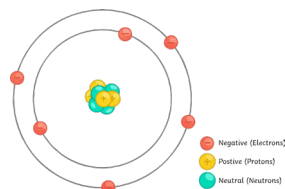


Atomic Structure and the Periodic Table – Foundation and Higher

Atoms

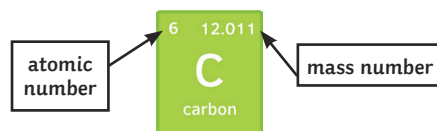
Contained in the nucleus are the **protons** and **neutrons**. Moving around the nucleus are the **electron** shells. They are negatively charged.

Particle	Relative Mass	Charge
proton	1	+1
neutron	1	0
electron	Very small	-1



Overall, atoms have no charge; they have the same number of protons as electrons. An ion is a charged particle - it does not have an equal number of protons to electrons.

Atomic Number and Mass Number



Equations and Maths

To calculate the **relative atomic mass**, use the following equation:

relative atomic mass (A_r) =

$$\frac{\text{sum of (isotope abundance} \times \text{isotope mass number)}}{\text{sum of abundances of all isotopes}}$$

Balancing Symbol Equations

There must be the same number of atoms on both sides of the equation:



$$\text{C} = 1$$

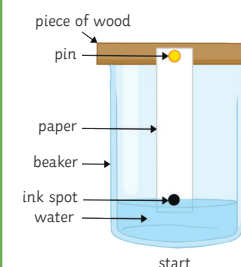
$$\text{O} = 4$$

$$\text{H} = 4$$

Mixtures, Chromatography and Separation

Mixtures – in a mixture there are no chemical bonds, so the elements are easy to separate. Examples of mixtures are air and salt water.

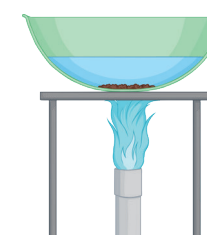
Chromatography – to separate out mixtures.



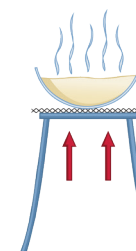
Filtration – to separate solids from liquids.



Evaporation – to separate a soluble salt from a solution; a quick way of separating out the salt.



Crystallisation – to separate a soluble salt from a solution; a slower method of separating out salt.



Separating out salt from rock salt:

1. Grind the mixture of rock salt.
2. Add water and stir.
3. Filter the mixture, leaving the sand in the filter paper
4. Evaporate the water from the salt, leaving the crystals.

Elements

Elements are made of atoms with the same atomic number. Atoms can be represented as symbols.

N = nitrogen **F** = fluorine **Zn** = zinc **Ca** = calcium

Isotopes – an isotope is an element with the **same number of protons** but a **different number of neutrons**. They have the same atomic number, but different mass number.

Isotope	Protons	Electrons	Neutrons
${}^1_1\text{H}$	1	1	$1 - 1 = 0$
${}^2_1\text{H}$	1	1	$2 - 1 = 1$
${}^3_1\text{H}$	1	1	$3 - 1 = 2$

Compounds – a compound is when two or more elements are chemically joined. Examples of compounds are carbon dioxide and magnesium oxide. Some examples of formulas are CO_2 , NaCl , HCl , H_2O , Na_2SO_4 . They are held together by chemical bonds and are difficult to separate.

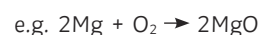
Chemical Equations

A chemical reaction can be shown by using a **word equation**.

e.g. magnesium + oxygen $\rightarrow$ magnesium oxide

On the left-hand side are the reactants, and the right-hand side are the products.

They can also be shown by a **symbol equation**.



Equations need to be **balanced**, so the same number of atoms are on each side. To do this, numbers are put in front of the compounds.

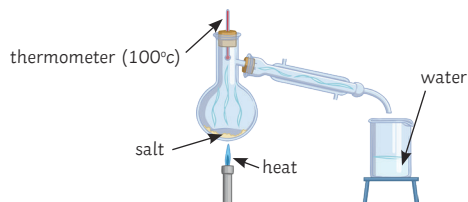


Atomic Structure and the Periodic Table – Foundation and Higher

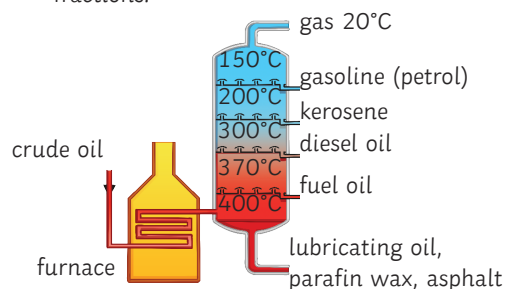
Distillation

To separate out mixtures of liquids.

1. **Simple distillation** – separating a liquid from a solution.



2. **Fractional distillation** – separating out a mixture of liquids. Fractional distillation can be used to separate out crude oil into fractions.



Metals and Non-metals

They are found at the **left** part of the periodic table. Non-metals are at the **right** of the table.

Metals

Are strong, malleable, good conductors of electricity and heat. They bond metallicity.

Non-Metals

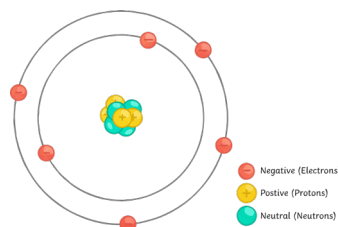
Are dull, brittle, and not always solids at room temperature.

History of the Atom

Scientist	Time	Discovery
John Dalton	start of 19 th century	Atoms were first described as solid spheres.
JJ Thomson	1897	Plum pudding model – the atom is a ball of charge with electrons scattered.
Ernest Rutherford	1909	Alpha scattering experiment – mass concentrated at the centre; the nucleus is charged. Most of the mass is in the nucleus. Most atoms are empty space.
Niels Bohr	around 1911	Electrons are in shells orbiting the nucleus.
James Chadwick	around 1940	Discovered that there are neutrons in the nucleus.

Electronic Structure

Electrons are found in shells. A maximum of two in the most inner shell, then eight in the 2nd and 3rd shell. The inner shell is filled first, then the 2nd then the 3rd shell.



Group 7 Elements and Noble Gases

Halogens

The halogens are **non-metals**: fluorine, chlorine, bromine, iodine. As you go down the group they become less reactive. It is harder to gain an extra electron because its outer shell is further away from the nucleus. The melting and boiling points also become higher.

Noble Gases

The **noble gases** (group 0 elements) include: **helium**, **neon** and **argon**. They are un-reactive as they have full outer shells, which makes them very stable. They are all colourless gases at room temperature.

The boiling points all increase as they go down the group – they have greater intermolecular forces because of the increase in the number of electrons.

Development of the Periodic Table

In the early 1800s, elements were arranged by atomic mass. The periodic table was not complete because some of the elements had not been found. Some elements were put in the wrong group.

Dimitri Mendeleev (1869) left gaps in the periodic table. He put them in order of **atomic mass**. The gaps show that he believed there was some undiscovered elements. He was right! Once found, they fitted in the pattern.

The Modern Periodic Table

Elements are in order of **atomic mass/proton number**. It shows where the metals and non-metals are. **Metals** are on the **left** and **non-metals** on the **right**. The **columns** show the **groups**. The **group number** shows the number of **electrons** in the **outer shell**. The rows are **periods** – each period shows another full shell of electrons.

The periodic table can be used to predict the reactivity of elements.

Alkali Metals

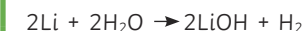
The alkali metals (**group 1** elements) are soft, very reactive metals. They all have **one electron** in their **outer shell**, making them **very reactive**. They are **low density**. As you go down the group, they become more reactive. They get bigger and it is easier to lose an electron that is further away from the nucleus.

They form ionic compounds with non-metals.

They react with water and produce hydrogen.

E.g.

lithium + water → lithium hydroxide + hydrogen



They react with chlorine and produce a metal salt.

E.g.

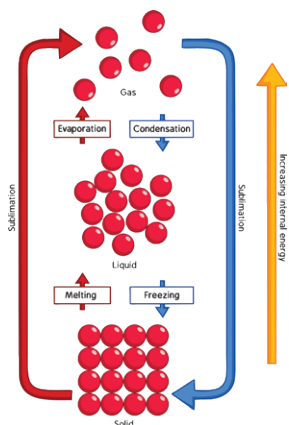
lithium + chlorine → lithium chloride



They react with oxygen to form metal oxides.



AQA GCSE Chemistry (Combined Science) Unit 2: Bonding, Structure and Properties of Matter



The three states of matter are **solid, liquid and gas**.

For a substance to change from one state to another, **energy** must be **transferred**.

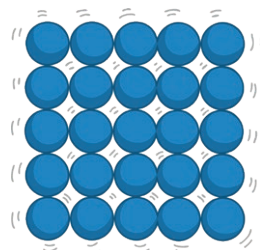
The particles gain energy. This results in the breaking of some of the **attractive forces** between particles during melting.

To evaporate or boil a liquid, more energy is needed to overcome the remaining chemical bonds between the particles.

Note the difference between **boiling** and **evaporation**. When a liquid **evaporates**, particles **leave the surface** of the liquid **only**. When a liquid **boils**, **bubbles** of gas form **throughout** the liquid before rising to the surface and escaping.

The amount of energy needed for a substance to change state is dependent upon the **strength** of the **attractive forces** between particles. The **stronger** the **forces of attraction**, the **more energy** needed to **break them apart**. Substances that have strong attractive forces between particles generally have **higher melting and boiling points**.

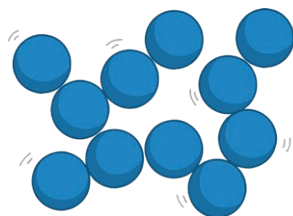
Solid



The particles in a **solid** are arranged in a regular pattern. The particles in a solid **vibrate** in a fixed position and are tightly packed together. The particles in a solid have a **low amount of kinetic energy**.

Solids have a **fixed shape** and are unable to flow like liquids. The particles **cannot be compressed** because the particles are very close together.

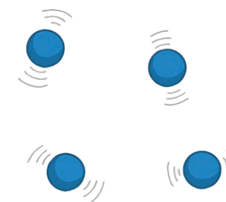
Liquid



The particles in a **liquid** are randomly arranged. The particles in a liquid are able to **move around** each other. The particles in a liquid have a **greater amount of kinetic energy** than particles in a **solid**.

Liquids are able to **flow** and can take the shape of the container that they are placed in. As with a solid, liquids **cannot be compressed** because the particles are close together.

Gas



The particles in a **gas** are randomly arranged. The particles in a gas are able to **move around very quickly** in all directions. Of the three states of matter, gas particles have the **highest amount of kinetic energy**.

Gases, like liquids, are able to **flow** and can fill the container that they are placed in. The particles in a gas are **far apart** from one another which allows the particles to move in any direction.

Gases can be **compressed**; when squashed, the particles have empty space to move into.

Limitations of the Particle Model (HT only)

The chemical bonds between particles are not represented in the diagrams above.

Particles are represented as solid spheres – this is not the case. Particles like atoms are mostly empty space. Particles are not always spherical in nature.

State Symbols

In chemical equations, the three states of matter are represented as symbols:

solid (**s**)

liquid (**l**)

gas (**g**)

aqueous (**aq**)

Aqueous solutions are those that are formed when a substance is dissolved in water.

Identifying the Physical State of a Substance

If the given temperature of a substance is **lower** than the **melting point**, the physical state of the substance will be **solid**.

If the given temperature of the substance is **between** the **melting point** and **boiling point**, the substance will be a **liquid**.

If the given temperature of the substance is **higher** than the **boiling point**, the substance will be a **gas**.



AQA GCSE Chemistry (Combined Science) Unit 2: Bonding, Structure and Properties of Matter

Formation of Ions

Ions are charged particles. They can be either positively or negatively charged, for example Na^+ or Cl^- .

When an element loses or gains electrons, it becomes an ion.

Metals **lose** electrons to become **positively charged**.

Non-metals **gain** electrons to become **negatively charged**.

Group 1 and 2 elements **lose** electrons and group 6 and 7 elements **gain** electrons.

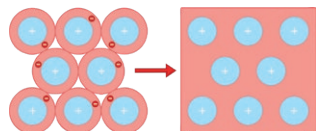
Group	Ions	Element Example
1	+1	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$
2	+2	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$
6	-2	$\text{Br} + \text{e}^- \rightarrow \text{Br}^-$
7	-1	$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$

Metals and Non-metals

Metals are found on the **left-hand side** of the **periodic table**. Metals are strong, shiny, malleable and good conductors of heat and electricity. On the other hand, non-metals are brittle, dull, not always solids at room temperature and poor conductors of heat and electricity. **Non-metals** are found on the **right-hand side** of the **periodic table**.

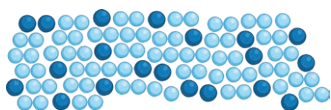
Metallic Bonding

Metallic bonding occurs between **metals only**. Positive metal ions are surrounded by a **sea of delocalised electrons**. The ions are tightly packed and arranged in rows.



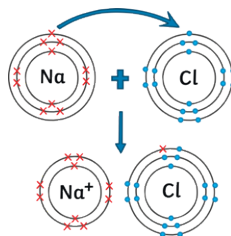
There are strong electrostatic forces of attraction between the positive metal ions and negatively charged electrons.

Pure metals are too soft for many uses and are often mixed with other metals to make alloys. The mixture of the metals introduces different-sized metal atoms. This **distorts the layers** and **prevents them from sliding over one another**. This makes it harder for alloys to be bent and shaped like pure metals.



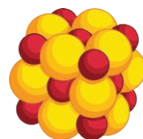
Ionic Bonding

Ionic bonding occurs between a metal and a non-metal. Metals lose electrons to become positively charged. Opposite charges are attracted by electrostatic forces – an ionic bond.



Ionic Compounds

Ionic compounds form structures called giant lattices. There are **strong electrostatic forces of attraction** that **act in all directions** and act between the **oppositely charged ions** that make up the giant ionic lattice.



Properties of Ionic Compounds

- High melting point – lots of energy needed to overcome the electrostatic forces of attraction.
- High boiling point
- **Cannot conduct electricity** in a **solid** as the ions are not free to move.
- Ionic compounds, when **molten** or in **solution**, can **conduct electricity** as the ions are free to move and can carry the electrical current.

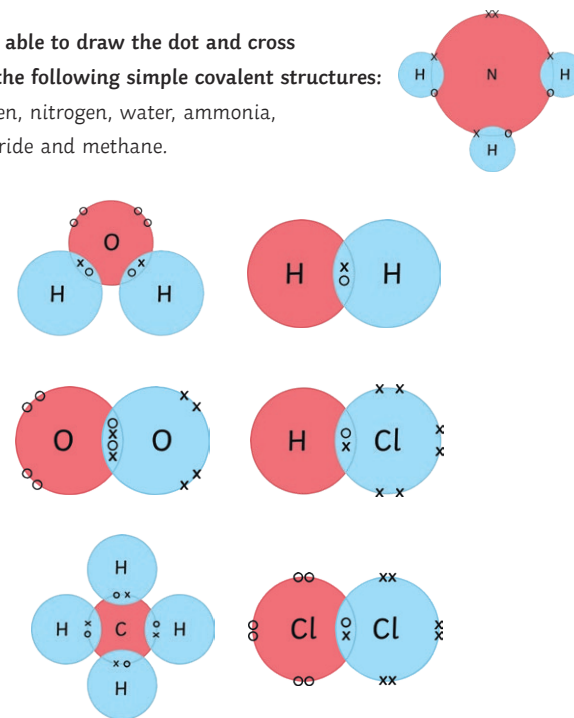
Covalent Bonding

Covalent bonding is the sharing of a pair of electrons between atoms to gain a full outer shell. This occurs between **non-metals only**. Simple covalent bonding occurs between the molecules below. Simple covalent structures have **low melting and boiling points** – this is because the **weak intermolecular forces** that hold the molecules together break when a substance is heated, not the strong covalent bonds between atoms. They **do not conduct electricity** as they do not have any free delocalised electrons.

Dot and cross diagrams are useful to show the **bonding in simple molecules**. The **outer electron shell** of each atom is represented as a **circle**, the circles from each atom overlap to show where there is a **covalent bond**, and the electrons from each atom are either drawn as **dots** or **crosses**. There are **two different types of dot and cross diagram** – one with a circle to represent the outer electron shell and one without.

You should be able to draw the dot and cross diagrams for the following simple covalent structures:

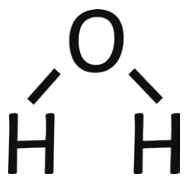
chlorine, oxygen, nitrogen, water, ammonia, hydrogen chloride and methane.



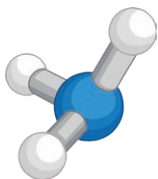
AQA GCSE Chemistry (Combined Science) Unit 2: Bonding, Structure and Properties of Matter

Structural Formulae

In this type of diagram, the element symbol represents the type of atom and the straight line represents the covalent bonding between each atom.



The structure of small molecules can also be represented as a 3D model.

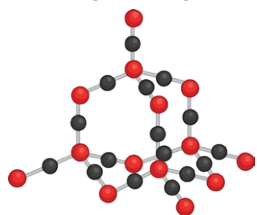


Giant Covalent Structure – Diamond

Each **carbon** atom is **bonded** to **four** other carbon atoms, making diamond very strong. Diamond has a high melting and boiling point. **Large** amounts of **energy** are needed to break the strong covalent bonds between each carbon atom. Diamond **does not conduct** electricity because it has **no free electrons**.

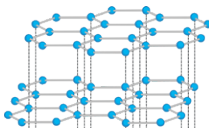


Silicon dioxide (silicon and oxygen atoms) has a similar structure to that of diamond, in that its atoms are held together by **strong covalent bonds**. Large amounts of energy are needed to break the strong covalent bonds therefore silicon dioxide, like diamond, has a high melting and boiling point.



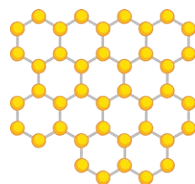
Giant Covalent Structure – Graphite

Graphite is made up of layers of **carbon** arranged in **hexagons**. Each carbon is bonded to **three** other carbons and has **one free delocalised electron** that is able to move between the layers. The layers are held together by weak intermolecular forces. The layers of carbon can slide over each other easily as there are no strong covalent bonds between the layers. Graphite has a high melting point because a lot of energy is needed to break the covalent bonds between the carbon atoms. Graphite can **conduct** electricity.



Giant Covalent Structure – Graphene

Graphene is one layer of graphite. It is very **strong** because of the covalent bonds between the carbon atoms. As with graphite, each carbon in graphene is bonded to three others with one **free delocalised electron**. Graphene is able to **conduct electricity**. Graphene, when added to other materials, can make them even stronger. Useful in electricals and composites.



Nanoscience

Nanoscience refers to structures that are **1–100nm** in size, of the order of a few hundred atoms. Nanoparticles have a **high surface area to volume ratio**. This means that smaller amounts are needed in comparison to normal sized particles. As the side length of a cube decreases by a factor of 10, the surface area to volume ratio increases approximately

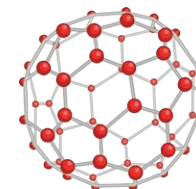
Name of Particle	Diameter
nanoparticle	1–100nm
fine particles (PM _{2.5})	100–2500nm
coarse particles (PM ₁₀)	2500–10000nm

Polymers

Polymers are long chain molecules that are made up of many smaller units called **monomers**. Atoms in a polymer chain are held together by **strong covalent bonds**. Between polymer molecules, there are **intermolecular forces**. Intermolecular forces **attract** polymer chains towards each other. Longer polymer chains have stronger forces of attraction than shorter ones therefore making stronger materials.

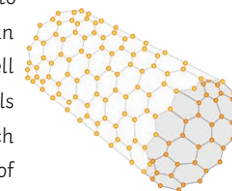
Fullerenes and Nanotubes

Molecules of carbon that are shaped like hollow tubes or balls, arranged in hexagons of five or seven carbon atoms. They can be used to **deliver drugs into the body**.



Buckminsterfullerene has the formula C₆₀

Carbon Nanotubes are tiny carbon cylinders that are very long compared to their width. Nanotubes can conduct electricity as well as strengthening materials without adding much weight. The properties of carbon nanotubes make them useful in electronics and nanotechnology.



Possible Risks of Nanoparticles

As nanoparticles are so **small**, it makes it possible for them to be inhaled and enter the lungs. Once inside the body, nanoparticles may **initiate harmful reactions** and toxic substances could bind to them because of their large surface area to volume ratio. Nanoparticles have many applications. These include medicine, cosmetics, sun creams and deodorants. They can also be used as catalysts.

Modern nanoparticles are a relatively new phenomenon therefore it is difficult for scientists to truly determine the risks associated with them.



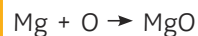
AQA GCSE Chemistry (Combined Science) Unit 5.3: Quantitative Chemistry Knowledge Organiser - Higher

Conservation of Mass

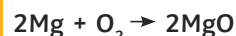
No atoms can be created or made during a chemical reaction, so the mass of the reactants will equal the mass of the product.

Reactions can be shown as a word or symbol equation.

magnesium + oxygen → magnesium oxide



Symbol equations should also be balanced; they should have the same number of atoms on each side.



Relative Formula Mass

The relative formula mass (M_r) is the sum of all the relative atomic masses (A_r) of the atoms in the formula.

Examples:

HCl

$$A_r \text{ of H} = 1$$

$$A_r \text{ of Cl} = 35.5$$

$$M_r \text{ of HCl} = 1 + 35.5 = 36.5$$

H₂SO₄

$$A_r \text{ of H} = 1$$

$$A_r \text{ of S} = 32$$

$$A_r \text{ of O} = 16$$

$$M_r \text{ of H}_2\text{SO}_4 = (1 \times 2) + 32 + (16 \times 4)$$

$$M_r \text{ of H}_2\text{SO}_4 = 2 + 32 + 64$$

$$M_r \text{ of H}_2\text{SO}_4 = 98$$

Calculating Percentage Mass of an Element in a Compound

percentage mass of an element in a compound =

$$A_r \times \frac{\text{number of atoms of that element}}{M_r \text{ of the compound}}$$

Find the percentage mass of oxygen in magnesium oxide.

$$A_r \text{ of magnesium} = 24$$

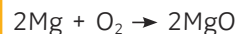
$$A_r \text{ of oxygen} = 16$$

$$M_r \text{ of MgO} = 24 + 16 = 40$$

$$\% \text{ mass} = \frac{A_r}{M_r} = \frac{16}{40} = 0.4 \quad 0.4 \times 100 = 40\%$$

Conservation of Mass

Show that mass is conserved in a reaction.



$$(2 \times 24) + (2 \times 16) \rightarrow 2(24 + 16)$$

$$48 + 32 \rightarrow 2 \times 40$$

$$80 \rightarrow 80$$

Total M_r on the left-hand side of the equation is the same as the M_r on the right-hand side.

Calculate the mass of the product.

6g of magnesium reacts with 4g of oxygen:

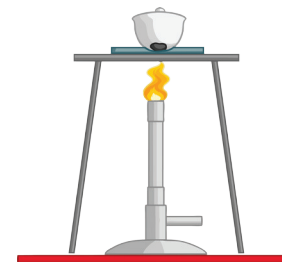
$$6 + 4 = 10\text{g of magnesium oxide}$$

During a reaction the mass can change. If one of the reactants is a gas, the mass can go up.

E.g.



Oxygen from the air is added to the magnesium (making the product) which will be heavier in mass.

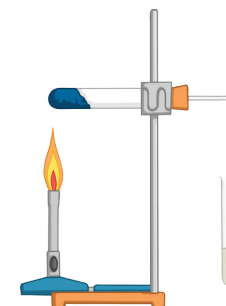


If one of the products is a gas, the mass can go down.

E.g.



When sodium carbonate is thermally decomposed, carbon dioxide gas is produced and released into the atmosphere.



Concentration of Solutions

Concentration is the amount of a substance in a specific volume of a solution. The more substance that is dissolved, then the more concentrated the solution is.

It is possible to calculate the concentration of a solution with the following equation:

$$\text{concentration (g/dm}^3\text{)} = \text{mass (g)} \div \text{volume of solvent (dm}^3\text{)}$$

The equation can be rearranged to find the mass of the dissolved substance:

$$\text{mass (g)} = \text{concentration (g/dm}^3\text{)} \times \text{volume (dm}^3\text{)}$$



The Mole

The Avogadro constant, 6.02×10^{23} , is the number of molecules of a substance that make up one mole of that substance.

Iron has an A_r of 56, so 1 mole of iron has a mass of 56g.

Oxygen (O_2) gas has an M_r of 32, so 1 mole of oxygen has a mass of 32g.

Ammonia (NH_3) has an M_r of 17, so 1 mole of ammonia has a mass of 17g.

$$\text{number of moles} = \frac{\text{mass in g (of an element or compound)}}{M_r \text{ (of the element or compound)}}$$

Moles and Equations

Write a balanced symbol equation for the reaction in which 5.6g of iron reacts with 10.65g of chlorine to form iron chloride.

Work out the M_r of all the substances.

A_r of Fe = 56 and A_r of Cl = 35.5

Divide the mass of each substance by its M_r to calculate how many moles of each substance reacted or produced.

$$\text{moles Fe} = 5.6/56 = 0.1$$

$$\text{moles Cl} = 10.65/35.5 = 0.3$$

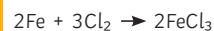
Divide by the smallest number of moles

$$\text{Fe} = \frac{0.1}{0.1} = 1 \qquad \text{Cl} = \frac{0.3}{0.1} = 3$$

Write down the balanced symbol equation.



Chlorine exists as Cl_2 so the whole thing must be multiplied by 2.



Limiting Reactions

If one reactant gets used up in a reaction before the other, then the reaction will stop. The reactant that has been used up is limiting.

If you halve the amount of reactant then the amount of product will also be halved.



AQA GCSE Chemistry (Combined Science) Unit 4: Chemical Changes Knowledge Organiser

The Reactivity Series

Here's a mnemonic to help you learn the order:

purple (potassium)

slime (sodium)

can (calcium)

make (magnesium)

a (aluminium)

careless (carbon)

zebra (zinc)

insane (iron)

try (tin)

learning (lead)

how (hydrogen)

camels (copper)

surprise (silver)

gorillas (gold)

	potassium
	sodium
	calcium
	magnesium
	aluminium
carbon	zinc
	iron
	tin
	lead
hydrogen	copper
	silver
	gold
	platinum

The reactivity series is a league table for metals. The more reactive metals are near the top of the table with the least reactive near the bottom. In chemical reactions, a more reactive metal will displace a less reactive metal.

Reactions of Metals with Water

Metals, when reacted with water, produce a metal hydroxide and hydrogen.

lithium + water $\rightarrow$ lithium hydroxide + hydrogen

$2\text{Li} + 2\text{H}_2\text{O} \rightarrow 2\text{LiOH} + \text{H}_2$

The more reactive a metal is, the faster the reaction.

Reactions of Metals with Dilute Acid

Metals, when reacted with acids, produce a salt and hydrogen.

Sodium + hydrochloric acid $\rightarrow$ sodium chloride + hydrogen

$2\text{Na} + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2$

Metals that are below hydrogen in the reactivity series **do not** react with dilute acids.

Reactions of Acids

The general formula for the reaction between an acid and a metal is:
acid + metal $\rightarrow$ salt + hydrogen

For example: hydrochloric acid + sodium $\rightarrow$ sodium chloride + hydrogen

$2\text{HCl} + 2\text{Na} \rightarrow 2\text{NaCl} + \text{H}_2$

When an acid reacts with an alkali, a neutralisation reaction takes place and a salt and water are produced.

The general formula for this kind of reaction is as follows:

acid + alkali $\rightarrow$ salt + water

hydrochloric acid + sodium hydroxide $\rightarrow$ sodium chloride + water

$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

Naming Salts

The first part comes from the metal in the metal carbonate, oxide or hydroxide. The second part of the name comes from the acid that was used to make it.

Acid Used	Salt Produced
hydrochloric	chloride
nitric	nitrate
sulfuric	sulfate

For example, sodium chloride.

Redox Reactions (Higher Tier Only)

When metals react with acids, they undergo a redox reaction. A **redox reaction** occurs when both **oxidation** and **reduction** take place at the same time.

For example:

$2\text{H}^+ + \text{Ca} \rightarrow \text{Ca}^{2+} + \text{H}_2$

The ionic equation can be further split into two half equations.

$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$

Oxidation is loss of electrons.

$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$

Reduction is gaining of electrons.

Reactions with Bases

The general formula for the reaction between an acid and a metal oxide is:
acid + metal oxide $\rightarrow$ salt + water

sulfuric acid + copper oxide $\rightarrow$ copper sulfate + water

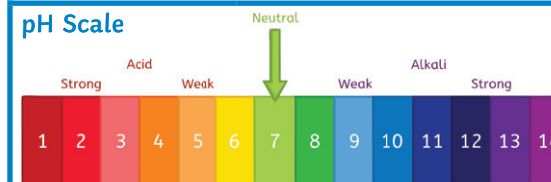
$\text{H}_2\text{SO}_4 + \text{CuO} \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}$

Reactions with Carbonates

The general formula for the reaction between an acid and a carbonate is:
acid + carbonate $\rightarrow$ salt + water + carbon dioxide

hydrochloric acid + calcium carbonate $\rightarrow$ calcium chloride + water + carbon dioxide

pH Scale



In aqueous solutions, acids produce H^+ ions and alkalis produce OH^- ions.

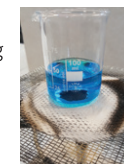
Neutral solutions are pH7 and are neither acids or alkalis.

For example, in neutralisation reactions, hydrogen ions from an acid react with hydroxide ions from an alkali to produce water:

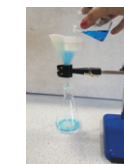
$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$

Making Soluble Salts

1. Make a saturated solution by stirring copper oxide into the sulfuric acid until no more will dissolve.



2. Filter the solution to remove the excess copper oxide solid.



3. Half fill a beaker with water and set this over a Bunsen burner to heat the water. Place an evaporating dish on top of the beaker.



4. Add some of the solution to the evaporating basin and heat until crystals begin to form.



5. Once cooled, pour the remaining liquid into a crystallising dish and leave to cool for 24 hours.



6. Remove the crystals with a spatula and pat dry between paper towels.



Strong and Weak Acids (Higher Tier Only)

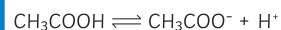
A **strong** acid **completely dissociates** in a solution. For example: $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

Hydrochloric acid is able to completely dissociate in solution to form hydrogen and chloride ions.

Examples of strong acids include nitric acid (HNO_3) and sulfuric acid (H_2SO_4).

Weak acids in comparison only partially dissociate.

For example acetic acid **partially dissociates** to form a hydrogen and acetate ion.



The **double arrow** symbol indicates that the reaction is **reversible**. Both the forward and reverse reaction occur at the same time and the reaction never goes to completion.

The Process of Electrolysis

Electrolysis is the **splitting up** of an ionic substance using **electricity**.

On setting up an electrical circuit for electrolysis, two **electrodes** are required to be placed in the electrolyte. The electrodes are **conducting rods**. One of the rods is connected to the **positive** terminal and the other to the **negative** terminal.

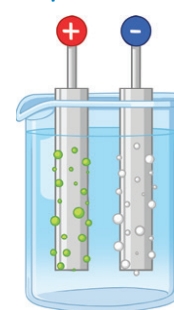
The **electrodes** are **inert** (this means they do not react in the reaction) and are often made from **graphite** or **platinum**.

During the process of electrolysis, **opposites attract**. The positively charged ions will be attracted toward the negative electrode. The negatively charged ions will be attracted towards the positive electrode.

When ions reach the electrodes, the charges are lost and they become elements.

The **positive** electrode is called the **anode**.

The **negative** electrode is called the **cathode**.

Electrolysis of Aqueous Solutions

Gases may be given off or metals deposited at the electrodes. This is dependent on the reactivity of the elements involved.

If the metal is **more reactive** than **hydrogen** in the reactivity series, then **hydrogen** will be **produced** at the **negative cathode**. At the **positive anode**, negatively charged ions **lose** electrons. This is called **oxidation** and you say that the ions have been oxidised.

Using Electrolysis to Extract Metals

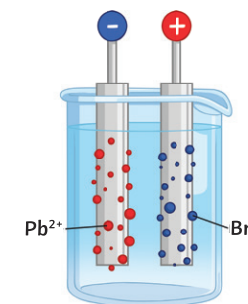
Metals are extracted by electrolysis if the metal in question reacts with carbon or if it is too reactive to be extracted by reduction with carbon. During the extraction process, large quantities of energy are used to melt the compounds.

Aluminium is manufactured by the process of electrolysis. Aluminium oxide has a high melting point and melting it would use large amounts of energy. This would increase the cost of the process, therefore molten **cryolite** is added to aluminium oxide to lower the melting point and thus reduce the cost.

Electrolysis of Molten Ionic Compounds – Lead Bromide

Lead bromide is an **ionic** substance. Ionic substances, when solid, are **not** able to conduct electricity. When molten or in solution, the ions are free to move and are able to carry a charge.

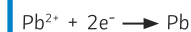
The **positive** lead ions are attracted toward the **negative cathode** at the same time as the **negative bromide** ions are attracted toward the **positive anode**.



Oxidation is the **loss** of electrons and reduction is the **gaining** of electrons. **OIL RIG** (Higher Tier Only).

We represent what is happening at the electrodes by using **half equations** (Higher Tier Only).

The lead ions are attracted towards the negative electrode. When the **lead ions** (Pb^{2+}) reach the cathode, each ion **gains two electrons** and becomes a neutral atom. We say that the lead ions have been **reduced**.



The bromide ions are attracted towards the positive electrode. When the **bromide ions** (Br^-) reach the anode, each ion **loses one electron** to become a neutral atom. Two bromine atoms are then able to bond together to form the covalent molecule Br_2 .



AQA GCSE Chemistry (Combined) Unit 5 Energy Changes Knowledge Organiser

Exothermic and Endothermic Reactions

When a chemical reaction takes place, **energy** is involved. Energy is transferred when chemical **bonds are broken** and when new **bonds are made**.

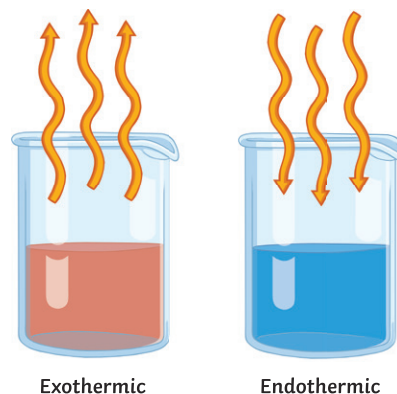
Exothermic reactions are those which involve the transfer of energy **from the reacting chemicals** to the surroundings. During a practical investigation, an exothermic reaction would show an **increase in temperature** as the reaction takes place.

Examples of exothermic reactions include **combustion, respiration and neutralisation** reactions. Hand-warmers and self-heating cans are examples of everyday exothermic reactions.

Endothermic reactions are those which involve the transfer of energy **from the surroundings** to the reacting chemicals. During a practical investigation, an endothermic reaction would show a **decrease in temperature** as the reaction takes place.

Examples of endothermic reactions include the **thermal decomposition** of calcium carbonate.

Eating **sherbet** is an everyday example of an endothermic reaction. When the sherbet dissolves in the saliva in your mouth, it produces a cooling effect. Another example is **instant ice packs** that are used to treat sporting injuries.



Activation Energy – the minimum amount of energy required for a chemical reaction to take place.

Catalysts – increase the rate of a reaction. Catalysts provide an alternative pathway for a chemical reaction to take place by **lowering** the activation energy.

Bond Making and Bond Breaking

In an **endothermic** reaction, energy is needed to break chemical bonds. The **energy change (ΔH)** in an endothermic reaction is **positive**.

You may also find, in some textbooks, ΔH referred to as the **enthalpy change**.

In an **exothermic** reaction, energy is needed to form chemical bonds. The **energy change (ΔH)** in an exothermic reaction is **negative**.

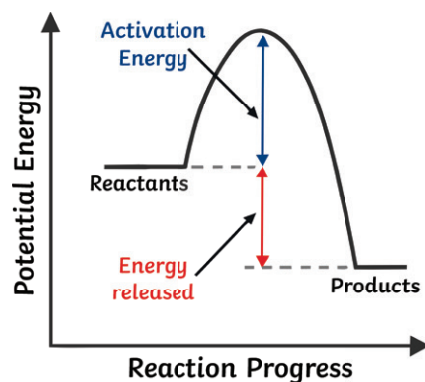
Bond energies are measured in **kJ/mol**.

Reaction Profiles – Exothermic

Energy level diagrams show us what is happening in a particular chemical reaction. The diagram shows us the **difference in energy** between the reactants and the products.

In an exothermic reaction, the **reactants** are at a **higher** energy level than the products.

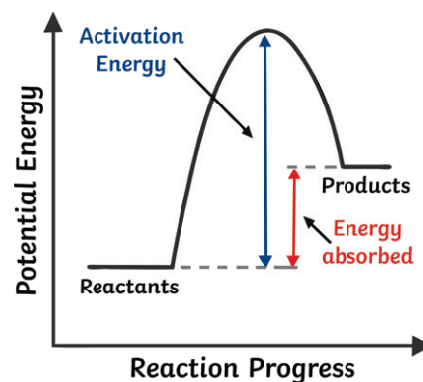
In an **exothermic** reaction, the difference in energy is **released** to the surroundings and so the **temperature** of the surroundings **increases**.



Reaction Profiles – Endothermic

In an **endothermic** reaction, the **reactants** are at a **lower** energy level than the products.

In an **endothermic** reaction, the difference in energy is **absorbed** from the surroundings and so the **temperature** of the surroundings **decreases**.



Calculations Using Bond Energies (Higher Tier Only)

Bond energies are used to calculate the change in energy of a chemical reaction.

Calculate the change in energy for the reaction: $2\text{H}_2\text{O}_2 \longrightarrow 2\text{H}_2\text{O} + \text{O}_2$

The first step is to write the symbol equation for the reaction.

Once you have done this, work out the bonds that are breaking and the ones that are being made.



Bond	Bond Energy kJ/mol
H-O	464
O-O	146
O=O	498

On the **left-hand side** of the equation, the **bonds are breaking**.

There are two **O-H** bonds and one **O-O** bond.

$$\text{So } 464 + 146 + 464 = 1074$$

There are two moles of H_2O_2 therefore the answer needs to be multiplied by two.

$$\text{So } 1074 \times 2 = 2148$$

On the **right-hand side** of the equation, the **bonds are made**.

There are two **H-O** bonds

$$\text{So } 464 + 464 = 928$$

Two moles of H_2O are made therefore the answer needs to be multiplied by two.

$$\text{So } 928 \times 2 = 1856$$

There is also one **O=O** bond with a bond energy of 498

$$\text{So } 1856 + 498 = 2354$$

$$\Delta H = \text{sum (bonds broken)} - \text{sum (bonds made)}$$

$$\Delta H = 2148 - 2354 = -206 \text{ kJ/mol}$$

The reaction is exothermic as ΔH is negative.

Required Practical**Aim**

To investigate the variables that affect temperature changes in reacting solutions, e.g. acid plus metals, acid plus carbonates, neutralisations and displacement of metals.

Equipment

- polystyrene cup
- measuring cylinder
- thermometer
- 250cm³ glass beaker
- measuring cylinder
- top pan balance

Method

1. Gather the equipment.
2. Place the polystyrene cup inside the beaker. This will prevent the cup from falling over.
3. Using a measuring cylinder, measure out 30cm³ of the acid. Different acids such as hydrochloric or sulfuric acid may be used. Pour this into the polystyrene cup.
4. Record the temperature of the acid using a thermometer.
5. Using a top pan balance, measure out an appropriate amount of the solid (for example, 10g) or use one strip of a metal such as magnesium.
6. Add the solid to the acid and record the temperature. You may choose to record the temperature of the acid and metal every minute for 10 minutes.

