 **W** Tungsten 74	186 **Re** Rhenium 75	190 **Os** Osmium 76	192 **Ir** Iridium 77	195 **Pt** Platinum 78	197 **Au** Gold 79	201 **Hg** Mercury 80	204 **Tl** Thallium 81	207 **Pb** Lead 82	209 **Bi** Bismuth 83	210 **Po** Polonium 84	210 **At** Astatine 85	222 **Rn** Radon 86

Key:

a **X** Y b	*relative atomic mass* ***SYMBOL*** ***Name*** *atomic number*