Chemistry

LET ALL OUR BRIGHT COLOURS SHINE

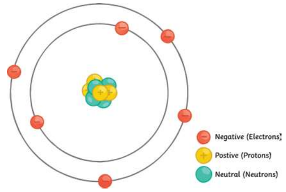
Year 10 Knowledge Organisers

Atomic Structure and the Periodic Table – Foundation and Higher (Separate)

Atoms

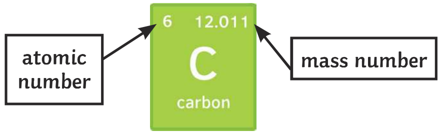
Contained in the nucleus are the **protons** and **neutrons**. Moving around the nucleus are the **electron** shells. They are negatively charged.

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) =

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

Balancing Symbol Equations

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

$CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$

C = 1

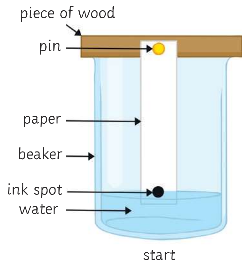
O = 4

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.

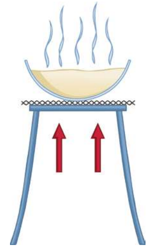
Chromatography – to separate out mixtures.



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
$^{1}_{1}H$	1	1	1 - 1 = 0
$^{2}_{1}H$	1	1	2 - 1 = 1
$^{3}_{1}H$	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 → 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.**

e.g. $2Mg + O_2 \rightarrow 2MgO$

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.

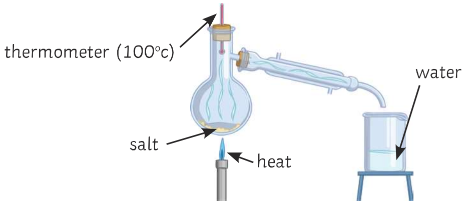
$CH_4 + \mathbf{2}O_2 \rightarrow \mathbf{2}H_2O + CO_2$

Atomic Structure and the Periodic Table – Foundation and Higher (Separate)

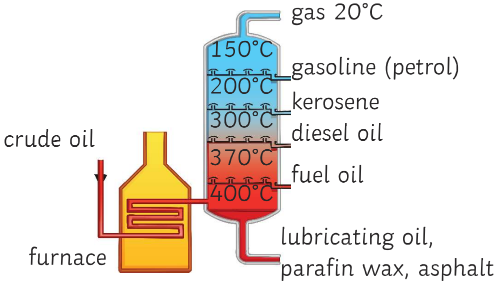
Distillation

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 metallically.

Non-Metals

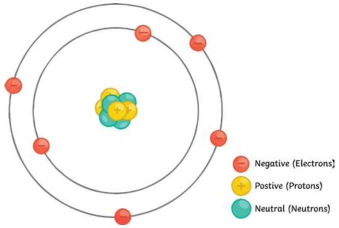
Are dull, brittle, and not always solids at room temperature.

History of the Atom

Scientist	Time	Discovery
John Dalton	start of 19th 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

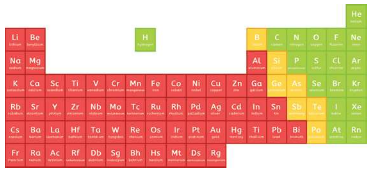
In the early 1800s, elements were arranged by atomic mass. The periodic table was not complete because some of the elements had not been found. Some elements were put in the wrong group.

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

$2Li + 2H_2O \rightarrow 2LiOH + H_2$

They react with chlorine and produce a metal salt.

E.g.

lithium + chlorine ➔ lithium chloride

$2Li + Cl_2 \rightarrow 2LiCl$

They react with oxygen to form metal oxides.

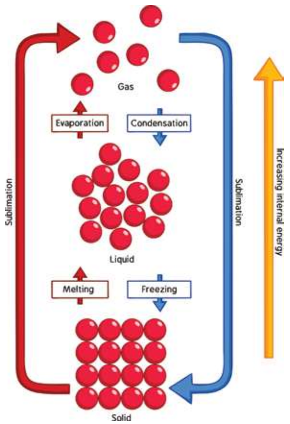
The Transition Metals

The transition metals are a block of elements found between groups 2 and 3 in the middle of the periodic table. Examples of transition metals include copper, nickel and iron with many more included. They have all the properties you would expect metals to have, such as being strong, shiny and conductors of electricity and heat. Transition metals make very good catalysts; this means they speed up a reaction without being used up themselves. Iron is used as a catalyst during the Haber process when making ammonia.

Transition metals can form more than one ion. For example, copper can take the form of Cu^{+}, Cu^{2+} and iron can be Fe^{2+} and Fe^{3+}. The ions are often coloured and the compounds they are found in are also coloured.

																	He helium
Li lithium	Be beryllium					H hydrogen						B boron	C carbon	N nitrogen	O oxygen	F fluorine	Ne neon
Na sodium	Mg magnesium											Al aluminium	Si silicon	P phosphorous	S sulfur	Cl chlorine	Ar argon
K potassium	Ca calcium	Sc scandium	Ti titanium	V vanadium	Cr chromium	Mn manganese	Fe iron	Co cobalt	Ni nickel	Cu copper	Zn zinc	Ga gallium	Ge germanium	As arsenic	Se selenium	Br bromine	Kr krypton
Rb rubidium	Sr strontium	Y yttrium	Zr zirconium	Nb niobium	Mo molybdenum	Tc technetium	Ru ruthenium	Rh rhodium	Pd palladium	Ag silver	Cd cadmium	In indium	Sn tin	Sb antimony	Te tellurium	I iodine	Xe xenon
Cs caesium	Ba barium	La lanthanum	Hf hafnium	Ta tantalum	W tungsten	Re rhenium	Os osmium	Ir iridium	Pt platinum	Au gold	Hg mercury	Ti thallium	Pb lead	Bi bismuth	Po polonium	At astatine	Rn radon
Fr francium	Ra radium	Ac actinium	Rf rutherfordium	Db dubnium	Sg seaborgium	Bh bohrium	Hs hassium	Mt meitnerium	Ds darmstadtium	Rg roentgenium							

AQA GCSE Chemistry (Separate 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**.

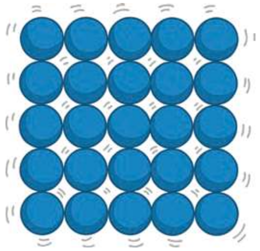
The particles gain energy. This results in the breaking of some of the **attractive forces** between particles during melting.

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**.

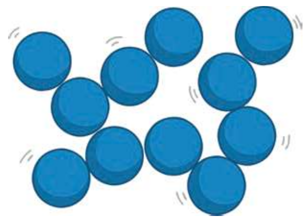
Solid



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.

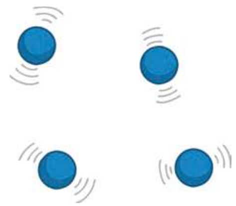
Liquid



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 (Separate Science) Unit 2: Bonding, Structure and Properties of Matter

Formation of Ions

Ions are charged particle. 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.

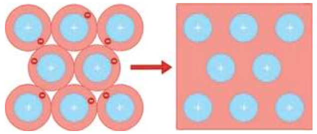
Group	Ions	Element Example
1	+1	$Li \rightarrow Li^+ + e^-$
2	+2	$Ca \rightarrow Ca^{2+} + 2e^-$
6	-2	$Br + e^- \rightarrow Br^-$
7	-1	$O + 2e^- \rightarrow O^{2-}$

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 electron**s. The ions are tightly packed and arranged in rows.

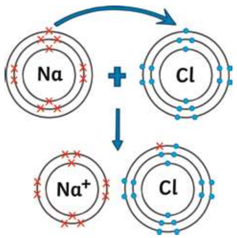


There are strong electrostatic forces of attraction between the positive metal ions and negatively charged electrons.

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.

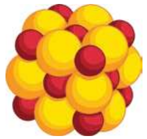
Ionic Bonding

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

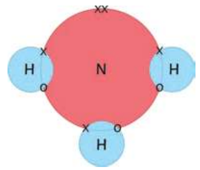
- High melting point – lots of energy needed to overcome the electrostatic forces of attraction.
- High boiling point
- **Cannot conduct electricity** in a **solid** as the ions are not free to move.
- Ionic compounds, when **molten** or in **solution**, can **conduct electricity** as the ions are free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

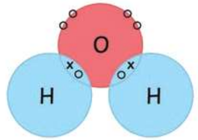
Covalent Bonding

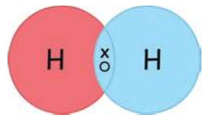
Covalent bonding is the sharing of a pair of electrons between atoms to gain a full outer shell. This occurs between **non-metals only**. Simple covalent bonding occurs between the molecules below. Simple covalent structures have **low melting and boiling points** – this is because the **weak intermolecular 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.

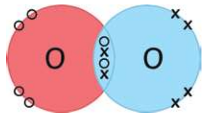
Dot and cross diagrams are useful to show the **bonding** in **simple molecules**. The **outer electron shell** of each atom is represented as a **circle**, the circles from each atom overlap to show where there is a **covalent bond**, and the electrons from each atom are either drawn as **dots or crosses**.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

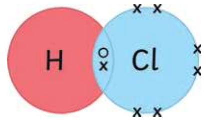
You should be able to draw the dot and cross diagrams for the following simple covalent structures: chlorine, oxygen, nitrogen, water, ammonia, hydrogen chloride and methane.

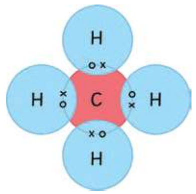


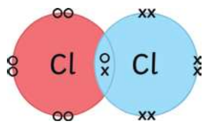








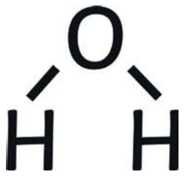




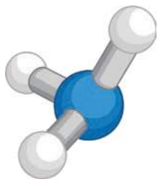
AQA GCSE Chemistry (Separate Science) Unit 2: Bonding, Structure and Properties of Matter

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.

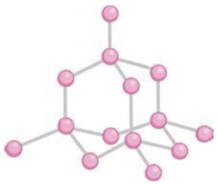


The structure of small molecules can also be represented as a 3D model.

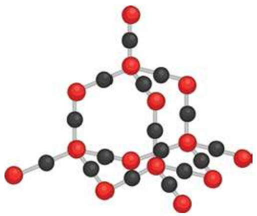


Giant Covalent Structure – Diamond

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**.

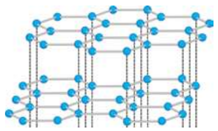


Silicon dioxide (silicon and oxygen atoms) has a similar structure to that of diamond, in that its atoms are held together by **strong covalent bonds**. Large amounts of energy are needed to break the strong covalent bonds therefore silicon dioxide, like diamond, has a high melting and boiling point.



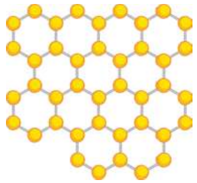
Giant Covalent Structure – Graphite

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

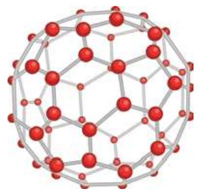
Name of Particle	Diameter
nanoparticle	1–100nm
fine particles ($PM_{2.5}$)	100–2500nm
coarse particles (PM_{10})	2500–10000nm

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.

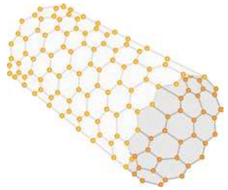
Fullerenes and Nanotubes

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 C_{60}

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 (Separate Science) Unit 3: Quantitative Chemistry

Relative Formula Mass (M_r)

The **relative atomic mass (A_r)** of an element is an elements relative mass compared to the mass of an atom of carbon-12. A_r values are given in the periodic table.

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

Example 1: hydrochloric acid (HCl) consists of one hydrogen atom (A_r1) and one chlorine atom (A_r 35.5).

The M_r of HCl = 1 + 35.5 = 36.5

Example 2: sulfuric acid (H_2SO_4) consists of two hydrogen atoms (A_r 1), one sulfur atom (A_r 32) and four oxygen atoms (A_r 16).

The M_r of H_2SO_4 = (1 × 2) + 32 + (16 × 4) = 98

Neither A_r or M_r values have any units.

Law of Conservation of Mass

The **law of conservation of mass** states that during a chemical reaction, no atoms are lost or made.

For example: $2Mg + O_2 \longrightarrow 2MgO$

In a chemical reaction, mass is never lost or gained. What **goes in** must **come out**. The **total mass of the reactants** at the beginning of the chemical reaction **equals** the **total mass of the products** made at the end of the reaction.

For example, imagine if we used building bricks to represent the atoms in a chemical reaction: atoms, like building bricks, can be completely rearranged. However, the total mass of the atoms will stay the same. Rearranging the building blocks in different structures takes a little **energy**, just like in a chemical reaction.

Reactions in Closed and Non-Enclosed Systems

If a reaction occurs in a **closed system**, the **mass** in a chemical reaction will remain **constant**.

In an **non-enclosed system, changes in mass can occur,** such as when a gas is released. It is important to remember that **no atoms are created or destroyed,** they are just **rearranged**. If a gas escapes a non-enclosed system, the total mass will look as if it has decreased. Similarly, if a gas is gained, the total mass will look as if it has increased. However, the **total mass** will **remain the same** if the mass of the gas is included in the reaction calculation.

In this **closed system** (the classroom), the mass in the reaction remains constant. As the system is a closed one, no children are allowed to leave or enter.

In this **non-enclosed system** (the classroom), the mass in the reaction can look as if it has changed as children are allowed to leave the classroom at any time.

Uncertainty

Whenever a measurement is made, there is always some degree of **uncertainty** about the result. Uncertainty is a **measure** of the **variability** in scientific data.

Uncertainty can be measured by considering the **resolution** of the scientific equipment being used or from the **range** of a set of scientific data.

There are two types of errors: **random error** and **systematic error.**

Random errors may be caused by **human error** such as a poor technique when taking measurements or by **equipment** that is **faulty**. For example, three mass balances all showing different mass values for the same object. Random errors are **random** and not something that can be predicted.

Systematic errors are errors that are produced **consistently**: if the experiment is repeated, the **same error** will occur. For example, not taring a mass balance properly or problems with the experimental method.

$$\text{uncertainty} = \frac{\text{range of results}}{2}$$

The **range** is the difference between the **largest** and **smallest** value.

For example, student A carried out a practical to determine how much dilute sulfuric acid is needed to react with exactly $50.0cm^3$ of a sodium hydroxide solution.

Repeat	1	2	3	Mean
Volume of H_2SO_4 needed to react with $50.0cm^3$ of NaOH.	23.13	24.00	23.56	23.56

Calculate the range:

range = 24.00 – 23.13 = $0.87cm^3$

Calculate the uncertainty of the mean:

uncertainty = 0.87 ÷ 2 = $0.44cm^3$

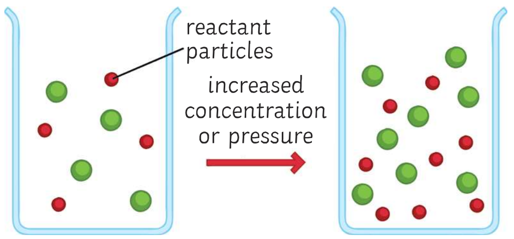
The mean with uncertainty:

23.56 ± $0.44cm^3$

AQA GCSE Chemistry (Separate Science) Unit 3: Quantitative Chemistry

Concentration of Solutions

Concentration is a **measure** of the amount of a **substance** in a **volume** of liquid. The higher the concentration, the more particles of a substance are present in the solution.



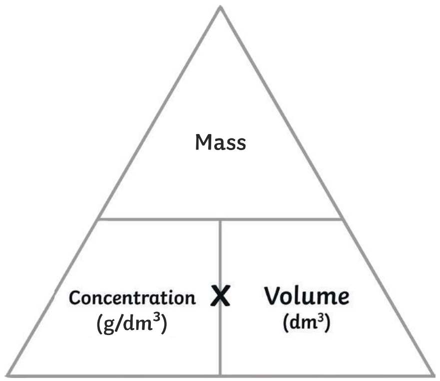
In chemistry, there are two ways to measure the concentration of a solution. This can be done by calculating the **mass** of the substance in grams or by calculating the number of **moles**.

In order to calculate concentration, you must be working in dm^3.

If it is not, it may mean that you need to do a conversion.

$cm^3 \longrightarrow dm^3 = \div 1000$

$m^3 \longrightarrow dm^3 = \times 1000$



Calculate the **concentration** of a solution with a mass of 2.15g and a volume of $5dm^3$.

concentration = mass ÷ volume

concentration = 2.15g ÷ $5dm^3$

concentration = $0.43g/dm^3$

Calculate the **mass** of sodium chloride that you would need to dissolve in $400cm^3$ of water to make a $20g/dm^3$ volume solution.

mass = concentration × volume

convert $cm^3 \longrightarrow dm^3$

$400cm^3 \div 1000 = 0.40dm^3$

mass = $20g/dm^3 \times 0.40dm^3$ = 8g

Calculate the **volume** of liquid required to add to 8.80g of a solid to make $42g/dm^3$ solution.

volume = mass ÷ concentration

volume = 8.80g ÷ $42g/dm^3$

concentration = $0.210dm^3$

The Mole – Higher Tier Only

When we talk about moles, we are not talking about the moles that live underground.

A **mole** (mol) is a **measurement** that is used in chemistry.

Example 1

Look at this reaction:

$2Mg + O_2 \longrightarrow 2MgO$

The reaction shows that **two moles** of magnesium react with oxygen to produce **two moles** of magnesium oxide. Using moles in a **balanced symbol equation** shows the **ratio** of **reactants** to **products**.

Avogadro's Constant

1 mole = 6.02×10^{23}

The number is known as **Avogadro's constant or Avogadro's number** and is named after the Italian scientist Amedeo Avogadro. The mole is abbreviated to **mol**.

This number is very important and one that you should remember. The mass of one mole of a substance in grams is equal to its relative formula mass. For example, one mole of carbon-12 has a mass of 12g

A mole is the amount of a substance that contains **6.02×10^{23}** particles of that substance. The particles could be atoms, molecules, ions or electrons.

For example, 1 mole of carbon will contain the same number of atoms **(6.02×10^{23})** as you would have molecules in 1 mole of water.

Calculating the Number of Particles

The number of particles can be calculated using Avogadro's constant if the number of moles is known.

In chemistry, Avogadro's constant is given the symbol N_A. To calculate the number of particles in a substance, the following equation can be used:

$N = n \times N_A$

N = the number of particles in a substance

n = the number of moles (mol)

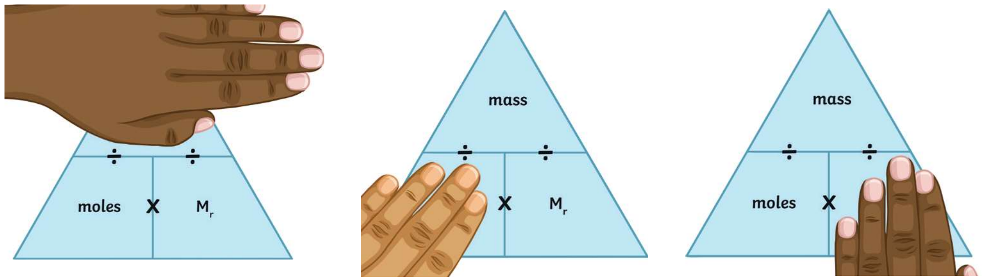
N_A = Avogadro's constant 6.02×10^{23}

For example, calculate the number of helium molecules in 10 mol of helium.

$N = n \times N_A$

N = $10 \times (6.02 \times 10^{23}) = 6.022 \times 10^{24}$

Calculating Moles, Mass and M_r



Calculating Moles, Mass and M_r

Calculate the number of **moles** in 330g of K_2S.

K_2S consists of two potassium atoms (A_r 39) and one sulfur atom (A_r 32).

Calculate the M_r of the compound = $(39 \times 2) + 32 =$ **110**

moles = mass ÷ M_r

moles = 330 ÷ 110 = 3 moles

Calculate the **mass** of 0.9 moles of $Fe(NO_3)_3(H_2O)_9$.

Calculate the Mr of the compound.

$(16 \times 3) + 14 = 62$

$62 \times 3 =$ **186**

$(1 \times 2) + 16 = 18$

$18 \times 9 =$ **162**

$56 + 186 + 162 =$ **404**

mass = moles × M_r

mass = 0.9 × 404 = 363.6g

Relative Atomic Mass (A_r)

iron (Fe) = 56

oxygen (O) = 16

nitrogen (N) = 14

hydrogen (H) = 1

Amount of Substances in Equations – Higher Tier Only

How do we know the masses involved in the equation?

To work out the masses involved, write in the relative atomic mass (A_r) for an element and the relative formula mass (M_r) for a compound.

Example

Step 1: Write down the **balanced** symbol equation. $2Mg + O_2 \longrightarrow 2MgO$

Step 2: Write in the relative atomic and relative formula masses for the **reactants** and **products** involved in the chemical reaction.

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

$48g + 32g \longrightarrow 80g$

Masses in Equations

$Fe + 2HCl \longrightarrow FeCl_2 + H_2$

$56 + 2\ (1 + 35.5) \longrightarrow (56 + 2(35.5) + (1 \times 2))$

$56g + 73g \longrightarrow 127g + 2g$

One mole of iron **reacts** with two moles of hydrochloric acid to **produce** one mole of iron chloride and one mole of hydrogen.

Calculate the **mass of water** made when burning **300g of methane**.

Step 1: Balance the equation.

$CH_4 + 2O_2 \longrightarrow 2H_2O + CO_2$

Step 2: Write down the relative formula mass of each compound.

$(12 + (1 \times 4)) + 2(16 \times 2) \longrightarrow 2((1 \times 2) + 16) + (12 + (16 \times 2))$

$16g + 64g \longrightarrow 36g + 44g$

Relative Atomic Mass (A_r)

Carbon (C) = 12

oxygen (O) = 16

hydrogen (H) = 1

We know from the equation that **16g of methane** reacts to produce **36g of water**.

The question asks us to calculate the mass of water made when burning **300g of methane**.

$\frac{\text{known mass}}{M_r} \times M_r \text{ of unknown mass}$ $\qquad$ $\frac{300}{16} \times 36 = 675g$

Limiting Reactants

A chemical reaction ends once one of the **reactants** is used up. The other reactants have nothing to react with and so some are left over.

The **limiting reactant** is the reactant that is **completely used up** in a chemical reaction. This reactant is the one that determines the amount of product that is made.

The reactant in **excess** is the one that is left over and could further react if there was another reactant to react with.

The **amount of product** that is produced during a chemical reaction is **dependent** upon the **amount of the limiting reactant**.

Calculating the maximum mass of a product formed during a chemical reaction can be done by the following:

a. Writing a balanced equation.

b. Calculating the mass (g) of the limiting reactant.

c. The A_r and M_r of the product and limiting reactant.

Determine the **maximum mass of hydrogen** that can be produced when 36g of magnesium (Mg A_r 24) reacts completely with excess hydrochloric acid (HCl) to produce magnesium chloride ($MgCl_2$) and hydrogen (H_2).

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

number of moles = mass ÷ A_r

number of moles = 36 ÷ 24 = 1.5 mol

From the equation, 1 mol of magnesium forms 1 mol of hydrogen. Therefore, 1.5 mol of magnesium forms 1.5 mol of hydrogen.

mass of hydrogen = M_r × number of moles

= 2 × 1.5

= 3g

Balancing Equations

By using the masses of the products and reactants, it is possible to work out the balancing numbers in an equation.

For example, 12g of magnesium (Mg A_r 24) reacts with 8g of oxygen (O_2 M_r 32) to produce magnesium oxide (MgO M_r 40). Determine the balanced symbol equation for the reaction.

Calculate the amount of each of the reactants.

Mg = 12 ÷ 24 = 0.5 mol

O_2 = 8 ÷ 32 = 0.25 mol

Divide both values by the smaller amount.

Mg = 0.5 ÷ 0.25 = 2

O_2 = 0.25 ÷ 0.25 = 1

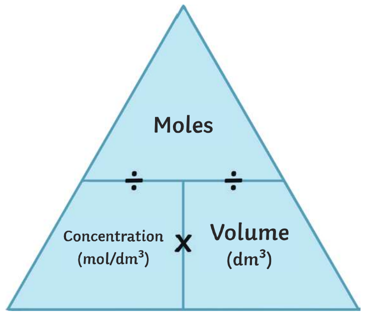
The equation shows that on the left-hand side of the equation, 2 mol of the reactant (Mg) reacts with 1 mol of oxygen. Using this information, it is then possible to balance the rest of the equation in the normal way.

$$2Mg + O_2 \longrightarrow 2MgO$$

Calculating Concentrations

The concentration of a solution can have the units g/dm³ or mol/dm³.

Concentration can be calculated using the mass of dissolved solute or the volume of the solvent or solution in dm³.



Example:

Student A dissolved 1 mol of sodium hydroxide in 4dm³ of water. Determine the concentration of the sodium hydroxide solution he made.

concentration = 1 mol ÷ 4dm³

concentration = 0.25mol/dm³

Converting between Units

To convert between g/dm³ and mol/dm³, the relative formula mass of the solute is used.

Multiply by the **M_r** to convert from **mol/dm³ to g/dm³**.

Divide by the **M_r** to convert from **g/dm³ to mol/dm³**.

Example:

Determine the concentration of 0.8mol/dm³ sodium hydroxide **(M_r 40)** solution in g/dm³.

concentration = 0.8 × 40 = 32g/dm³

Volumes of Solutions

By rearranging the concentration equation, it is possible to calculate the amount of a solute in a given volume of solution if the concentration is known.

amount of solute (mol) = concentration (mol/dm^3) × volume (dm^3)

Example:

Determine the amount of 0.2mol/dm^3 sodium hydroxide in 75cm^3 of solution.

Step 1: Convert the volume to dm^3.

75cm^3 = 75.0 ÷ 1000 = 0.075dm^3

Step 2: amount of solute (mol) = concentration (mol/dm^3) × volume (dm^3)

= 0.2mol/dm^3 × 0.075dm^3

= 0.015 mol

Calculating the Mass

Using the example above, calculate the mass of sodium hydroxide (M_r 40) in 75cm^3 of solution.

mass = amount × M_r

mass = 0.015 mol × 40

mass = 0.6g

Percentage Yield – Chemistry Only

The percentage yield can be calculated from the following equation.

$$\text{percentage yield} = \frac{\text{actual mass of product made}}{\text{maxium theoretical mass of product}} \times 100$$

The **theoretical yield** is **the maximum mass** that can be made during a chemical reaction. The law of conservation states that during a chemical reaction, no atoms are lost or made. It's not always possible to obtain the maximum calculated amount of product.

The loss of product may be due to some of the product being lost when filtered. Some of the reactants may not react as expected and so may not produce enough product. The reaction may be a reversible one and as a consequence, the reaction may not go to completion.

Example:

1.8g of copper sulfate crystals are made during a chemical reaction. The theoretical yield for this reaction is **2.0g**. Calculate the percentage yield of copper sulfate.

$$\text{percentage yield} = \frac{1.8\text{g}}{2.0} \times 100$$

percentage yield = 90%

Atom Economy – Chemistry Only

The percentage atom economy can be calculated from the following equation.

$$\text{atom economy} = \frac{\text{relative formula mass of desired product from equation}}{\text{sum of relative formula masses of all reactants from equation}}$$

The **atom economy** is a measure of the amount of starting materials (reactants) that end up as **useful products**. It is important for sustainable development and for economic reasons to use reactions with **high atom economy**. However, not all atoms end up as the desired product and may form other products. We call these **byproducts**.

Example:

When glucose (M_r 180) is fermented, ethanol (M_r 46) is produced.

$$C_6H_{12}O_6 \longrightarrow 2CH_3CH_2OH + 2CO_2$$

Calculate the atom economy for this reaction.

$$\text{atom economy} = \frac{2 \times 46}{180} \times 100$$

atom economy = 51.1%

Reaction Pathways – Higher Tier Only

There is often more than one way to make a substance. Reaction pathways **describe** the **reactions** that have taken place to form the **desired product**. Choosing a particular pathway is dependent upon a number of factors:

1. percentage yield
2. atom economy
3. rate of reaction
4. position of the equilibrium
5. usefulness of any byproducts

The raw materials needed for a particular reaction may affect its chosen pathway. For example, crude oil is a **non-renewable resource**; the resource will run out if we continue to use it. However, plant sugars are **renewable** and can be replenished as long as other plants are replanted.

Ethanol can be made through the fermentation of glucose or the hydration of ethene.

Method of Ethanol Production	Percentage Yield (%)	Atom Economy (%)	Rate of Reaction
fermentation	15	51.1	low
hydration	95	100	high

The hydration of ethene has a 100% atom economy; all atoms react to form the desired product. On the other hand, fermentation has an atom economy of 51.1%. However, its rate of reaction is low in comparison to the hydration method and only has a percentage yield of 15%. Therefore, hydration is the best method for making ethanol.

A byproduct of the fermentation process is carbon dioxide. The gas is sold to fizzy drinks manufacturers to provide the bubbles for some well known fizzy drinks. As the byproduct produced is one that can be useful, it means that the atom economy can be increased to 100%.

Ethene hydration is a reversible reaction. The position of the equilibrium lies to the left. Therefore, only 5% of the ethene supplied to the reaction is actually converted to ethanol. A 95% yield is achieved by recirculating the unreacted ethene.

Avogadro's Law – Higher Tier Only

When the temperature and pressure stay the same, Avogadro's law states that different gases that have the same volume contain equal numbers of molecules.

For example, 1 mol of methane gas occupies the same volume as 1 mol of argon gas.

$H_2 + Cl_2 \longrightarrow 2HCl$

When hyrdogen and chlorine react, hydrogen chloride is produced. In terms of the molar ratio, 10cm^3 of hydrogen reacts completely with 10cm^3 of chlorine. Therefore, the ratio between hydrogen and chlorine is 1:1.

The molar ratio between hydrogen and hydrogen chloride is 1:2. For example, 10cm^3 of hydrogen reacts to produce 20cm^3 of hydrogen.

Molar Gas Volume

The volume of one mole of any gas at room temperature and pressure (20°C and 1 atmosphere pressure) is 24dm^3 (**24 000 cm^3**).

To calculate a known volume of a gas:

volume = amount in mol × molar volume

For example, determine the volume of 0.55 mol of carbon monoxide at room temperature and pressure.

volume = amount in mol × molar volume

volume = 0.55 × 24

= 13.2dm^3

Calculating the Amount of Gas

By **rearranging the equation,** it is possible to calculate the amount of a gas in moles.

For example, determine the amount of hydrogen gas that occupies 198cm^3 at room temperature and pressure.

$$\text{amount in mol} = \frac{\text{volume}}{\text{molar volume}}$$

$$\text{amount in mol} = \frac{198}{24\,000}$$

amount in mol = 0.0083 mol

Calculating a Volume from a Mass

When 3.5g of sodium reacts with water it produces sodium hydroxide and hydrogen gas.

$2Na + 2H_2O \longrightarrow 2NaOH + H_2$

1. Determine the molar amount of sodium (A_r 23).

$$\text{amount in mol} = \frac{\text{mass}}{\text{atomic mass}}$$

$$\text{amount in mol} = \frac{3.5}{23}$$

amount in mol = 0.15 mol

2. Determine the molar amount of hydrogen.

The molar ratio of **sodium to hydrogen**, according to the balanced symbol equation, is **2:1**.

Therefore, 0.15 mol of sodium produces 0.075 mol of hydrogen.

3. Determine the volume of hydrogen.

volume = amount in mol × molar volume

volume = 0.075 × 24dm^3

= 1.8dm^3

4. Calculating the mass from a volume.

Sodium reacts with chlorine to produce sodium chloride.

$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

5. Determine the mass of sodium chloride (Mr 58.5) that can be produced from 685cm^3 of chlorine.

amount of chlorine = 685cm^3 ÷ 24 000 = 0.029 mol

From the equation, the mole ratio between chlorine and sodium chloride is I:2. Therefore, 0.029 moles of chlorine would produce (0.029 × 2) = 0.058 mol.

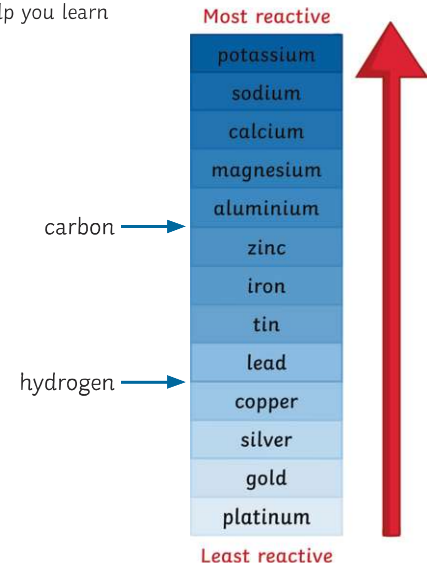
mass of sodium chloride = 0.058 × 58.5 = 3.393g

AQA GCSE Chemistry (Separate Science) Unit 4: Chemical Changes

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)



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 → lithium hydroxide + hydrogen

$2Li + 2H_2O \rightarrow 2LiOH + H_2$

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 → sodium chloride + hydrogen

$2Na + 2HCl \rightarrow 2NaCl + H_2$

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 → salt + hydrogen

For example: hydrochloric acid + sodium → sodium chloride + hydrogen

$\mathbf{2HCl + 2Na \rightarrow 2NaCl + H_2}$

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 acid + alkali → salt + water

hydrochloric acid + sodium hydroxide → sodium chloride + water

$HCl + NaOH \rightarrow NaCl + H_2O$

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:

$2H^+ + Ca \rightarrow Ca^{2+} + H_2$

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

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

Oxidation is loss of electrons.

$2H^+ + 2e^- \rightarrow H_2$

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 → salt + water

sulfuric acid + copper oxide → copper sulfate + water

$H_2SO_4 + CuO \rightarrow CuSO_4 + H_2O$

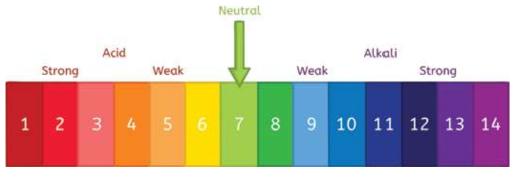
Reactions with Carbonates

The general formula for the reaction between an acid and a carbonate is:

acid + carbonate → salt + water + carbon dioxide

hydrochloric acid + calcium carbonate → 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 nor alkalis.

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

$H^+ + OH^- \rightarrow H_2O$

AQA GCSE Chemistry (Separate Science) Unit 4: Chemical Changes

Making Soluble Salts

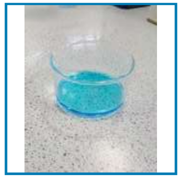
1. Make a saturated solution by stirring copper oxide into the sulfuric acid until no more will dissolve.

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: $HCl \rightarrow H^+ + Cl^-$

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.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

The **double arrow** symbol indicates that the reaction is **reversible**.

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.

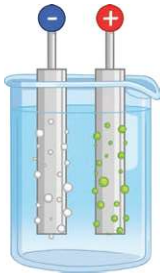
During the process of electrolysis, **opposites attract**. The positively-charged ions will be attracted toward the negative electrode. The negatively-charged ions will be attracted towards the positive electrode.

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

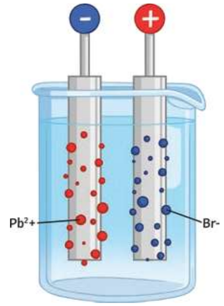
Metals are extracted by electrolysis if the metal in question reacts with carbon or if it is too reactive to be extracted by reduction with carbon. During the extraction process, large quantities of energy are used to melt the compounds.

Aluminium is manufactured by the process of electrolysis. Aluminium oxide has a high melting point and melting it would use large amounts of energy and 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 **i**s the **l**oss of electrons and **r**eduction **i**s the **g**aining 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**.

$Pb^{2+} + 2e^- \rightarrow Pb$

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.

$2Br^- \rightarrow Br_2 + 2e^-$

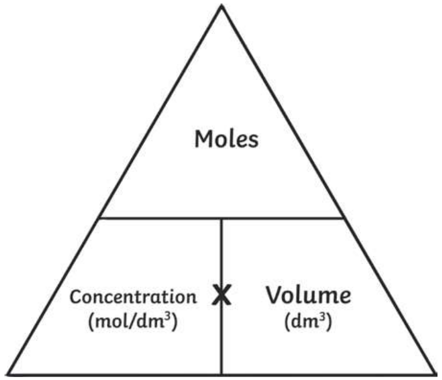
Titration Method (Chemistry Only)

1. Using the pipette and pipette filler, measure $25cm^3$ sodium hydroxide solution and pour into a conical flask.
2. Add several drops of phenolphthalein to the sodium hydroxide solution.
3. Swirl the flask and the mixture should be pink.
4. Place the conical flask on a white tile.
5. Place the burette into its stand, ensuring the tap is closed. Using the funnel, fill the burette with sulfuric acid to the $0cm^3$ line. Should you go above this line, open the tap and allow the excess to run off into a beaker.
6. Once the burette is correctly filled, place over the conical flask.
7. Carefully open the tap so the acid flows slowly into the conical flask. Swirl the flask and look for the indicator changing from pink to colourless.
8. Continue adding the acid to the flask until the indicator is permanently colourless.
9. Record the total volume of acid added to the sodium hydroxide in the results table.
10. Repeat the experiment twice more.

Titration Method (Chemistry Only)

Using the results from a titration experiment, it is possible to calculate the concentration of a solution or the volume of solution required to neutralise an acid or alkali.

Worked Example



In a titration, $20cm^3$ of $1.0mol/dm^3$ sulfuric acid reacted with $25cm^3$ of sodium hydroxide. What was the concentration of sodium hydroxide?

Write out the symbol equation for the reaction.

$NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

Check that the equation is balanced.

$\mathbf{2}NaOH + H_2SO_4 \rightarrow Na_2SO_4 + \mathbf{2}H_2O$

To convert cm^3 to dm^3, just divide by 1000.

Draw a table like the one below and fill it in with the information that you know from the question.

	Acid (H_2SO_4)	Alkali (NaOH)
number of moles		
concentration mol/dm³	1.0	
volume (dm³)	0.02	0.025

As the values for the **concentration** and **volume** of the acid are known, it is possible to now work out the **number of moles** of H_2SO_4.

number of moles = concentration × volume

number of moles = 1.0 × 0.02 = **0.02 moles**

From the balanced symbol equation, we know that there is double the amount of NaOH compared to H_2SO_4, therefore to calculate the number of moles of the alkali, we double the number of moles of the acid. 0.02 × 2 = **0.04 moles**.

	Acid (H_2SO_4)	Alkali (NaOH)
number of moles	0.02	0.04
concentration mol/dm³	1.0	
volume (dm³)	0.02	0.025

The question asks you to calculate the **concentration** of sodium hydroxide. As the number of moles and volume is now known, it is possible to calculate the concentration.

concentration = number of moles ÷ volume

concentration = 0.04 ÷ 0.025

concentration = **1.6mol/dm³**

AQA GCSE Chemistry (Separate Science) Unit 5 Energy Changes Knowledge Organiser

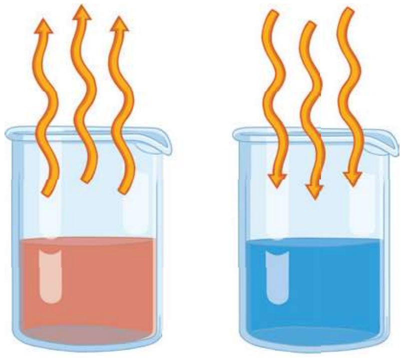
Exothermic and Endothermic Reactions

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 i**ncrease 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.



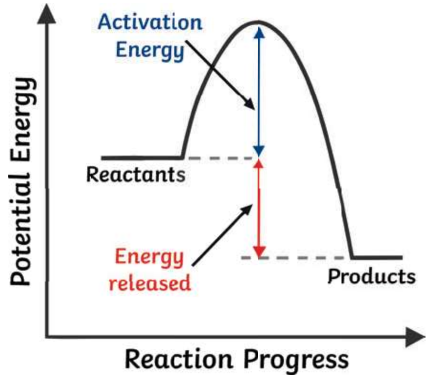
Examples of endothermic reactions include the **thermal decomposition** of calcium carbonate.
Eating **sherbet** is an everyday example of an endothermic reaction. When the sherbet dissolves in the saliva in your mouth, it produces a cooling effect. Another example is **instant ice packs** that are used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

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.

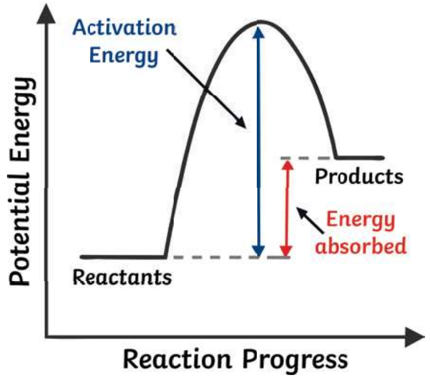
In an **exothermic** reaction, the difference in energy is **released** to the surroundings and so the **temperature** of the surroundings **increases**.



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**.



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.**

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: $2H_2O_2 \longrightarrow 2H_2O + O_2$

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.

2H–O–O–H ⟶ 2H–O–H + O=O

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.

So 464 + 146 + 464 = 1074

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

So 1074 × 2 = 2148

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

There are two **H–O** bonds

So 464 + 464 = 928

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

So 928 × 2 = 1856

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

So 1856 + 498 = 2354

ΔH = sum (bonds broken) – sum (bonds made)

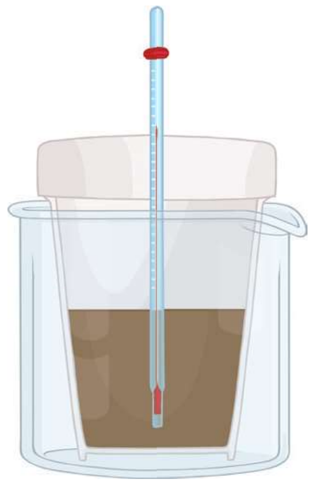
ΔH = 2148 – 2354 = -206 kJ/mol

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
- 250cm^3 glass beaker
- measuring cylinder
- top pan balance

Method

Reaction between a metal and an acid.

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 30cm^3 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.

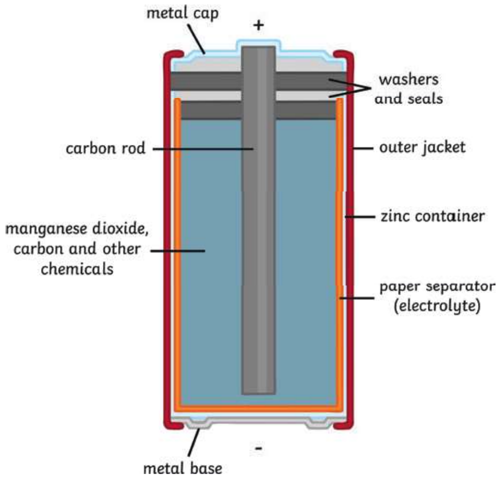
Chemical Cells

A chemical cell converts **chemical energy** into **electrical energy.** More than one cell connected in series is called a battery.

There are two types of chemical cell, **rechargeable** and **non-rechargeable.**

Non-rechargeable cells will produce a **voltage** until the chemicals inside are used up. Once this occurs, the cell is no longer useful and can then be recycled.

Rechargeable cells and batteries can be recharged multiple times. An electrical current is passed backwards through the cell. This works by **reversing** the chemical reactions and the cell or battery can then be used again to produce more electricity. Mobile phones contain rechargeable batteries.

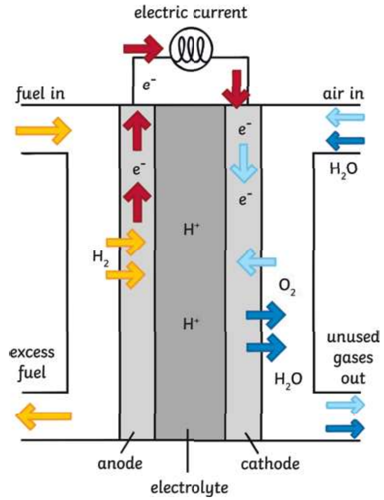


Fuel Cells

Fuels cells work differently to chemical cells in that they need to be supplied with a fuel and oxygen. The constant supply of these two ingredients will allow a fuel cell to produce a voltage continuously.

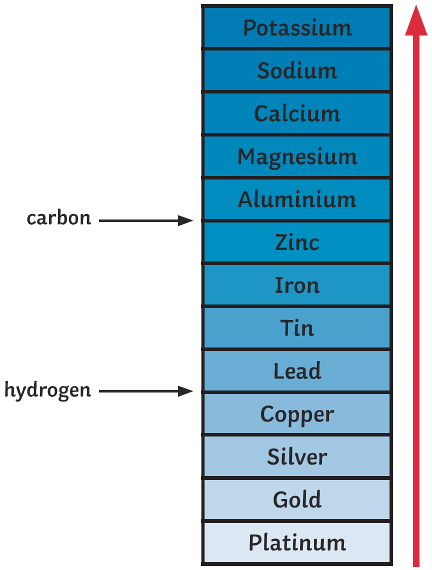
Inside the fuel cell, hydrogen is **oxidised** electrochemically; the fuel is **not combusted.** This allows the reaction to take place at a lower temperature.

Hydrogen-oxygen fuel cells are an alternative to rechargeable batteries and cells as the only product that is produced is water.



Voltage

The voltage of a cell is affected by the combination of metals used inside it. The bigger the difference in the **reactivity** of the two metals, the bigger the **voltage** produced. For example, if the metals used inside the cell are magnesium and zinc, then the voltage produced will be **small** as the two metals are **close together** in the **reactivity series.** By comparison, if magnesium and copper are used, then the voltage produced will be **larger** as the metals are **further apart** in the **reactivity series.**



Ionic Equations

hydrogen + oxygen ⟶ water

$2H_2 + O_2 \longrightarrow 2H_2O$

At the **cathode:** $2H_2 + 4OH^- \longrightarrow 4H_2O + 4e^-$

At the **anode:** $O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$

In the fuel cell, **oxygen** is being **reduced** (reduction is the gaining of electrons) whilst **hydrogen** is being **oxidised** (oxidation is the loss of electrons). Oxidation and reduction happen simultaneously – this is called a **redox reaction.**

Calculating Rates of Reactions

Reactions happen at **varying rates**. For example, a firework exploding is a fast reaction whereas a piece of iron rusting would take place over a longer period of time.

The **rate of a chemical reaction** tells us how quickly a **product is formed** or how quickly a **reactant is used up**.

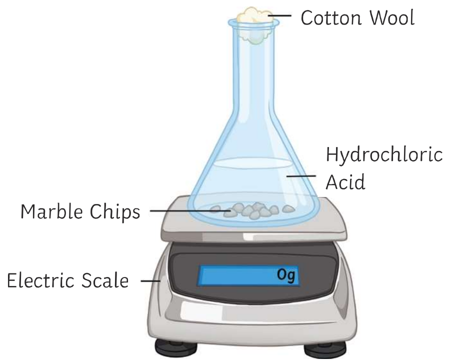
For a chemical reaction to occur, the reactant particles must collide with enough energy. Those collisions that produce a chemical reaction are called successful collisions.

$$\text{mean rate of reaction} = \frac{\text{quantity of reactant used}}{\text{time taken}}$$

$$\text{mean rate of reaction} = \frac{\text{quantity of product formed}}{\text{time taken}}$$

Measuring the Mass of a Reaction Mixture

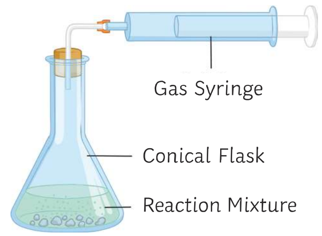
The changing mass of a reaction mixture can be measured during a reaction. This method is particularly useful when gases, such as carbon dioxide, are given off. **Gas escapes during the reaction and the mass of the reaction mixture decreases.** The mass can be measured at regular time intervals.

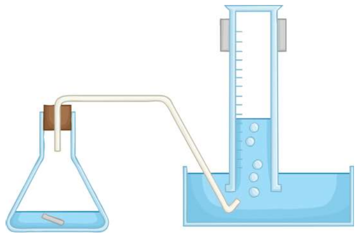


units = g/s or g/min

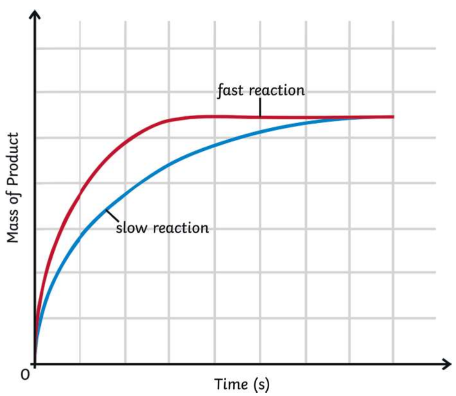
Measuring the Volume of a Reaction Mixture

The changing volume of a reaction mixture can be measured during a reaction. This method is particularly useful when gases, such as carbon dioxide, are given off. The gas can be collected and its volume measured at regular time intervals. Different types of measuring equipment can be used to collect the gas such as a gas syringe, measuring cylinder or upside-down burette.





units = cm^3/s or cm^3/min



Graphs are a useful way to **analyse** the results from a rate of reaction investigation. The graph above shows two lines, one red and one blue.

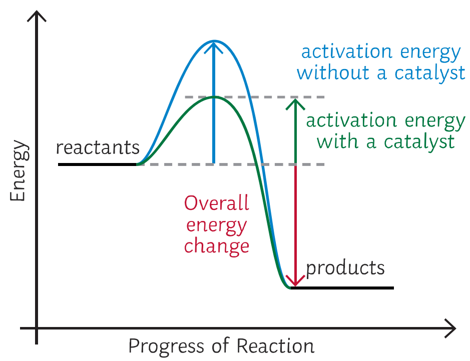
The red line represents a fast reaction and the blue line a slow reaction. We know the fast reaction occurs at a much faster rate as the line is steep. The fast reaction finishes before the slow reaction as the line plateaus sooner.

Factors Affecting the Rate of a Chemical Reaction

- concentration and pressure
- catalyst
- surface area
- temperature

The rate of a chemical reaction will be increased if there are more frequent successful collisions between reactant particles.

Catalyst



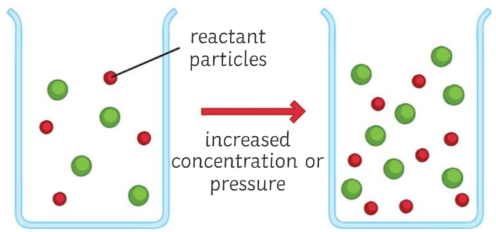
A catalyst is a **substance** that speeds up a chemical reaction without getting used up itself. Catalysts are able to offer an **alternative pathway** at a **lower activation energy**.

Biological catalysts are called **enzymes**.

When a catalyst is used in a chemical reaction (not all reactions have a catalyst that is suitable to use), the **frequency of collisions** is **unchanged**. More **particles** are able to react. The particles have **energy greater** than that of the **activation energy**. Consequently, there is in an **increase** in the **rate successful of collisions**.

Concentration and Pressure

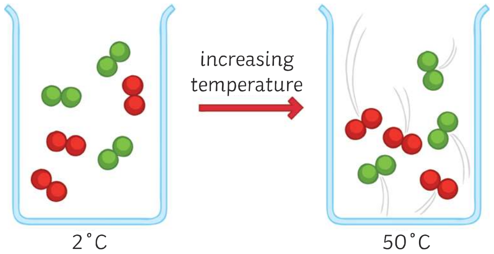
If the **number of reactant particles** in a given space is **doubled**, there will be **more frequent successful collisions** between reactant particles, therefore, **increasing** the **rate of reaction**.



AQA GCSE Chemistry (Combined Science) Unit 6: The Rate and Extent of Chemical Change

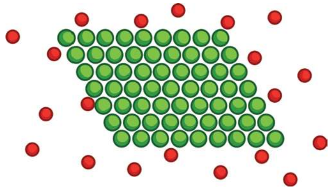
Temperature

When the temperature of the reaction mixture is increased, the reactant particles **gain kinetic energy** and move much more quickly. This results in **more frequent successful collisions** between the reactant particles, therefore, **increasing the rate of the reaction**.



Surface Area

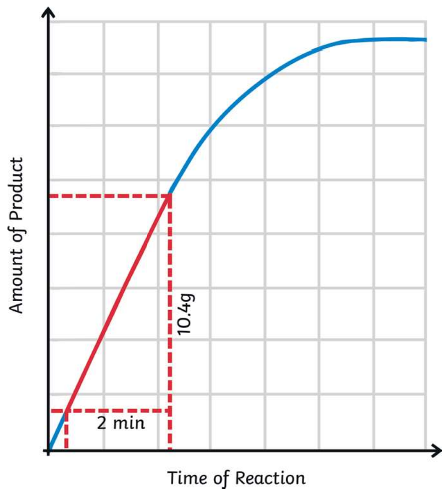
Large lumps of a solid have a **small surface area to volume ratio**. If the solid is broken up into smaller lumps or crushed into a powder, this will increase the surface area to volume ratio.



A larger area of the solid is now exposed to other reactant particles. This increases the frequency of successful collisions thus increasing the rate of reaction.

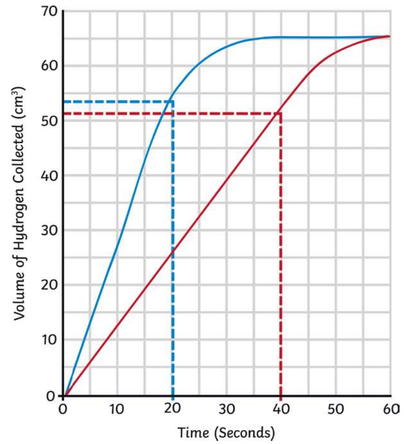
Calculating Gradient (Higher Tier Only) gradient = $\frac{y}{x}$

On the graph, draw construction lines on the part of the graph that has a straight line. Measure the values of x and y.



In the graph below, the gradient of the first line is much steeper than the second line. This indicates that a faster reaction is taking place. Remember, the steeper the line, the faster the reaction.

To calculate the reaction rate at a specific time period, construction lines must first be drawn on the straightest part of the graph.



For the first line, what is the rate of reaction at 20 seconds?

54 ÷ 20 = 2.7cm^3/s

For the second line, what is the rate of reaction at 40 seconds?

52 ÷ 40 = 1.3cm^3/s

Dynamic Equilibrium

In a **closed system** (this means nothing can get in or out), a reversible reaction can reach **dynamic equilibrium**. This is where the **forward** and **reverse reactions** are occurring at the **same rate** and the **concentrations** of all the substances that are reacting remain constant.

Changing Conditions and the Effect on the Position of Equilibrium (Higher Tier Only)

The reaction between nitrogen and hydrogen to make ammonia is an industrial process called the Haber process. It requires a high temperature, high pressure and an iron catalyst.

The symbol equation for the reaction is as follows:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

According to **Le Chatelier's Principle**, the position of equilibrium can be altered by changing the conditions of the reaction i.e. the pressure, concentration and/or the temperature. The **position** of the **equilibrium** will shift to **counteract** any changes made.

Increasing the **temperature** of the reaction in the forward direction (exothermic) will result in the equilibrium shifting in favour of the reverse direction (endothermic) to reduce the temperature.

From the equation, it is clear that on the **left-hand side,** there are **four molecules** and on the **right-hand side,** there are **two molecules**. If the **pressure** in the system were **increased**, the equilibrium **position would shift to the right** as there are fewer molecules. If the pressure in the system were **decreased**, the equilibrium **position would shift to the left** as there are a larger number of molecules.

If the **concentration** of one of the **reactants were increased**, then the equilibrium position would move in **favour of the products**. This would result in more product being produced. If the concentration of the **products were decreased**, equilibrium would shift to **favour the products**. More reactants would react until equilibrium is reached.

Reversible Reactions

A reversible reaction is one in which the **reactants form products**. The products are then able to react together to **reform the reactants**.

For example:
A reacts with B to form C and D.
C and D are able to react to form A and B.

The equation would be as follows (where the **double arrow symbol** represents a **reversible reaction** is taking place):

$A + B \rightleftharpoons C + D$

The **forward reaction** goes to the **left** and the **backwards reaction** goes to the **right**. For example, if the forward reaction is exothermic then the backward reaction will be endothermic. The amount of energy that is transferred is the same for both the forward and reverse reaction.

Hydrated copper sulfate is a blue substance. We say that the copper sulfate is hydrated as it **contains water**. The copper sulfate is heated and the water evaporates leaving a white substance known as **anhydrous** copper sulfate. Anhydrous meaning **no water**.

The word equation for the reaction is as follows:

hydrated copper sulfate $\rightleftharpoons$ anhydrous copper sulfate + water

$CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4(s) + H_2O(l)$

The reaction can be reversed when water is added to the anhydrous copper sulfate.

Required Practical 5: Measuring the Production of a Gas

This method outlines one way to carry out an investigation to collect a gas from a chemical reaction.

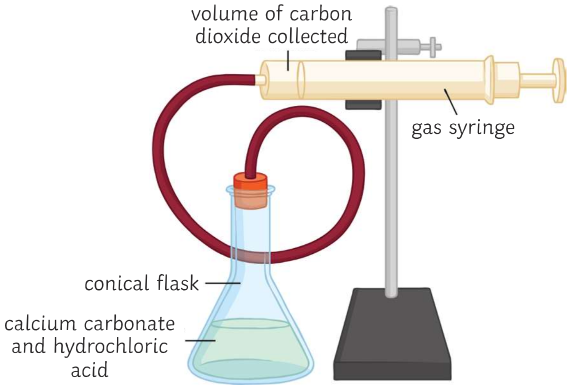
The practical involves changing the concentration of hydrochloric acid and measuring the volume of carbon dioxide gas produced when the acid reacts with calcium carbonate.

The word equation for the reaction is as follows:

calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

The symbol equation for the reaction is:

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



Method

Step 1 – Clamp a gas syringe to a retort stand using a boss and clamp. Ensure the syringe is a quarter of the way from the top of the stand. Place the delivery tube to the end of the gas syringe.

Step 2 – Measure out 50ml of hydrochloric acid using a measuring cylinder and pour into a conical flask.

Step 3 – Using a top pan balance, measure out 0.5g of powdered calcium carbonate and place in the conical flask.

Step 4 – Immediately connect the bung and delivery tube to the conical flask. Start the stopwatch.

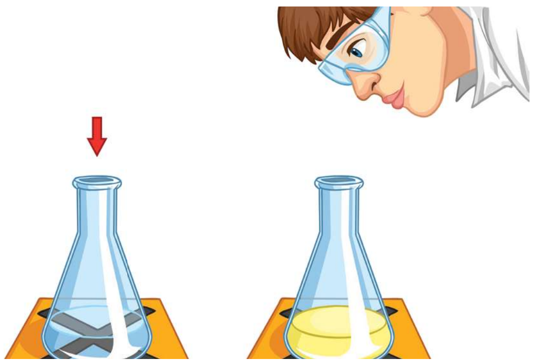
Step 5 – Record the volume of carbon dioxide gas produced every 10 seconds.

Step 6 – When the reaction has finished and there are no more bubbles of gas being produced, clean the equipment and repeat using four other different concentrations of hydrochloric acid.

When analysing the results from the practical investigation, plot a graph of Time (s) against Volume of Gas Produced (cm^3). Draw a curve of best fit through the points. A graph should be plotted for each concentration of acid.

Calculate the mean rate of reaction (cm^3/s) for each concentration of acid used. This can be calculated by dividing the total mass of gas produced (cm^3) by the reaction time (s).

Required Practical 5: Investigating a Change in Colour



This method outlines one way to carry out an investigation into the effect of increased temperature on the rate of a reaction.

The word equation for this reaction is as follows:

sodium thiosulfate + hydrochloric acid → sodium chloride + water + sulfur dioxide + sulfur

The symbol equation for this reaction is:

$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + H_2O + SO_2 + S$

The reaction between sodium thiosulfate and hydrochloric acid produces a **precipitate**. **Sulfur** is responsible for the formation of the precipitate. A precipitate is a **solid** that is formed in a solution. It is the formation of this precipitate that causes the reaction mixture to become **cloudy**; the cloudiness is a way to measure the **reaction time**.

Method

Sodium thiosulfate from three different temperatures may be used, for example, ice cold, room temperature and hot.

Step 1 – Place a black cross on a white tile.

Step 2 – Using the first temperature, measure out $35cm^3$ of sodium thiosulfate using a measuring cylinder. Place the liquid in a conical flask and position over the black cross on the white tile.

Step 3 – Measure out $5cm^3$ of water and $10cm^3$ of hydrochloric acid in separate measuring cylinders.

Step 4 – Pour the water and acid into the conical flask.

Step 5 – Pour the measured amount of sodium thiosulfate into the conical flask and immediately start the stopwatch.

Step 6 – Look down through the conical flask to the black cross below. When the black cross is no longer visible, stop the stopwatch and record the results in a table.

Step 7 – Repeat the steps with the remaining temperatures of sodium thiosulfate.