



Year 11

**F Yj]g]cb`A UhYf]U`g`Zcf
.....Chemistry**

Pages from the Year 10 revision guide -
needed for Year 11 mock exam.

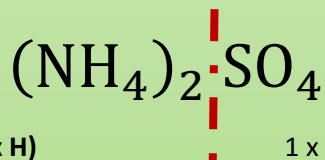
Using chemical formulae

Compounds are written using a **formula** of the **symbols** of the **elements** in it.

Compound	Formula	No. of elements	No. of atoms
Sodium Chloride	NaCl	2	2 (1 Na, 1 Cl)
Sodium Hydroxide	NaOH	3	3 (1 Na, 1 O, 1 H)
Sodium Oxide	Na ₂ O	2	3 (2 Na, 1 O)
Sodium Sulfate	Na ₂ SO ₄	3	7 (2 Na, 1 S, 4 O)
Calcium Carbonate	CaCO ₃	3	5 (1 Ca, 1 C, 3 O)

Worked Example Question

What is the total number of atoms in this formula (NH₄)₂SO₄, ammonium sulfate?



$$2 \times \text{NH}_4 = (2 \times \text{N}) + (2 \times \text{H}_4) = (2 \times \text{N}) + (8 \times \text{H})$$

$$1 \times \text{SO}_4 = (1 \times \text{S}) + (1 \times \text{O}_4) = (1 \times \text{S}) + (4 \times \text{O})$$

$$\text{So... } (2 \times \text{N}) + (8 \times \text{H}) + (1 \times \text{S}) + (4 \times \text{O})$$

$$\text{TOTAL} = 15 \text{ Atoms}$$

Chemical Reactions

Atoms are **rearranged** during a chemical reaction. None are created or destroyed.

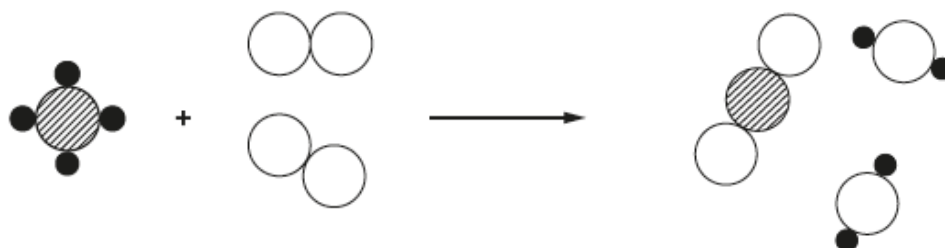
Representing simple molecules using a diagram and key

Sometimes it's hard to **visualise** what's happening during a reaction, so it's useful to represent simple molecules using a diagram:

Methane burns in air forming carbon dioxide and water. This reaction is represented by the symbol equation:



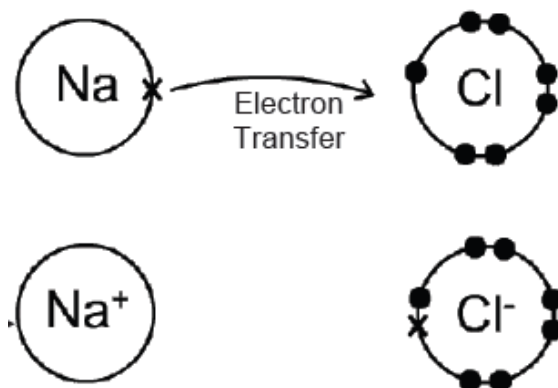
It can also be represented by the diagram equation:



Ionic compounds

When a chemical reaction takes place, **new bonds are formed**.

Ionic compounds form by the **transfer of electrons** from **metal** to **non-metal** atom. Charged particles called **ions** are formed.



e.g. When sodium chloride (NaCl) forms, one electron is transferred from the outer orbit (*shell*) of the **sodium atom** into the outer orbit of the **chlorine atom**.

This will form a **full stable outer orbit** (*shell*) for the **two** particles.

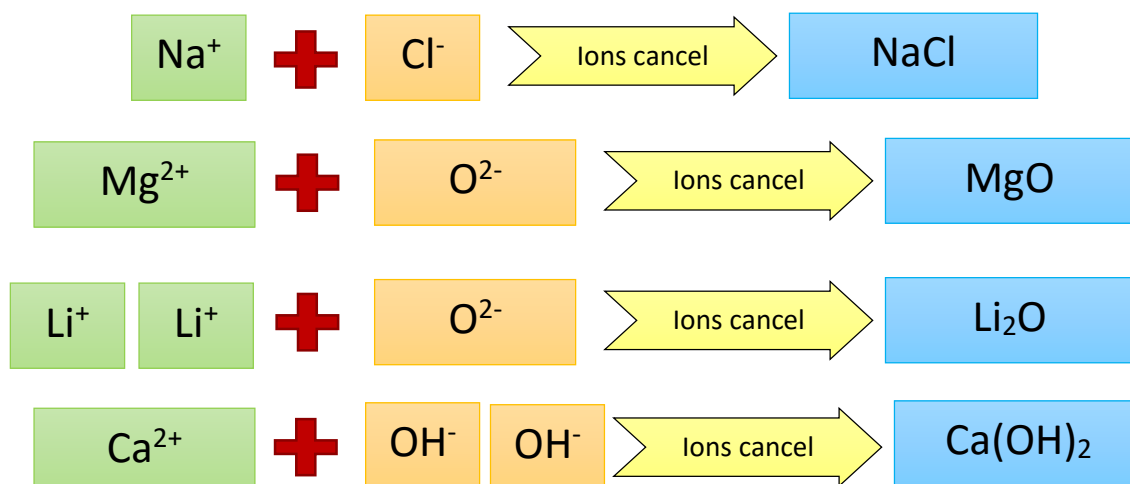
The **sodium atom** becomes a **positive sodium ion (Na⁺)**.

The **chlorine atom** becomes a **negative chloride ion (Cl⁻)**.

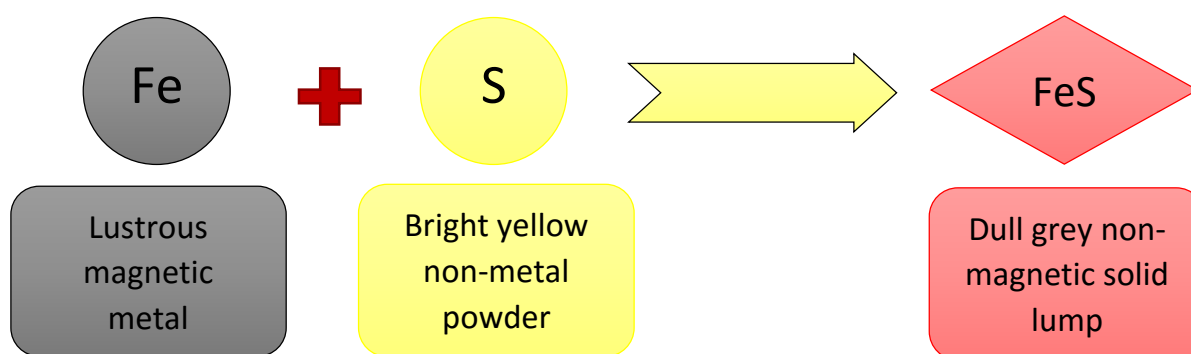
It's the strong **electrostatic attraction** between the **two charges** that holds ionic compounds together.

Overall there is no charge as the **positive** and **negative** charges cancel.

Using ions to create formulae



Compounds have completely different properties to their constituent elements



Percentage (%) composition of compounds

After calculating the M_r it's possible to work out the percentage of an element in a compound using the formula:

$$\% \text{ Percentage mass} = \frac{A_r \times \text{No. of atoms (of that element)}}{M_r \text{ (of the whole compound)}} \times 100$$

The best way to demonstrate this is by example:

Calculate the percentage mass of magnesium in magnesium carbonate, MgCO_3 .

A_r of Mg = 24

A_r of C = 12

A_r of O = 16

M_r of MgCO_3 =

$24 + 12 + (3 \times 16)$

= 84

Now:

$$\% \text{ mass} = \frac{A_r \times N}{M_r} \times 100 = \frac{24 \times 1}{84} \times 100 = \underline{\underline{28.6\%}}$$

This tells us that 28.6% (by mass) of MgCO_3 is magnesium.

Chemical Equations

Atoms **aren't made or lost** during a **chemical reaction** – just re-arranged.

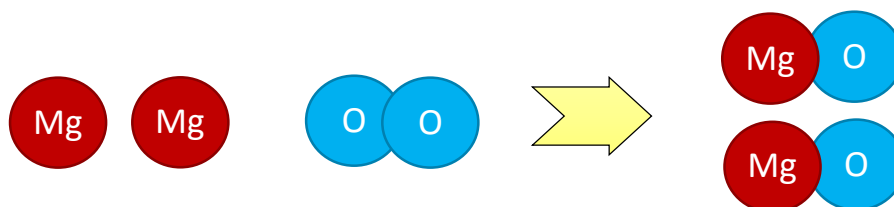
The reactions can be shown using **word**, **symbol** or **diagram** equations.

e.g.

Word: Magnesium + Oxygen → Magnesium Oxide

Balanced Symbol: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

Diagram:



There must always be the **same number of atoms** on **both sides** of the equation – so that they are equal.

Left-hand side (LHS):

Mg = 2

O = 2

Right-hand side (RHS):

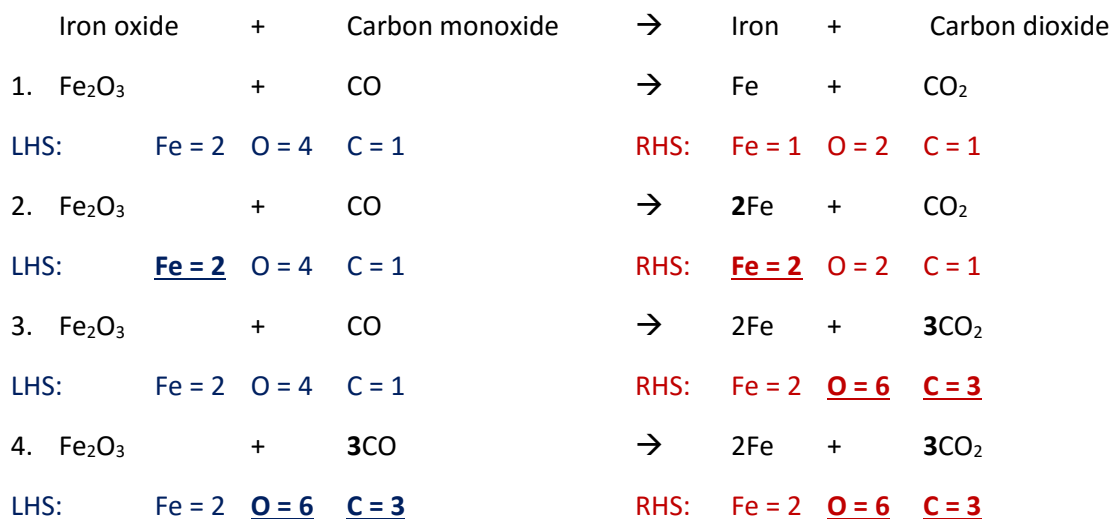
Mg = 2

O = 2

Balancing equations

We balance equations by putting numbers **in front** of the formulas when needed.

e.g.

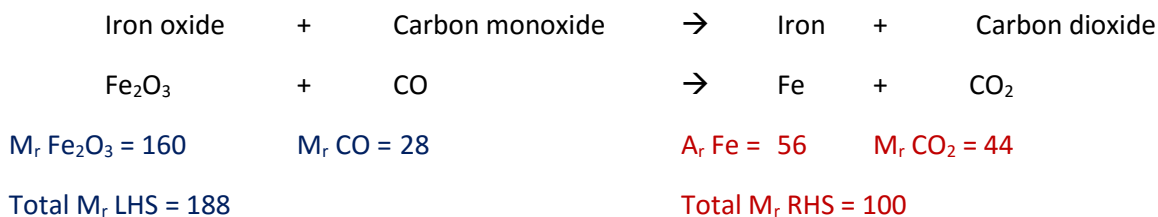
Blast furnace reaction:

The number of atoms now balance!

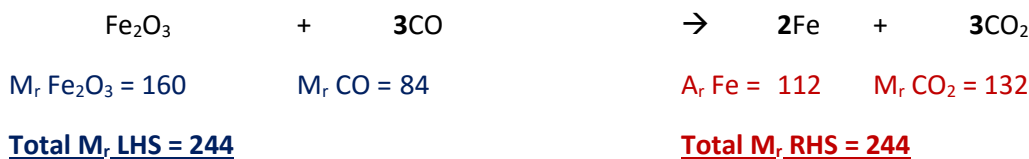
Balanced mass

We can see above that the **number of atoms** that **go into a reaction** is equal to the **number of atoms** that **come out** of a reaction.

This is true for **the mass** also.

Blast furnace reaction:

We can see here that the mass going in does not equal the mass coming out.



The mass now balances!

Moles

Higher Tier

The mole is a term that **describes a specific number** – like the word ‘dozen’ represents the number 12. The mole however is a much larger number 6.02×10^{23} atoms. (6 followed by 23 zeros).

This number is also called **Avogadro constant** or **Avogadro’s number**.

A mole is defined as the number of atoms in exactly 12 grams of Carbon-12 (^{12}C). 12 is the mass number of carbon, so one mole of carbon atoms has a mass of 12 grams.

So... the A_r or M_r of a substance **in grams** is known as **one mole of that substance**.

e.g.

Iron has an A_r of 56.

So, one mole of iron has a mass of 56g.

Nitrogen gas has an M_r of 28 (2×14).

So, one mole of nitrogen gas has a mass of 28g.

Calcium carbonate (CaCO_3) has an M_r of 100.
($40 + 12 + (16 \times 3)$) = 100

So, one mole of CaCO_3 has a mass of 100g.

To calculate the number of moles, we use this equation:

$$\text{Number of moles} = \frac{\text{Mass in g (of element or compound)}}{M_r(\text{of element or compound})}$$

Example 1:

How many moles of atoms are there in 4.8 g of carbon?

$$\text{moles} = \frac{\text{mass}}{A_r} = \frac{4.8\text{g}}{12} = 0.4 \text{ moles}$$

$A_r \text{ C} = 12$

Example 2:

How many moles are there in 640 g of oxygen (O_2)?

$$\text{moles} = \frac{\text{mass}}{M_r} = \frac{640\text{g}}{32} = 20 \text{ moles}$$

$A_r \text{ O} = 16$

$M_r \text{ O}_2 = 16 \times 2 = 32$

Converting moles into mass

You can **rearrange the equation** to form: $\text{mass} = \text{moles} \times M_r$

Example 3:

What is the mass of 0.6 moles of chlorine molecules (Cl_2)?

$$\text{mass} = \text{moles} \times M_r = 0.6 \times 71 = 42.6\text{g}$$

$A_r \text{ Cl} = 35.5$ $M_r \text{ Cl}_2 = 35.5 \times 2 = 71$

Example 4:

What is the mass of 0.1 moles of calcium carbonate (CaCO_3)?

$$\text{mass} = \text{moles} \times M_r = 0.1 \times 100 = 10.0\text{g}$$

Calculating the M_r from moles and mass

You can **rearrange the equation** to form:

$$M_r = \frac{\text{mass}}{\text{moles}}$$

Example 5:

0.5 moles of a compound weighs 80g, calculate its M_r .

$$M_r = \frac{\text{mass}}{\text{moles}} = \frac{80}{0.5} = 160$$

Calculations

Calculating the percentage yield (%) of a chemical reaction

The amount of product we get from a chemical reaction is called the **yield**. The more reactants we put in, the higher the **actual yield** will be.

The **percentage yield (%)** tells us the overall success of the experiment. It compares the **predicted yield** (what we should get) with the **actual yield** (what we actually get in practice).

$$\text{Percentage yield (\%)} = \frac{\text{actual yield (in g)}}{\text{predicted yield (in g)}} \times 100$$

e.g. In a manufacturing process 12 tonnes of product are predicted but only 10 tonnes are obtained. What is the percentage yield?

$$\text{Percentage yield (\%)} = \frac{10}{12} \times 100 = 83.33 \%$$

Calculating the masses of reactants or products from a balanced chemical equation

Higher Tier

By using relative atomic masses (A_r) and relative molecular masses (M_r) it is possible to **calculate how much of a product is produced** or how much reactants are needed.

Product calculation

What **MASS OF MAGNESIUM OXIDE** is produced when 60g of magnesium is burned in air?

Symbol Equation:	$2Mg + O_2 \rightarrow 2MgO$	
M_r:	2x24	2 (24+16)
	48	80
Therefore	48g (or tonnes) will produce 80g	
So for every 1g...	1g	$80 \div 48 = 1.67g$
So...	60g will produce $60 \times 1.67 = \underline{\underline{100.2g}}$	

Product
Reactant

Reactant calculation (*working backwards*)

What is the **MASS OF MAGNESIUM** needed to produce 90g of magnesium oxide?

Symbol Equation:	$2Mg + O_2 \rightarrow 2MgO$	
M_r:	2x24	2 (24+16)
	48	80
Therefore	48g (or tonnes) will produce 80g	
So for every 1g...	1g	$48 \div 80 = 0.6g$
So...	90g will produce $90 \times 0.6 = \underline{\underline{54g}}$	

Reactant
Product

Calculating the formula of a compound from reacting mass data (*Empirical Formula*)**Higher Tier****Example 1**

When **4 g of copper oxide** is reduced in a **steam of hydrogen**, **3.2 g of copper** remains.

Work out how much oxygen was contained in the copper oxide.

1. First step

Find the mass difference

$$4 - 3.2 = \mathbf{0.8\text{ g}}$$

So for every copper oxide:

	<u>Cu</u>	<u>O</u>
Mass	3.2	0.8
2. Divide with A_r	÷ 64 (A _r Cu)	÷ 16 (A _r O)
	= 0.05	= 0.05
3. Divide with smallest	÷ 0.05	÷ 0.05
	= 1	= 1

So ratio is

1 Cu : 1 O

For every 1 Cu there is 1 O

→

Formula = CuO

Example 2

Find the formula of iron oxide produced when 44.8g of iron reacts with 19.2g of oxygen.

So, for every iron oxide:

	<u>Fe</u>	<u>O</u>
Mass	44.8	19.2
Divide with A _r	÷ 56 (A _r Fe)	÷ 16 (A _r O)
	= 0.8	= 1.2
Divide with the smallest value	0.8 ÷ 0.8	1.2 ÷ 0.8
	= 1	= 1.5

A formula must have **whole numbers** therefore the ratio is:

2 Fe : 3 O

For every 2 Fe there are 3 O

→

Formula = Fe₂O₃

Pages from the Year 11 revision guide -
needed for Year 11 mock exam.

Topic 1 - BONDING, STRUCTURE AND PROPERTIES

Metals and Non-metals

One of the easiest ways to put elements into groups is to split them into **metals** and **non-metals**.

[illegible]

Metals appear on the left side of the periodic table, whereas **non-metals** appear on the right. Many elements in Groups 3, 4 & 5 show **metallic** and **non-metallic** properties.

Physical properties of metals and non-metals

How can we tell if a substance is a metal or not? One of the first things we look at are the **properties** of the metal, i.e. a distinctive attribute or quality of a substance.

Metals

- Conducts electricity
- Conducts heat
- High melting point
- High boiling point
- Malleable – Can be hammered into sheets
- Ductile – Can be drawn out into wire or threads
- Lustrous / Shiny

Non-metals

- **DOES NOT** conduct electricity
- **DOES NOT** Conduct heat
- Low melting point
- Low boiling point
- Non-malleable
- Brittle – breaks or fractures easily
- Dull

Ionic and covalent bonding

We have already seen how bonding happens in a metallic element, but what about in non-metals and within compounds?

When a **chemical reaction** takes place, new bonds are formed.

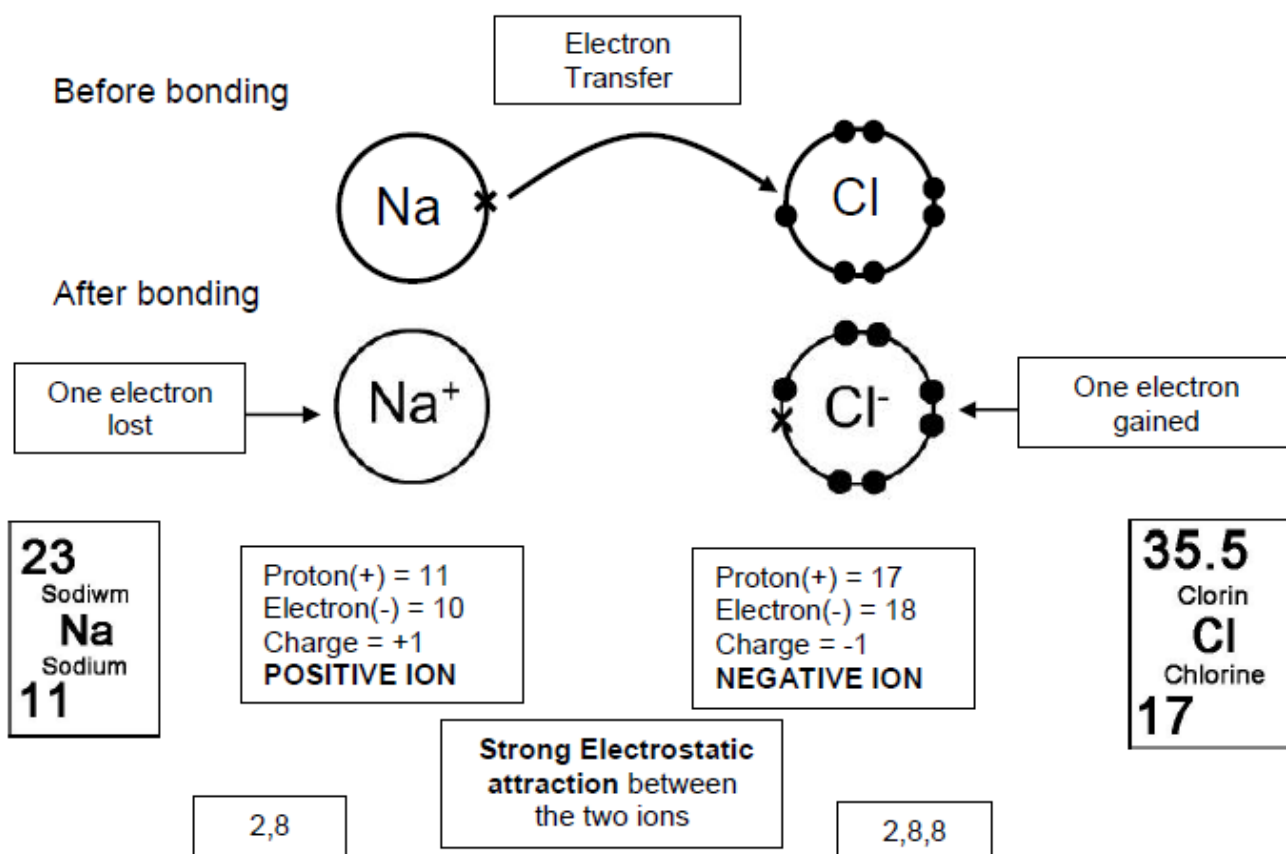
Ionic compounds form by the **transfer of electrons** from a **metal** to a **non-metal** atom. Charged particles called **ions** are formed. (see Chemistry 1)

Covalent bonding occurs between 2 or more **non-metals**. When these bonds form, the atoms **share** electrons.

Ionic bonding

Charged particles called ions are formed when electrons are transferred between atoms during chemical bonding.

e.g. When sodium chloride (NaCl) forms, one electron is transferred to chlorine. This will form a full stable outer shell (like noble gasses) for the two particles.

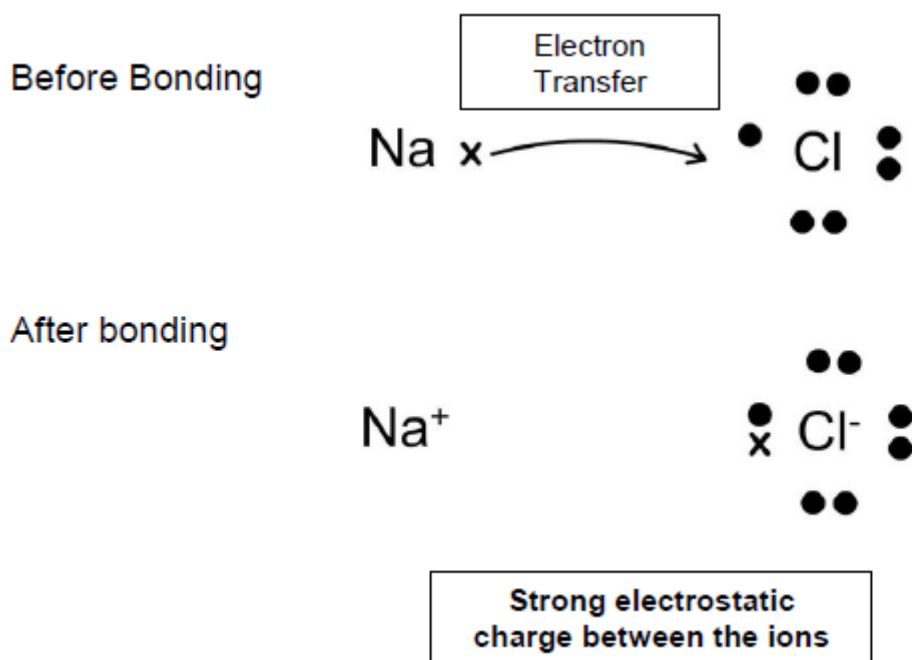


The above diagram shows the **dot and cross representation** which is commonly accepted by chemists and scientists across the world.

Examples of ionic bonding:

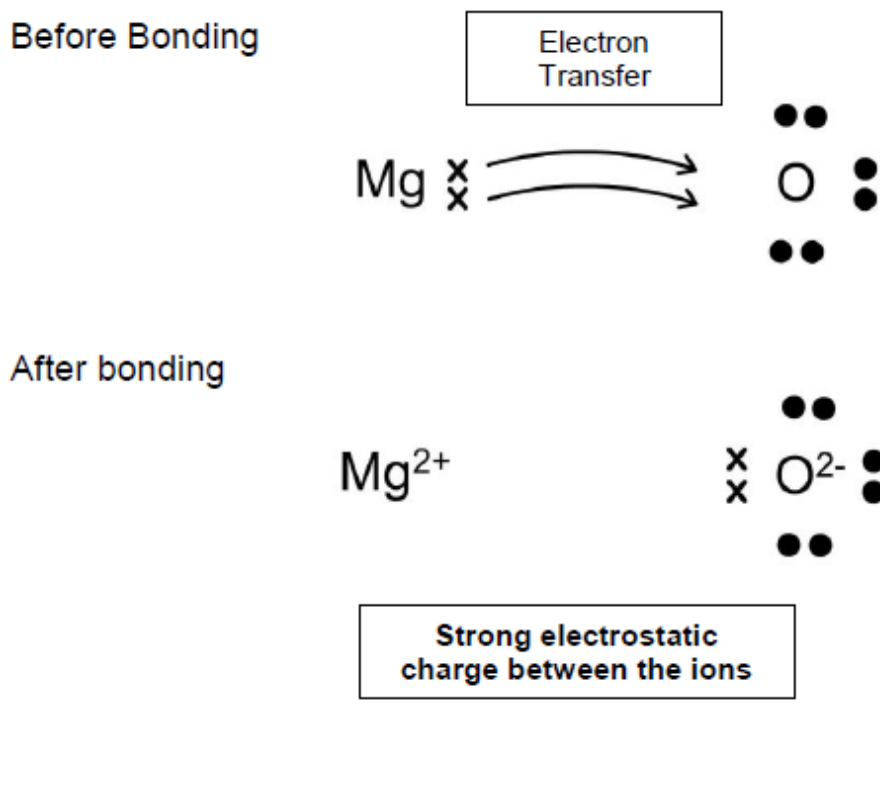
1. Sodium chloride (NaCl)

Group 1 metal with a group 7 non-metal (halide)



2. Magnesium oxide (MgO)

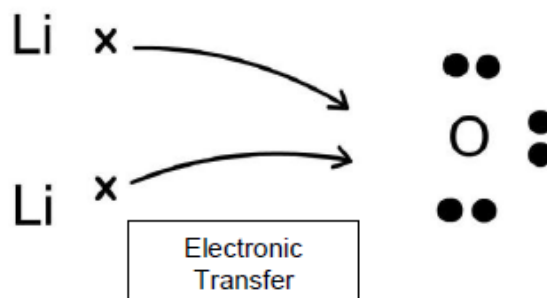
Group 2 metal with a group 6 non-metal



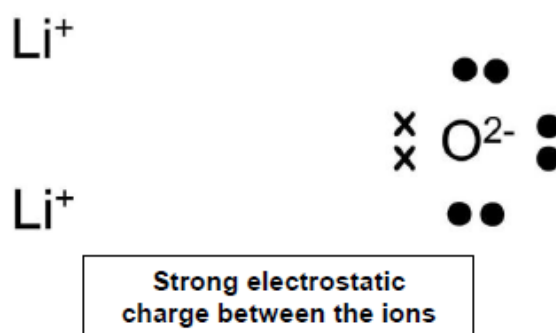
3. Lithium Oxide (Li_2O)

Group 1 metal with a group 6 non-metal

Before Bonding



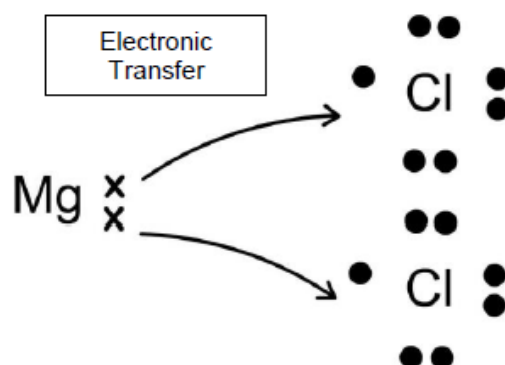
After bonding



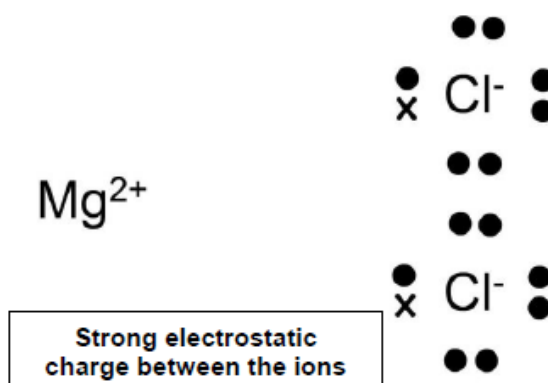
4. Magnesium Chloride (MgCl_2)

Group 2 metal with a group 7 non-metal (halide)

Before Bonding



After bonding



Simple and Giant structures

1. Giant ionic structures

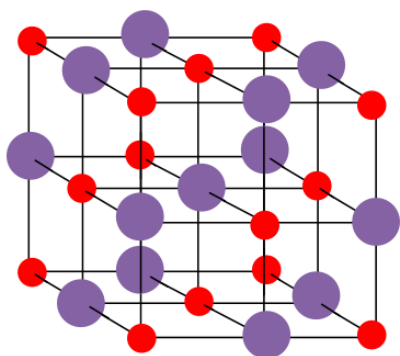
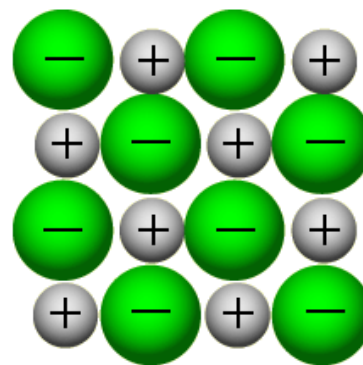
There are many ionic bonds in an ionic compound such as sodium chloride, arranged in ***giant lattice structures***. Ionic compounds have **high melting and boiling points** because of the strength of the **electrostatic forces** of attraction between oppositely-charged ions.

The oppositely-charged ions are arranged in a regular way to form a giant ionic lattice. It is a 'lattice' because the arrangement is a regular one and 'giant' because the arrangement is repeated many times with large numbers of ions.

Ionic compounds often form **crystals** as a result.

The number of ions in an ionic compound is such that the **overall charge** of a sample of the compound is **zero**.

For an ionic substance to **conduct electricity**, its ions must be free to move so that they can carry charge from place to place. Ions are free to move when an ionic compound is a **molten liquid** or **in solution** (i.e. dissolved in water or another solvent) but **not when it is solid**.



Higher tier

The melting point of sodium chloride is lower than that of magnesium oxide because it has weaker ionic bonds, which need less heat energy to break/overcome.

This is for two reasons:

1. The Na^+ and Cl^- ions in sodium chloride have fewer charges than the Mg^{2+} and O^{2-} ions in magnesium oxide.
2. Na^+ ions are larger than Mg^{2+} and cannot get as close to the negatively charged ions.

Topic 2 - ACIDS, BASES AND SALTS

Acids and alkalis are some of the most widely used chemicals both in industry and our homes.

Common acids and alkalis

Most people think **acids** are dangerous, and **some** are. Concentrated **sulfuric acid** is **very corrosive**, and must be handled with care. But other acids are common chemicals that we use every day. **Acetic acid** - also called **ethanoic acid** - is present in vinegar. **Citric acid** is found in fruits such as oranges and lemons. These acids are safe to eat.

It is a similar story with alkalis. A concentrated solution of the alkali **sodium hydroxide** is **very caustic**. It can be used to remove skin from bones to obtain skeletons. But **magnesium hydroxide** is such a mild alkali it can be taken as an indigestion remedy.

Acids and alkalis that are dangerous are stored in containers which have hazard warning signs to show they are **corrosive**. These chemicals need to be handled with care, with goggles and gloves worn.



Acid, Alkali or Neutral - The pH scale

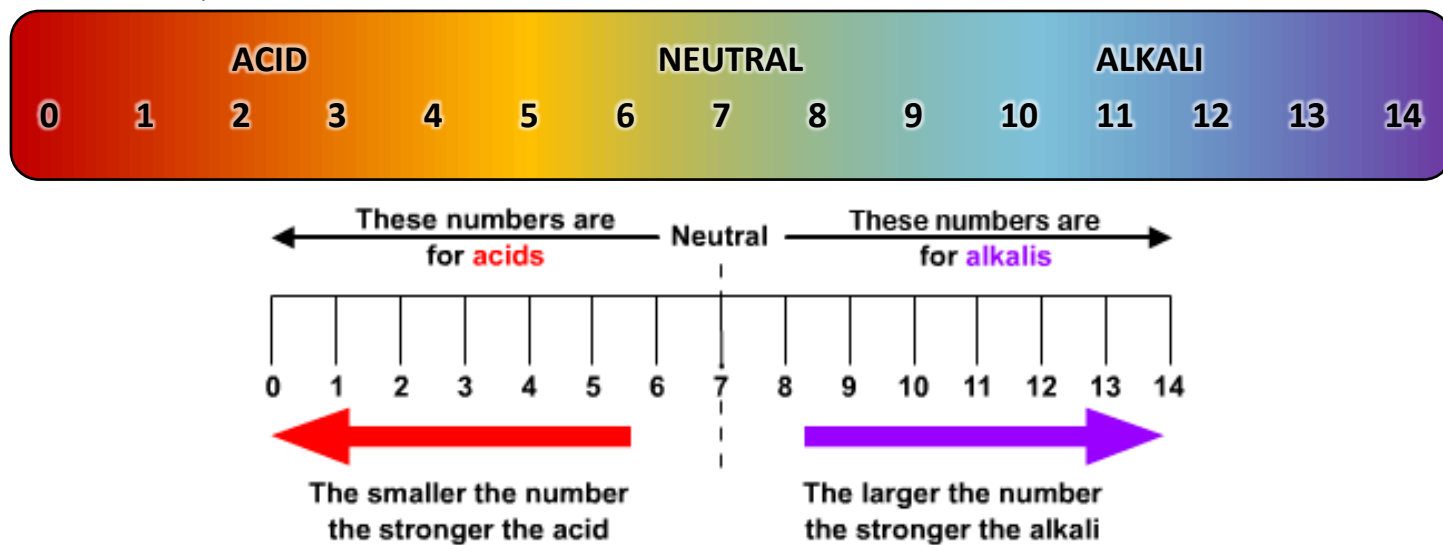
It is possible to tell if a solution is **acidic** or **alkaline** by using an **indicator**. An indicator is a substance which has different colours when it is in acidic or alkaline conditions.

Litmus is probably the most well-known indicator. This is **red in acids** and **blue in alkalis**. Litmus can be used as a liquid, or as litmus paper.

Solutions of acids and alkalis can vary widely in their acidity and alkalinity. It is useful to know not just whether a solution is an acid or an alkali, but **how** acidic or how alkaline it is.

To measure acidity and alkalinity, we can use the pH scale.

The easiest way to do this is to use **Universal indicator**. This is a mixture of several different indicators, and can be used as a liquid or paper. It has many different colour changes. The colour of the Universal indicator shows the pH value of the solution.



The pH scale runs from **pH 0 (strongest acid)** to **pH 14 (strongest alkali)**.

Strong acid pH 0-2 **Weak acid pH 3-5** **Neutral pH 7** **Weak alkali pH 9-11** **Strong alkali pH 12-14**

Acids, Alkalis and Bases

Acid
Release H^+ ions

Base:
Absorb H^+ ions
Alkali:
Soluble base

Metal oxides and **metal hydroxides** are known as **bases** and that an **alkali** is a **soluble base**.

e.g. MgO (magnesium oxide) is a base, but is insoluble in water.
 NaOH (sodium hydroxide) is an **alkali**, as it is a base that is **soluble** in water.

Neutralisation

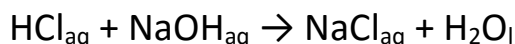
When an acidic compound dissolves in water it produces hydrogen ions, H^+ . These ions are responsible for the **acidity** of the solution.

When an alkaline compound dissolves in water it produces hydroxide ions, OH^- . These ions are responsible for the **alkalinity** of the solution.

Acids react with alkalis to form salts. These are called **neutralisation reactions**. In each reaction, water is also formed:



e.g. hydrochloric acid + sodium hydroxide \rightarrow sodium chloride + water



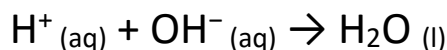
Hydrochloric acid contains **hydrogen ions** (H^+) and chloride ions (Cl^-) dissolved in water.

Sodium hydroxide solution contains sodium ions (Na^+) and **hydroxide ions** (OH^-) dissolved in water.

Higher tier

Hydrogen and hydroxide ions

The only change that does produce something different during the neutralisation is the reaction between hydrogen ions and hydroxide ions, which produces **water molecules**.



This is the ionic equation that represents the **neutralisation** reaction between any **acid** and any **alkali**.
The **name of the salt** produced in the neutralisation depends on **which acid** reacts with **which alkali**.

But the ionic equation for the neutralisation is always the same.

In the above reaction, the **salt** sodium chloride is formed when the acid and alkali are mixed together. This salt is produced as sodium ions and chloride ions in solution. There is no change to the sodium ions and chloride ions during the reaction to make sodium chloride. They were dissolved in water in the acid and in the alkali, and they are still dissolved in water as the salt.

Reactions of acids

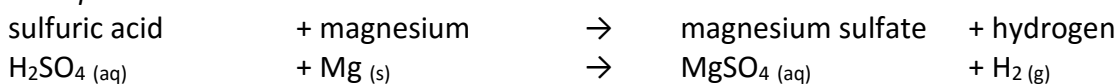
Acids **react** with metals, metal oxides, metal hydroxides and metal carbonates. During each of these reactions a salt is made.

1. Acid with metal

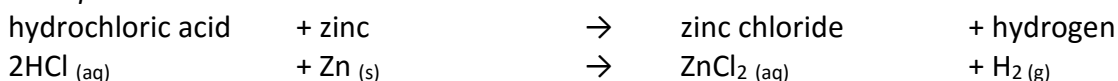
Acids will react with reactive **metals**, such as magnesium and zinc, to make a **salt** and **hydrogen**. This reaction is **exothermic** – releases heat.



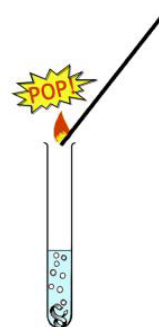
Example 1:



Example 2:

**Hydrogen test**

The hydrogen causes bubbling during the reaction. It can be detected using a **lighted splint**, which causes the gas to burn with a **squeaky pop**.



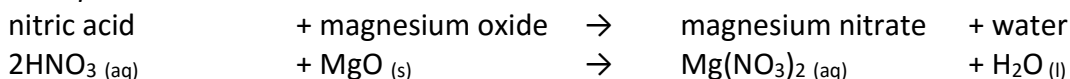
2. Acid with base

a. Acid with an insoluble base (metal oxide)

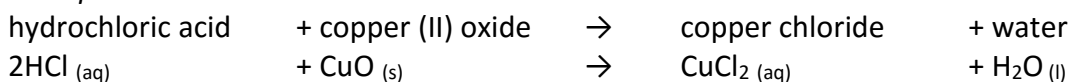
Acids react with **metal oxides**, such as magnesium oxide, to produce a **salt** and **water**. This reaction is **exothermic**.



Example 1:

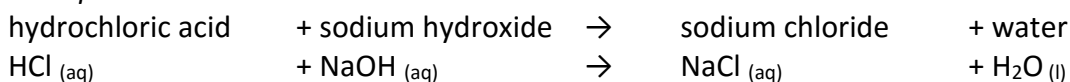
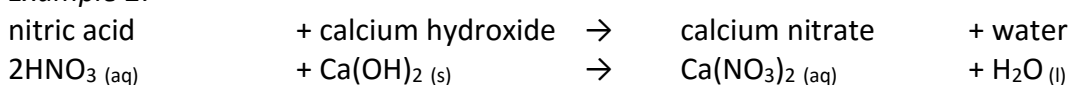


Example 2:

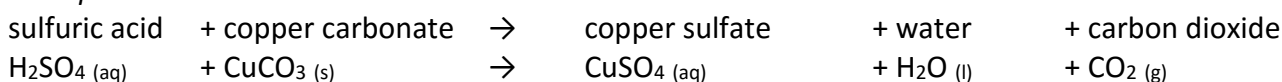
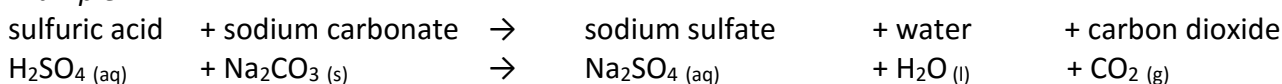
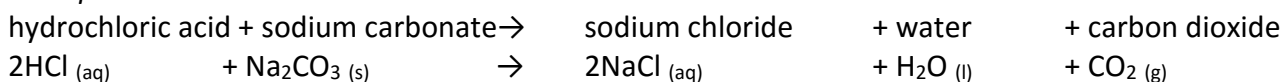


b. Acid with an alkali (metal hydroxide)

Acids react with **metal hydroxides**, such as sodium hydroxide, to produce a **salt** and **water**. This reaction is **exothermic**.

*Example 1:**Example 2:**c. Acid with a carbonate (metal carbonate)*

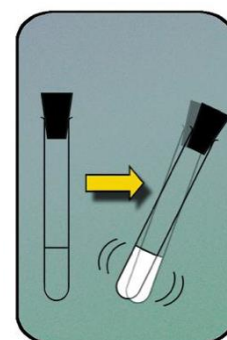
Acids react with **metal carbonates**, such as copper carbonate, to produce a **salt**, **water** and **carbon dioxide**. This reaction is **exothermic** and **effervesces (fizzes)**.

*Example 1:**Example 2:**Example 3:***Carbonate test**

When acid reacts with a carbonate **fizzing** is observed. Bubbles are seen, as CO₂ is a gas.

Carbon dioxide test

The carbon dioxide (CO₂) causes **bubbling/fizzing** during the reaction. It can be detected using **limewater**, which turns a **milky white** when **carbon dioxide** is bubbled through it.

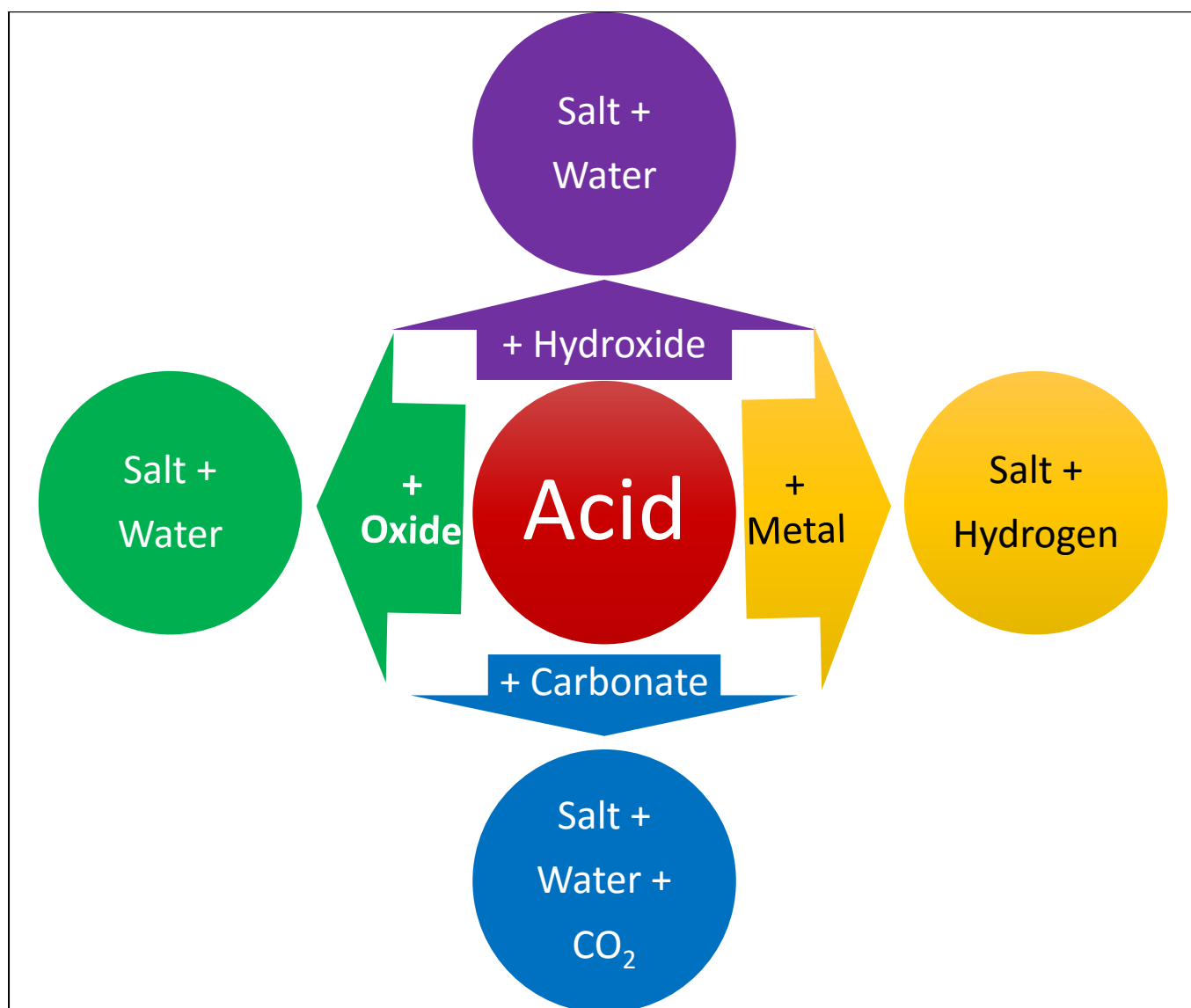


Summary of reactions of acids

Naming salts is an important aspect of acid reactions. You need to be able to name the salts formed from the acid.

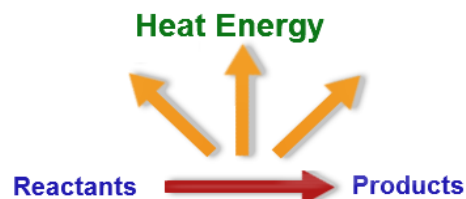
Acid**Salt formed**

1. Hydrochloric acid → metal chloride
2. Sulfuric acid → metal sulfate
3. Nitric acid → metal nitrate



Exothermic reactions

All neutralisation reactions are **exothermic** – heat is released.
This change can be measured by using a **thermometer**.

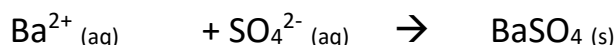


Sulfate test

Another test that you need to know is the **sulfate test**. This involves adding **barium chloride** solution to the test solution.

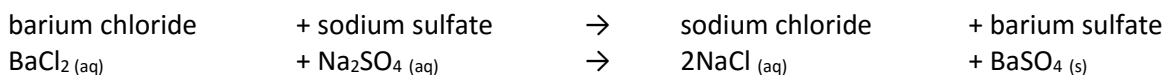
If there are **sulfate ions (SO_4^{2-})** present \rightarrow **white** precipitate forms

Ionic equation:



Example:

If sodium sulfate solution is tested:

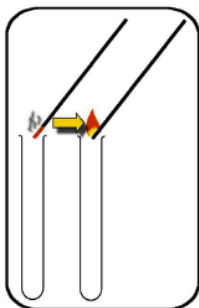


Tests for gases

Often, during chemical reactions, a gas will be given off, we need to be able to identify these gases.

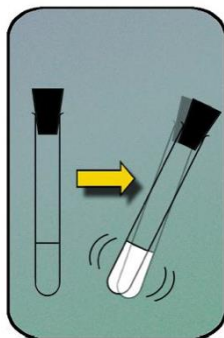
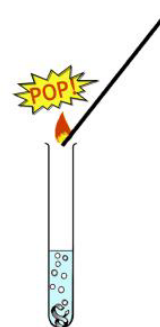
Oxygen

Oxygen will re-light a glowing splint



Hydrogen

Can be detected using a **lighted splint**, which causes the gas to burn with a **squeaky pop**.



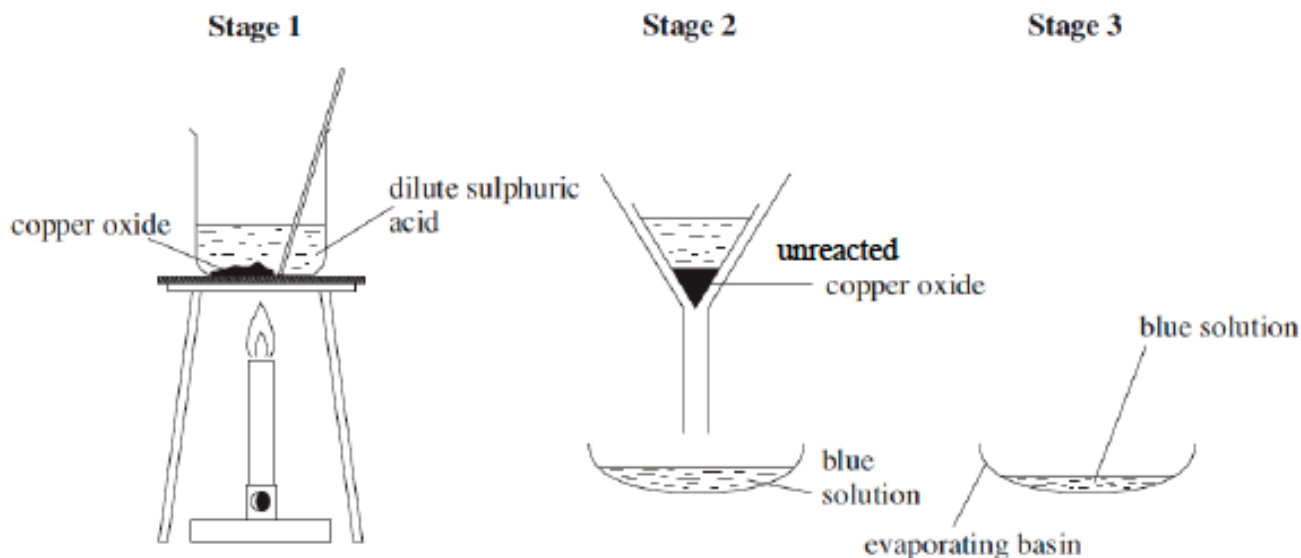
Carbon dioxide

Can be detected using **limewater**, which turns a **milky white** when **carbon dioxide** is bubbled through it.

Preparation of salt crystals

Using the neutralisation reactions of insoluble bases and carbonates, we can form salt crystals. You need to remember this method.

Example method – copper oxide and sulphuric acid to form copper sulfate crystals



Stage 1:

Excess base (copper oxide) is added to the dilute acid to make sure all the acid has been reacted and used up. Heating and stirring will assist the process

Stage 2:

The excess (unreacted) base is removed by the process of **filtration**, using a filter funnel and filter paper

Stage 3:

Salt is obtained by **evaporation** – water evaporates and crystals of salt left behind.

Large crystals - water can be evaporated slowly near a window over a period of a few days.

Small crystals - with additional heating to evaporate until about 1/3 of the solution remains, and the rest should be left to cool and evaporate naturally.

Obtaining salt crystals from the **metal carbonate** and **acid** reaction

The only difference in the method is stage 1 – excess metal carbonate is used – to make sure all the acid has been used up.

Obtaining salt crystals from the **metal** and **acid** reaction

The only difference in the method is stage 1 – excess metal is used – to make sure all the acid has been used up.

Titration

Another important method that is used in chemistry is titration, this involves using a **burette** to slowly add a solution to another.

Carrying out a titration

The **concentration** of an acid or alkali can be calculated by carrying out an experiment called a titration. You should be able to identify the apparatus needed to carry out a simple acid-alkali titration, and to describe how it is done.

Materials

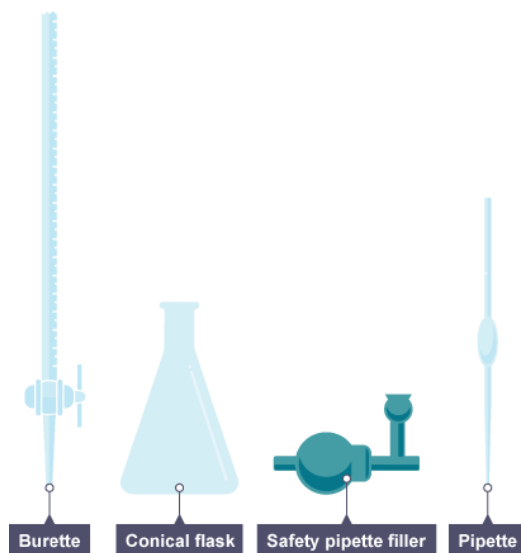
The apparatus needed includes a:

- **pipette** to accurately measure a certain volume of acid or alkali
- **pipette filler** to use the pipette safely
- **conical flask** to contain the liquid from the pipette
- **burette** to add small, measured volumes of one reactant to the other reactant in the conical flask

Example method

1. Use the pipette and pipette filler to add 25 cm³ of alkali to a clean conical flask.
2. Add a few drops of indicator and put the conical flask on a white tile (so you can see the colour of the indicator more easily).
3. Fill the burette with acid and note the starting volume.
4. Slowly add the acid from the burette to the alkali in the conical flask, swirling to mix.
5. Stop adding the acid when the end-point is reached (the appropriate colour change in the indicator happens). Note the final volume reading.
6. Repeat steps 1 to 5 until you get consistent readings

The same method works for adding an alkali to an acid - just swap around the liquids that go into the conical flask and burette.



The titre

The **difference** between the reading at the **start** and the **final** reading gives the volume of acid (or alkali) added. This volume is called the **titre**.

For example, if the reading at the start is 1.0 cm³ and the final reading is 26.5 cm³, then the titre is 25.5 cm³ (26.5 – 1.0). Note that the titre will depend upon the volume of liquid in the conical flask, and the concentrations of the acid and alkali used.

It is important to **repeat the titration several times** to check that your titre value is consistent so that your calculations are **reliable**.

Indicator

If universal indicator is used, the colour changes gradually through a range of colours. On the other hand, a single indicator like litmus or phenolphthalein gives a **sharp end-point** where the **colour changes suddenly**.

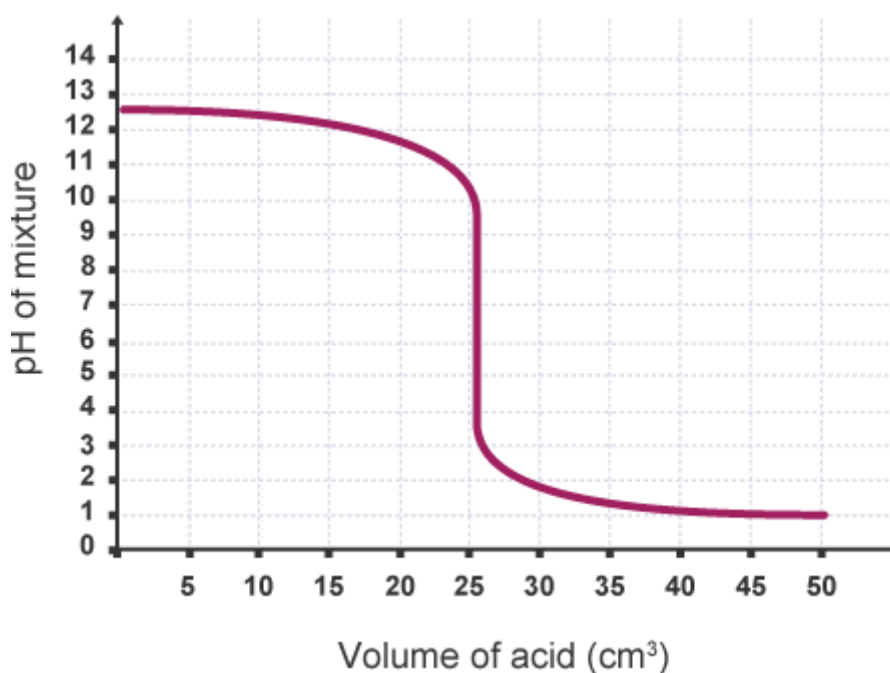
pH curves

The change in pH can be estimated using universal indicator, or measured more accurately using a **pH meter**, when an acid and an alkali are mixed.

A graph of pH (on the vertical axis) against total volume of acid or alkali added (on the horizontal axis) is called a **pH curve**. You need to be able to interpret a simple pH curve. If you are taking the Higher Tier paper you may have to sketch a pH curve.

1. Adding acid to alkali

The pH curve below shows what happens to the pH when a **strong acid** (such as hydrochloric acid) is added to 25 cm³ of a **strong alkali** (such as sodium hydroxide). The acid and the alkali started off at the same concentration.



Describing and explaining the graph

Graph shows how pH varies as increasing quantities of acid are added to an alkali. It starts at around 12.5 and decreases slowly until the volume of acid reaches 25 cm³, at which point pH falls sharply from 10 to 3 and then decreases slowly to 1, when volume of acid is 50 cm³.

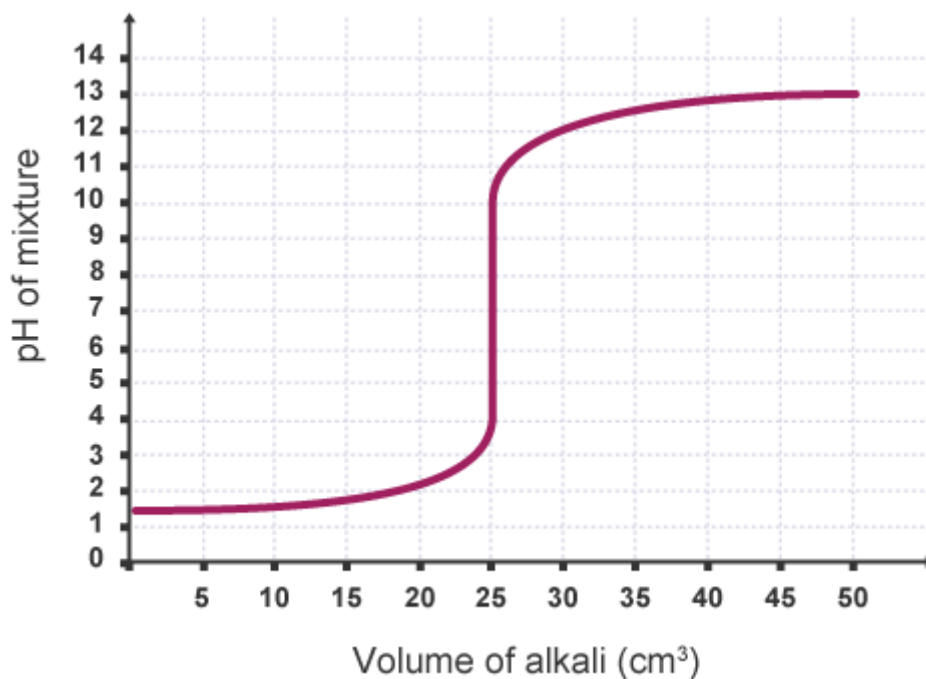
Note that the pH falls:

- slowly at first as acid is added to the alkali
- rapidly at the **end-point** (the point where the alkali is completely neutralised)
- slowly again once excess acid is being added

In this example, 25 cm³ of acid was needed to neutralise the alkali. If the acid had been more concentrated than the alkali, the volume needed would have been less than 25 cm³. The mixture was pH 12 at 20 cm³ and at pH 2 at 30 cm³. It was pH 7 at the **end-point**.

2. Adding alkali to acid

The pH curve below shows what happens to the pH when a **strong alkali** is added to 25 cm³ of a **strong acid**. As before, they both started off at the same concentration.



Describing and explaining the graph

Graph shows how pH varies as increasing quantities of alkali are added to an acid. It starts at around 1 and increases slowly until the volume of alkali reaches 25 cm³, at which point pH rises sharply from 3 to 10 and then increases slowly to 12.5, when volume of alkali is 50 cm³.

Note that the pH rises:

- slowly at first as alkali is added to the acid
- rapidly at the **end-point** (the point where the acid is completely neutralised)
- slowly again once excess alkali is being added

In this example, 25 cm³ of alkali was needed to neutralise the acid. If the acid had been more concentrated than the alkali, the volume needed would have been more than 25 cm³.

Preparing a salt crystals using titration then evaporation

You should know the method used to prepare crystals of soluble salts from the reaction of acids with alkalis.

The method is the same as the titration earlier in this topic, except:

- same fixed volume of acid/alkali in clean flask and exact volume of alkali/acid needed for neutralisation is added **but with no indicator**;

Then, we use the crystallisation method from earlier to evaporate the water to form crystals.



GCSE Chemistry ONLY

Double award - move on to **Topic 3**

Concentration

Following on from Unit 1, you should be able to convert number of moles into mass. From this we can work out the concentration of a solution.

Concentration is the number of moles (or the mass) per dm³.

Equations for working out the concentration:

$$\text{Concentration (mol/dm}^3\text{)} = \frac{\text{number of moles}}{\text{volume (dm}^3\text{)}}$$

$$\text{Concentration (g/dm}^3\text{)} = \frac{\text{mass (g)}}{\text{volume (dm}^3\text{)}}$$

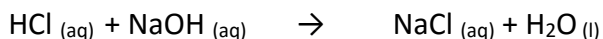
Higher tier

Titration calculations

You should be able to use **titration** results to calculate the **concentration** of an acid or alkali. If several runs have been carried out, any **irregular titres should be ignored** before calculating the **mean titre**.

Example 1 (1:1 mole ratio)

From an experiment, we record that 27.5 cm³ of 0.2 mol/dm³ hydrochloric acid is needed to titrate 25.0 cm³ of sodium hydroxide solution. What is the **concentration** of the sodium hydroxide solution?



Step 1: Convert all volumes to dm³

$$\text{HCl} \quad 27.5 \text{ cm}^3 = 27.5 \div 1000 = 0.0275 \text{ dm}^3$$

$$\text{NaOH} \quad 25.0 \text{ cm}^3 = 25.0 \div 1000 = 0.025 \text{ dm}^3$$

Step 2: Calculate the number of moles of the substance where the volume and concentration are known

Rearranging the above equation gives us:

$$\text{number of moles} = \text{concentration} \times \text{volume}$$

$$\text{number of moles of hydrochloric acid} = 0.2 \times 0.0275 = 0.0055 \text{ mol} \quad (5.5 \times 10^{-3} \text{ mol})$$

Step 3: Calculate the unknown concentration

Because this reaction is a **1:1 reaction**, i.e. 1 H⁺ from the HCl reacts with 1 OH⁻ from the NaOH, we can say that **0.0055 mol of acid** will react with **0.0055 mol of alkali**

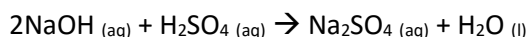
Rearranging the above equation again gives us:

$$\text{concentration} = \text{number of moles} \div \text{volume}$$

$$\text{concentration of alkali} = \text{moles} \div \text{volume} = 0.0055 \div 0.025 = 0.22 \text{ mol/dm}^3$$

Example 2 (2:1 mole ratio)

25.0 cm³ of sodium hydroxide (NaOH) solution of unknown concentration was titrated with dilute sulfuric acid (H₂SO₄) of concentration 0.050 mol/dm⁻³. 20.0 cm³ of the acid was required to neutralise the alkali. Find the concentration of the sodium hydroxide solution in mol/dm⁻³



Step 1: Convert all volumes to dm³

$$\text{H}_2\text{SO}_4 \quad 20.0 \text{ cm}^3 = 20.0 \div 1000 = \underline{0.020 \text{ dm}^3}$$

$$\text{NaOH} \quad 25.0 \text{ cm}^3 = 25.0 \div 1000 = \underline{0.025 \text{ dm}^3}$$

Step 2: Calculate the number of moles of the substance where the volume and concentration are known

Rearranging the above equation gives us:

$$\text{number of moles} = \text{concentration} \times \text{volume}$$

$$\text{number of moles of hydrochloric acid} = 0.050 \times 0.020 = \underline{0.001 \text{ mol}} \quad (1 \times 10^{-3} \text{ mol})$$

Step 3: Calculate the unknown concentration

Because this reaction is a **2:1 reaction**, i.e. 2 H⁺ from the H₂SO₄ are released, so **1 H₂SO₄ : 2 NaOH**

We can say that **0.001 mol of acid** will react with **0.002 mol of alkali** (0.001 x 2)

Rearranging the above equation again gives us:

$$\text{concentration} = \text{number of moles} \div \text{volume}$$

$$\text{concentration of alkali} = \text{moles} \div \text{volume} = 0.002 \div 0.025 = \underline{0.08 \text{ mol/dm}^3}$$

Try these:

1. A student reacted 25 cm³ of 1 mol/dm⁻³ HCl with 25 cm³ of 1 mol/dm⁻³ NaOH, is the solution alkali, acidic or neutral?
2. A student reacted 25 cm³ of acid (HCl) with 20 cm³ alkali (NaOH) to make a neutral solution, which had the highest concentration, the acid or the alkali?
3. A student reacted 10 cm³ of acid (HCl) with 40 cm³ alkali (NaOH) to make a neutral solution, which had the weakest concentration, the acid or the alkali?

Strong and weak acids

Acid solutions contain **hydrogen ions**. The higher the **concentration of hydrogen ions**, the **lower the pH**. Hydrochloric acid is a **strong acid** and ethanoic acid is a **weak acid**.

Strong acids are **fully ionised** but weak acids are only **partly ionised** in solution. At the same concentration, strong acids have a higher concentration of hydrogen ions than weak acids.

Acids ionise in water to produce hydrogen ions, H^+ .

Strong acids fully ionise. For example:

- hydrochloric acid: $HCl \rightarrow H^+ + Cl^-$
- nitric acid: $HNO_3 \rightarrow H^+ + NO_3^-$
- sulfuric acid: $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

Weak acids do not fully ionise. Instead, they form an **equilibrium** mixture. For example:

- ethanoic acid: $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

At the same concentration, strong acids have lower pH than weak acids.

pH and hydrogen ion concentration

The **pH** of a solution is related to its **concentration of hydrogen ions** - the higher the concentration of hydrogen ions H^+ the lower the pH.

At the same concentration of acid, the concentration of hydrogen ions will be higher in a strong acid than in a weak acid. This is why the pH of a strong acid like hydrochloric acid will be lower than the pH of a weak acid like ethanoic acid.

Acid strength vs concentration

The **strength of an acid** is a measure of the degree of its **ionisation** - strong acids are fully ionised but weak acids are only partly ionised. Remember: the opposite of strong is weak.

The **concentration of an acid** is a measure of the **number of moles of acid in 1 dm³ of acid solution**. For example, 2 mol/dm³ hydrochloric acid is twice as concentrated as 1 mol/dm³ hydrochloric acid or 1 mol/dm³ ethanoic acid. Remember: the opposite of concentrated is dilute.

Reactions of acids - further

Both strong acids and weak acids will react with magnesium or with calcium carbonate. However, the **rate of reaction** will be different – a **strong acid** produces a **faster reaction** than a weak acid at the same concentration.

Metal and acid

Magnesium reacts with acids to produce a magnesium salt and hydrogen:

- magnesium + hydrochloric acid → magnesium chloride + hydrogen
- magnesium + ethanoic acid → magnesium ethanoate + hydrogen

The rate of reaction is determined by measuring the volume of hydrogen produced as the reaction goes on.

Note that **ethanoic acid** produces **ethanoate** salts.

Carbonate and acid

Calcium carbonate reacts with acids to produce a calcium salt, water and carbon dioxide:

- calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide
- calcium carbonate + ethanoic acid → calcium ethanoate + water + carbon dioxide

The rate of reaction is determined by measuring the volume of carbon dioxide produced as the reaction goes on.

Volume of gas produced

If the **same concentration** and **volume** of acid is used in these reactions, the **same volume of gas** is produced whether hydrochloric acid or ethanoic acid is used.

This is because both acids contain the **same amount of acid reactant** (same number of moles). The strength of the acid does not matter here.

Rate of reaction

If the **same concentration** and **volume** of acid is used in these reactions, **ethanoic acid** will **react more slowly** than hydrochloric acid. The reaction will also be **less exothermic**.

This is because the ethanoic acid will contain fewer hydrogen ions in a solution (it will have a lower concentration of hydrogen ions). There will be fewer collisions between hydrogen ions and particles of the magnesium or calcium carbonate, so the rate of reaction will be lower than for hydrochloric acid.

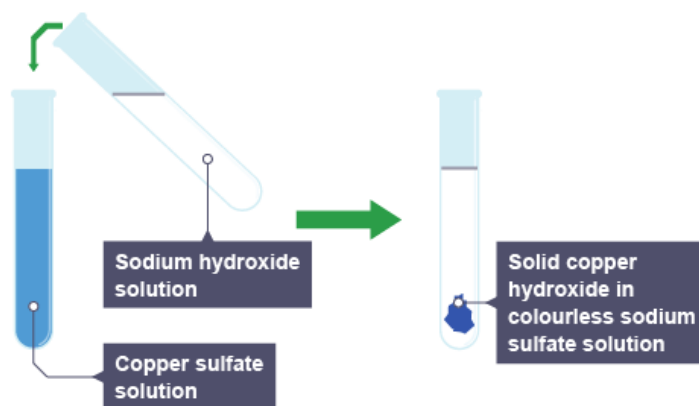
Precipitates

Ionic substances contain charged particles which are called **ions**. These are in fixed positions in solids, but they are free to move when they are molten or in solution. Most **precipitation** reactions involve ions from one solution reacting with ions from another solution.

In a precipitation reaction, **ions collide with one another to form an insoluble product** (one that does not dissolve in water). This is the **precipitate**.

Most precipitation reactions are very fast because there is a high chance of collisions between ions in solution. The precipitate forms as soon as two suitable solutions are mixed together

For example, copper sulfate solution is clear and blue, while sodium hydroxide solution is clear and colourless. A blue precipitate of copper hydroxide immediately forms when they are mixed.



Making an insoluble compound

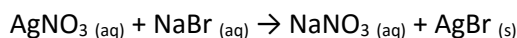
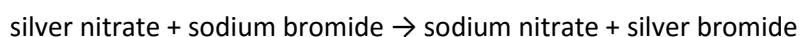
Insoluble compounds may be made by precipitation reactions.

Three main stages are involved:

1. Mixing the required reactant solutions.
2. Filtration to remove soluble impurities.
3. Washing and drying the residue (the insoluble compound that remains in the filter paper).

1. Mixing

All nitrates are soluble, and all sodium salts are soluble. This means that, for example, if you want to make insoluble silver bromide you can mix together silver nitrate solution and sodium bromide solution:

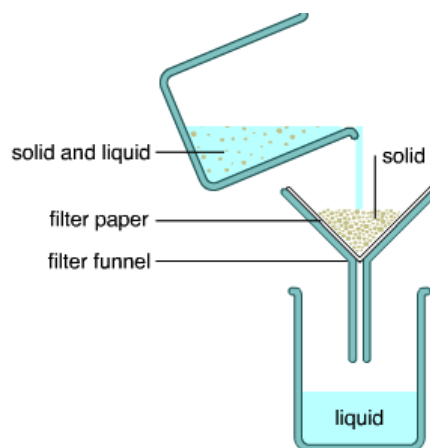


2. Filtration

The **insoluble precipitate** must be separated from the soluble impurities using **filtration**. The precipitate stays behind in the filter paper as a residue, while the soluble impurities pass through in the filtrate.

3. Washing and drying

The precipitate can be **washed** while it is still in the filter funnel. Water cannot dissolve the precipitate, but it can wash off any remaining soluble impurities. The filter paper can then be removed and opened out flat. The precipitate is then **dried in a warm oven**.



Topic 4 - CHEMICAL REACTIONS AND ENERGY

Exothermic and Endothermic reactions

Changes in temperature happen often during chemical reactions.

Exothermic reactions

An **exothermic** reaction is a reaction where the **temperature rises**. e.g. magnesium and acid.

Exothermic reactions will feel warm/hot. This is because energy is being **released** from the reaction to its surroundings.

Combustion and **neutralisation** are common examples of exothermic reactions.

Endothermic reactions

An **endothermic** reaction is a reaction where the **temperature falls**. e.g. ammonium nitrate and acid.

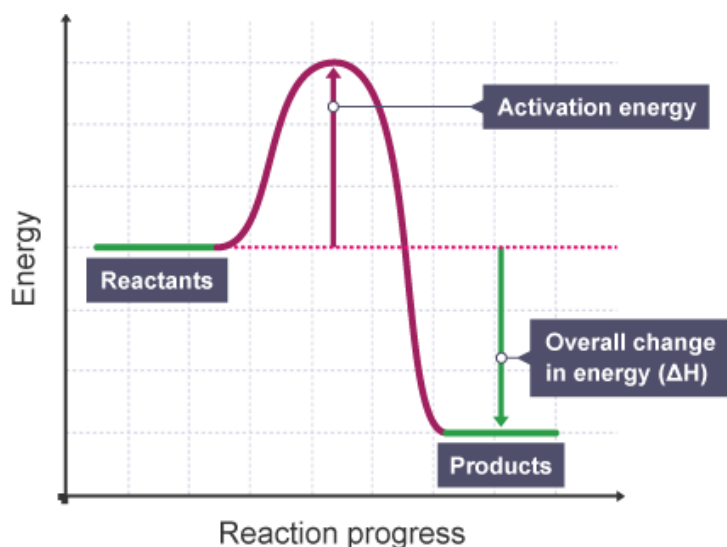
Endothermic reactions will feel cold. This is because energy is being **absorbed** by the reaction from its surroundings.

Examples are: electrolysis, the reaction between ethanoic acid and sodium carbonate and the thermal decomposition of calcium carbonate in a blast furnace.

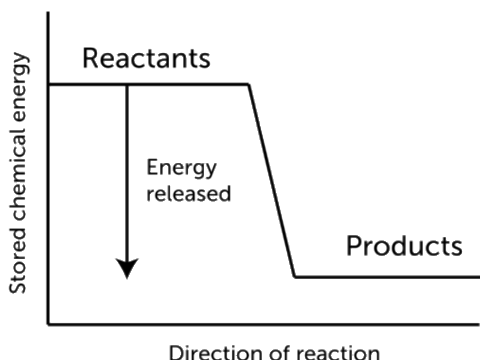
Activation energy

Simple energy level diagrams only show the **energy levels** at the **beginning** and **end** of a reaction (like those above).

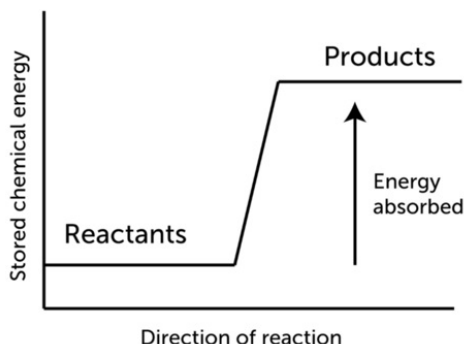
Energy levels **change gradually during a reaction**, and this can be shown using a curve between the reactant and product energy levels.



Exothermic Reaction



Endothermic Reaction



Notice that in this diagram, **energy is required to start the reaction**.

This is the **minimum amount** of energy required to start a reaction, and is known as the **activation energy**.

Notice that the overall change in energy during this diagram is **negative**...

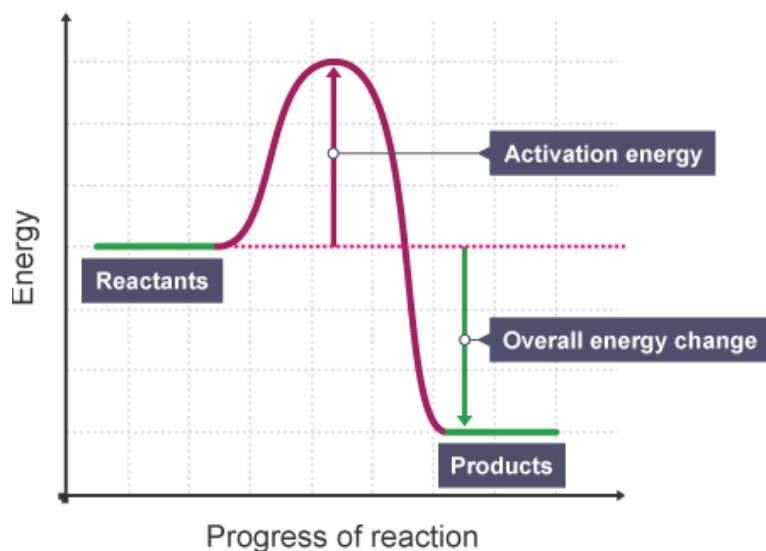
The products have **less energy** than the reactants. This “lost” energy has been given out as **heat**. So, this reaction is **exothermic**.

Energy profiles

You will need to be able to sketch energy profiles for your exam, including labels.

Here are two common energy profiles:

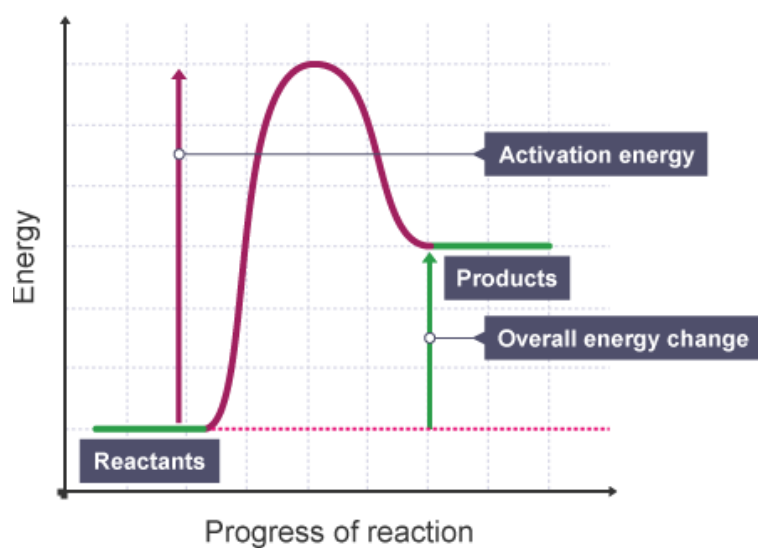
Exothermic:



Notice that the overall energy change is **negative**.

This means energy is **given out** to its surroundings.

Endothermic:



Notice that the overall energy change this time is **positive**.

This means energy is **taken from** its surroundings.

Often the activation energy required for endothermic reactions is large.

Bonds and energy

During a chemical reaction:

1. bonds in the **reactants** are broken
2. **new bonds are made** in the **products**

Energy is needed to **break bonds**, and **energy** is **released** when **bonds are made**.

In an **exothermic reaction**, more energy is released when new bonds are made than is needed to break existing bonds.

In an **endothermic reaction**, more energy is needed to break existing bonds than is released when new bonds are made.

Calculating bond energies

You can calculate the energy change in a reaction using **bond energies**. A bond energy is the amount of energy needed to break a mole of a particular bond. You will be given any bond energies you need in the exam.

Method

1. Add together all the bond energies to **break all the bonds in the reactants** – this is the '**energy in**'.
2. Add together the bond energies to **form all the bonds in the products** – this is the '**energy out**'.
3. Calculate the energy change: **energy in** – **energy out**.

Example 1 – an exothermic reaction

Hydrogen and chlorine react to form hydrogen chloride gas:



Bond energies relevant to this reaction:

Bond	Bond Energy (kJ/mole)
H-H	436
Cl-Cl	243
H-Cl	432

$$\text{Energy in} = 436 \text{ (breaking H-H)} + 243 \text{ (breaking Cl-Cl)} = \underline{679 \text{ kJ/mole (Bonds broken)}}$$

$$\text{Energy out} = 2 \times 432 \text{ (forming H-Cl – twice)} = \underline{864 \text{ kJ/mole (Bonds formed)}}$$

$$\text{Energy change} = \text{Energy in} - \text{Energy out} = 679 - 864 = \underline{-185 \text{ kJ/mole}}$$

The **energy change is negative**, showing that energy is released to the surroundings in an **exothermic** reaction.

Bonds formed is larger than **bonds broken** so the reaction is **exothermic**.

Example 2 – an endothermic reaction

Hydrogen bromide decomposes to form hydrogen and bromine:



Bond energies relevant to this reaction:

Bond	Bond Energy (kJ/mole)
H-Br	366
H-H	436
Br-Br	193

$$\text{Energy in} = 2 \times 366 = \underline{732 \text{ kJ/mole}}$$

$$\text{Energy out} = 436 + 193 = \underline{629 \text{ kJ/mole}}$$

$$\text{Energy change} = \text{Energy in} - \text{Energy out} = 732 - 629 = \underline{+103 \text{ kJ/mole}}$$

The **energy change is positive**, showing that energy is taken in from the surroundings in an **endothermic** reaction.

Bonds broken is larger than **bonds formed** so the reaction is **endothermic**.

Table of Ions

POSITIVE IONS		NEGATIVE IONS	
Name	Formula	Name	Formula
Aluminium	Al^{3+}	Bromide	Br^-
Ammonium	NH_4^+	Carbonate	CO_3^{2-}
Barium	Ba^{2+}	Chloride	Cl^-
Calcium	Ca^{2+}	Fluoride	F^-
Copper(II)	Cu^{2+}	Hydroxide	OH^-
Hydrogen	H^+	Iodide	I^-
Iron(II)	Fe^{2+}	Nitrate	NO_3^-
Iron(III)	Fe^{3+}	Oxide	O^{2-}
Lithium	Li^+	Sulfate	SO_4^{2-}
Magnesium	Mg^{2+}		
Nickel	Ni^{2+}		
Potassium	K^+		
Silver	Ag^+		
Sodium	Na^+		
Zinc	Zn^{2+}		

Periodic Table of Elements

1		2		Group										3		4		5		6		7		0			
				<div><div><div><div><div><div></div><div>${}^1_1\text{H}$</div><div>Hydrogen</div></div></div></div></div></div>																							
<div><div><div><div><div></div><div>${}^7_3\text{Li}$</div><div>Lithium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^9_4\text{Be}$</div><div>Beryllium</div></div></div></div></div>											<div><div><div><div><div></div><div>${}^{11}_5\text{B}$</div><div>Boron</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{12}_6\text{C}$</div><div>Carbon</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{14}_7\text{N}$</div><div>Nitrogen</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{16}_8\text{O}$</div><div>Oxygen</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{19}_9\text{F}$</div><div>Fluorine</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{20}_{10}\text{Ne}$</div><div>Neon</div></div></div></div></div>										
<div><div><div><div><div></div><div>${}^{23}_{11}\text{Na}$</div><div>Sodium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{24}_{12}\text{Mg}$</div><div>Magnesium</div></div></div></div></div>											<div><div><div><div><div></div><div>${}^{27}_{13}\text{Al}$</div><div>Aluminium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{28}_{14}\text{Si}$</div><div>Silicon</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{31}_{15}\text{P}$</div><div>Phosphorus</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{32}_{16}\text{S}$</div><div>Sulfur</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{35}_{17}\text{Cl}$</div><div>Chlorine</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{40}_{18}\text{Ar}$</div><div>Argon</div></div></div></div></div>										
<div><div><div><div><div></div><div>${}^{39}_{19}\text{K}$</div><div>Potassium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{40}_{20}\text{Ca}$</div><div>Calcium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{45}_{21}\text{Sc}$</div><div>Scandium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{48}_{22}\text{Ti}$</div><div>Titanium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{51}_{23}\text{V}$</div><div>Vanadium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{52}_{24}\text{Cr}$</div><div>Chromium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{55}_{25}\text{Mn}$</div><div>Manganese</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{56}_{26}\text{Fe}$</div><div>Iron</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{59}_{27}\text{Co}$</div><div>Cobalt</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{59}_{28}\text{Ni}$</div><div>Nickel</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{64}_{29}\text{Cu}$</div><div>Copper</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{65}_{30}\text{Zn}$</div><div>Zinc</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{70}_{31}\text{Ga}$</div><div>Gallium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{73}_{32}\text{Ge}$</div><div>Germanium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{75}_{33}\text{As}$</div><div>Arsenic</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{79}_{34}\text{Se}$</div><div>Selenium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{80}_{35}\text{Br}$</div><div>Bromine</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{84}_{36}\text{Kr}$</div><div>Krypton</div></div></div></div></div>										
<div><div><div><div><div></div><div>${}^{86}_{37}\text{Rb}$</div><div>Rubidium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{88}_{38}\text{Sr}$</div><div>Strontium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{89}_{39}\text{Y}$</div><div>Yttrium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{91}_{40}\text{Zr}$</div><div>Zirconium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{93}_{41}\text{Nb}$</div><div>Niobium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{96}_{42}\text{Mo}$</div><div>Molybdenum</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{99}_{43}\text{Tc}$</div><div>Technetium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{101}_{44}\text{Ru}$</div><div>Ruthenium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{103}_{45}\text{Rh}$</div><div>Rhodium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{106}_{46}\text{Pd}$</div><div>Palladium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{108}_{47}\text{Ag}$</div><div>Silver</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{112}_{48}\text{Cd}$</div><div>Cadmium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{115}_{49}\text{In}$</div><div>Indium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{119}_{50}\text{Sn}$</div><div>Tin</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{122}_{51}\text{Sb}$</div><div>Antimony</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{128}_{52}\text{Te}$</div><div>Tellurium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{127}_{53}\text{I}$</div><div>Iodine</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{131}_{54}\text{Xe}$</div><div>Xenon</div></div></div></div></div>										
<div><div><div><div><div></div><div>${}^{133}_{55}\text{Cs}$</div><div>Caesium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{137}_{56}\text{Ba}$</div><div>Barium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{139}_{57}\text{La}$</div><div>Lanthanum</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{179}_{72}\text{Hf}$</div><div>Hafnium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{181}_{73}\text{Ta}$</div><div>Tantalum</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{184}_{74}\text{W}$</div><div>Tungsten</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{186}_{75}\text{Re}$</div><div>Rhenium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{190}_{76}\text{Os}$</div><div>Osmium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{192}_{77}\text{Ir}$</div><div>Iridium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{195}_{78}\text{Pt}$</div><div>Platinum</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{197}_{79}\text{Au}$</div><div>Gold</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{201}_{80}\text{Hg}$</div><div>Mercury</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{204}_{81}\text{Tl}$</div><div>Thallium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{207}_{82}\text{Pb}$</div><div>Lead</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{209}_{83}\text{Bi}$</div><div>Bismuth</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{210}_{84}\text{Po}$</div><div>Polonium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{210}_{85}\text{At}$</div><div>Astatine</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{222}_{86}\text{Rn}$</div><div>Radon</div></div></div></div></div>										
<div><div><div><div><div></div><div>${}^{223}_{87}\text{Fr}$</div><div>Francium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{226}_{88}\text{Ra}$</div><div>Radium</div></div></div></div></div>	<div><div><div><div><div></div><div>${}^{227}_{89}\text{Ac}$</div><div>Actinium</div></div></div></div></div>	Key:																								

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