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States of Matter	Atoms, elements, compounds and mixtures	Structure of the Atom	Electronic Structure	Development of the atomic model	Isotopes	Isotope Calculations	Check 20	Misconceptions lesson
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Key Words

Key Word	definition
Atom	Smallest part of an element that can exist
Element	Made of 1 type of atom. Found on the periodic table
Molecule	Two or more atoms chemically bonded together
Compound	Two or more elements chemically combined in fixed proportions
Mixtures	Two or more elements or compounds not chemically combined
Isotopes	Atoms of the same element with the same number of protons and electrons but different numbers of neutrons

Misconceptions

Evaporation and boiling are not the same. Liquids can evaporate below their boiling point. e.g, water evaporates from a glass left on a windowsill

Compounds cannot be found on the periodic table. ONLY elements e.g. Carbon dioxide (CO₂) is a compound so is not on the periodic table

Molecules can be element or compounds e.g., O₂ is an element and a molecule. H₂O is a compound and a molecule

The nucleus of an atom is positive. This is because it contains neutrons (0 charge) and protons (+1 charge)

Key questions

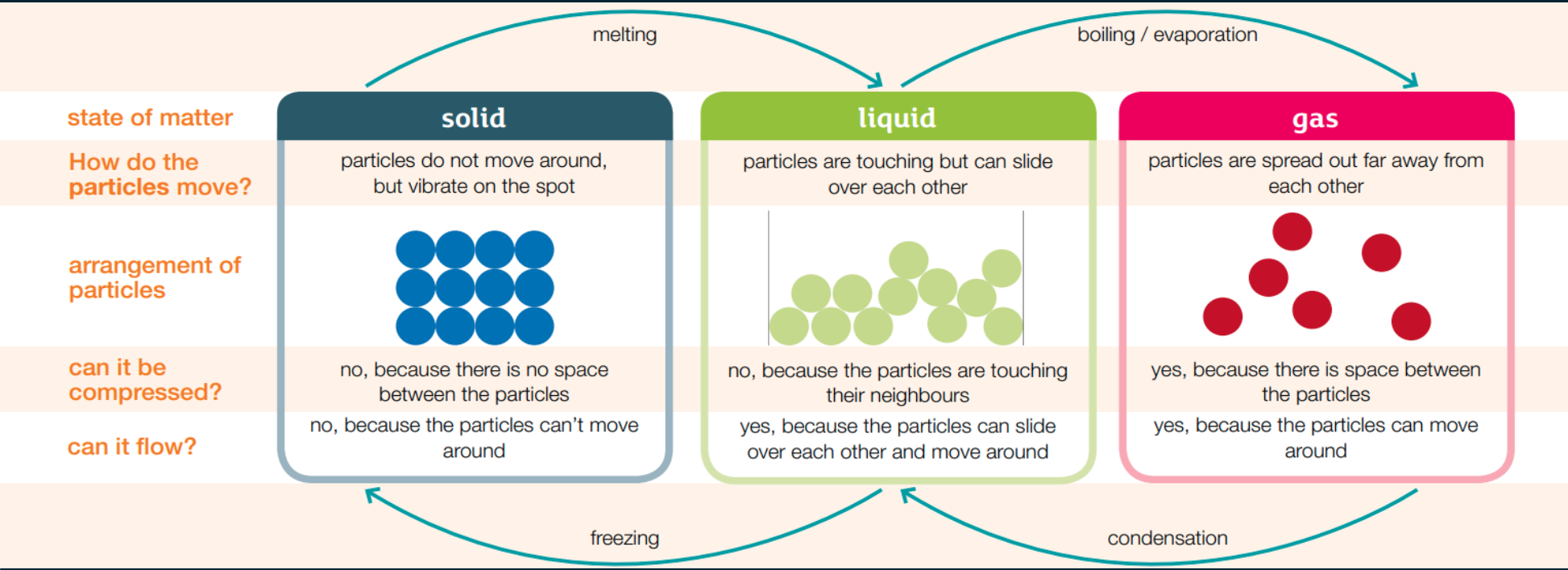
How does a solid become a liquid?

What is the overall charge of an atom?

What is the difference between a compound and a mixture?

What is an isotope?

States of Matter



Limitations of the particle model

The particle model assumes that particles are solid spheres with no forces between them.

However:

- particles are not solid, since atoms are mostly empty space
- many particles are not spherical

Changes of state

Changes of state happen when energy is transferred, this is done by heating or cooling

Heating
Particles gain kinetic energy. In a solid they vibrate faster, in a liquid they move faster, until the forces weaken enough for the substance to melt (solid to liquid) or boil (liquid to gas)

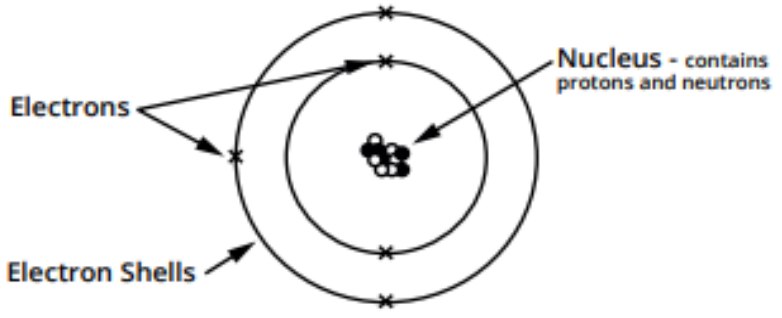
Cooling
Particles lose kinetic energy. Particles move more slowly, and forces strengthen until the substance condenses (gas to liquid) or freezes (liquid to solid)

State symbols

(s) = solid
(l) = liquid
(g) = gas
(aq) = aqueous*

*this means dissolved in water

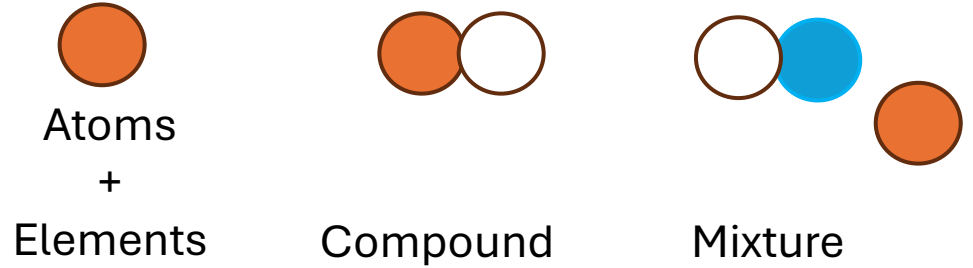
Structure of the atom:



Size of atoms

Atoms have a radius of 0.1nm ($1 \times 10^{-10} \text{m}$)
The radius of the nucleus is 1/10 000 of that of an atom ($1 \times 10^{-14} \text{m}$)

Atoms, elements, compounds and mixtures



Subatomic particles

Sub-atomic particle	Mass	Charge
Proton	1	+1
Electron	1/2000 or very small	-1
Neutron	1	0

Electron Shells

Electronic shell	Max number of electrons
1	2
2	8
3	8

Isotopes

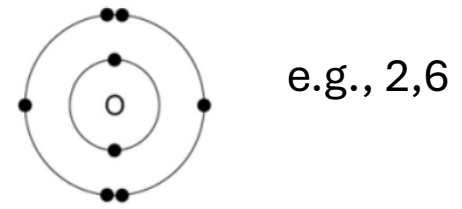
Atoms of the same element with the same number of protons and electrons but different numbers of neutrons

Charge of an atom

The number of protons (+) and electrons (-) are always equal in an atom – so atoms have no overall charge

Electron Configuration

Number of electrons on each shell



Isotopes Calculations

$$\text{Average relative atomic mass} = \frac{(\text{mass} \times \%) + (\text{mass} \times \%)}{100}$$

35
Cl
75%

37
Cl
25%

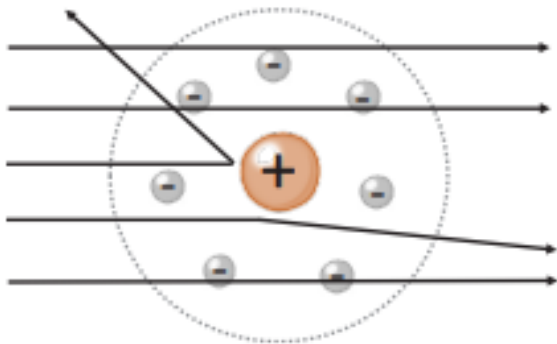
e.g.

$$\frac{(35 \times 75) + (37 \times 25)}{100} = 35.5$$

The development of the model of the atom:

Rutherford's scattering experiment

A beam of alpha particles are directed at a very thin gold foil



Conclusion:

- Most of the alpha particles passed straight through telling us the atom was mostly empty space
- A small number of the +alpha particles were deflected telling us there was a small positive mass at the center (the nucleus)

Dalton



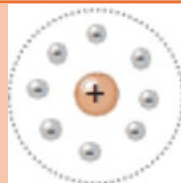
Atoms are solid indivisible spheres

JJ Thomson – Plum pudding model



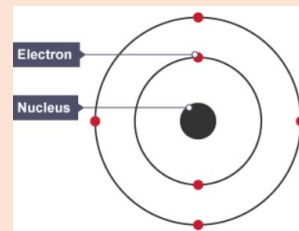
Solid ball of positive charge with electrons embedded

Rutherford – nuclear model



Atoms have a central positive nucleus (mass is concentrated here) and a cloud of electrons

Bohr

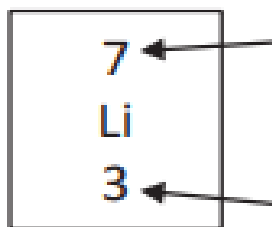


Put on electrons on shells at fixed distances from the nucleus

Chadwick

Discovered the neutron

Elements on periodic table



Mass number

The sum of the protons and neutrons in the nucleus

Atomic number

The number of protons in the atom

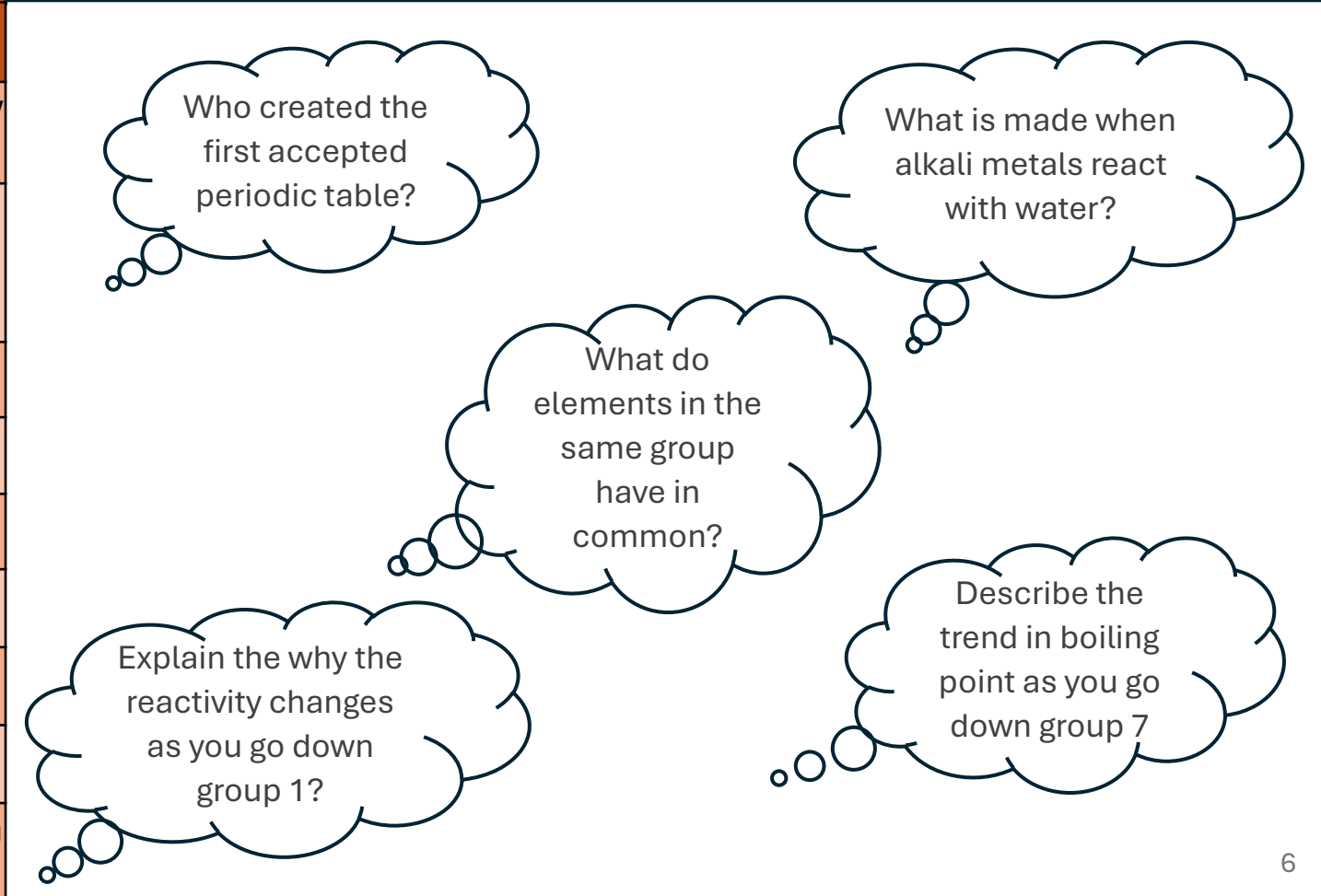
Number of electrons = number of protons

Periodic Table	Development of the periodic table	Metals and non-metals	Alkali Metals	Reactivity of alkali metals	Transition metals	Halogens	Reactivity of halogens	Group 0	Naming compounds	Check 20	Misconceptions lesson
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Key Words

Key questions

Key Word	Definition
Period	Row on the periodic table, it tells you how many electron shells
Group	Column on the periodic table, tells you the number of electrons on the outer shell
Alkali metals	Elements in group 1
Halogens	Elements in group 7
Noble Gases	Elements in group 0
Ions	Charges particles
Metals	Elements that react to form + ions
Non-metals	Elements that react to form - ions
Displacement reaction	When a more reactive element replaces a less reactive element in a compound



Who created the first accepted periodic table?

What is made when alkali metals react with water?

What do elements in the same group have in common?

Describe the trend in boiling point as you go down group 7

Explain the why the reactivity changes as you go down group 1?

Periodic table

Development of the periodic table

The periodic table is **ordered by atomic proton number**

1	2											3	4	5	6	7	0										
		<div style="border: 1px solid black; padding: 5px; display: inline-block;"> Key relative atomic mass atomic symbol name atomic (proton) number </div>																									
7 Li lithium 3	9 Be beryllium 4											11 B boron 5	12 C carbon 6	14 N nitrogen 7	16 O oxygen 8	19 F fluorine 9	20 Ne neon 10										
23 Na sodium 11	24 Mg magnesium 12											27 Al aluminium 13	28 Si silicon 14	31 P phosphorus 15	32 S sulfur 16	35.5 Cl chlorine 17	40 Ar argon 18										
39 K potassium 19	40 Ca calcium 20	45 Sc scandium 21	48 Ti titanium 22	51 V vanadium 23	52 Cr chromium 24	55 Mn manganese 25	56 Fe iron 26	59 Co cobalt 27	59 Ni nickel 28	63.5 Cu copper 29	65 Zn zinc 30	70 Ga gallium 31	73 Ge germanium 32	75 As arsenic 33	79 Se selenium 34	80 Br bromine 35	84 Kr krypton 36										
85 Rb rubidium 37	88 Sr strontium 38	89 Y yttrium 39	91 Zr zirconium 40	93 Nb niobium 41	96 Mo molybdenum 42	[97] Tc technetium 43	101 Ru ruthenium 44	103 Rh rhodium 45	106 Pd palladium 46	108 Ag silver 47	112 Cd cadmium 48	115 In indium 49	119 Sn tin 50	122 Sb antimony 51	128 Te tellurium 52	127 I iodine 53	131 Xe xenon 54										
133 Cs caesium 55	137 Ba barium 56	139 La* lanthanum 57	178 Hf hafnium 72	181 Ta tantalum 73	184 W tungsten 74	186 Re rhenium 75	190 Os osmium 76	192 Ir iridium 77	195 Pt platinum 78	197 Au gold 79	201 Hg mercury 80	204 Tl thallium 81	207 Pb lead 82	209 Bi bismuth 83	[209] Po polonium 84	[210] At astatine 85	[222] Rn radon 86										
[223] Fr francium 87	[226] Ra radium 88	[227] Ac* actinium 89	[267] Rf rutherfordium 104	[270] Db dubnium 105	[269] Sg seaborgium 106	[270] Bh bohrium 107	[270] Hs hassium 108	[278] Mt meitnerium 109	[281] Ds darmstadtium 110	[281] Rg roentgenium 111	[285] Cn copernicium 112	[286] Nh nihonium 113	[289] Fl flerovium 114	[289] Mc moscovium 115	[293] Lv livermorium 116	[293] Ts tennessine 117	[294] Og oganesson 118										

Before scientists knew about protons, electrons and neutrons, **early periodic tables were ordered by atomic weight**

Early periodic tables were incomplete, and some elements were in inappropriate groups if strict order of atomic weight was followed

Mendeleev

He fixed some problems:

- He left gaps for undiscovered elements
- Swapped some elements out of atomic weight order to keep elements with similar properties together

Mendeleev's Table of Elements

H																	B	C	N	O	F
Li	Be															Al	Si	P	S	Cl	
K	Ca	*	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	*	*	As	Se	Br					
Rb	Sr	Y	Zr	Nb	Mo	*	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I					
Cs	Ba	*	*	Ta	W	*	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi							

The columns are called **groups**. Elements in the same group have the **same number of electrons on their outer shell** and have **similar chemical properties** because of this.

Mendeleev's table was accepted when elements were discovered fitted into the gaps he left (he had predicted the properties of these)

- You need to know:
- Group 1 = alkali metals
 - Group 7 = halogens
 - Group 0 = noble gases

Isotopes – when these were discovered (after Chadwick discovered the neutron) it explained why an order based on atomic weight was not always correct

The rows are called **periods**. Elements in the same period have the **same number of electron shells**.

Metals and non-metals

Metals

Elements that react to form positive ions are metals. They are found on the left and towards the bottom of the periodic table.

Properties:

- Ductile (stretch into wires)
- Malleable (can be hammered into shape without breaking)
- Shiny
- High melting and boiling points
- Good conductors of heat and electricity
- Strong

Non-metals

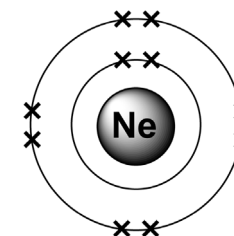
Elements that react to form negative ions are non-metals. They are found towards the right and top of the periodic table.

Properties:

- Weak and brittle (easily break/ shatter)
- Dull
- Poor conductors of heat and electricity
- Generally low melting and boiling points

Group 0 – noble gases

They are unreactive – this is because they have a full outer shell (stable arrangement of electrons)



Boiling point increases down the group, as the atoms get bigger, the intermolecular forces get stronger, more energy is needed to break the forces

Naming Compounds

When a metal reacts with a non-metals you write the name of the metal and change the ending of the non-metal to –ide

e.g., Magnesium + oxygen → magnesium oxide

e.g. Zinc + nitrogen → zinc nitride

When a metal reacts with 2 non-metals you write the name of the metal and change the ending of the non-metal to –ate

e.g., Magnesium + sulfur + oxygen → magnesium sulfate

e.g., Zinc + nitrogen + oxygen → zinc nitrate

Group 7 - halogens

Group 1 – alkali metals

Appearance

Chlorine
Pale green gas



Bromine
Orange-brown liquid



Iodine
Grey solid or purple vapor



Properties

- They are diatomic (they travel around in pairs) e.g., Cl₂ and Br₂
- Melting and boiling point increase as you go down the group
- Reactivity decreases as you go down the group

Physical properties

For metals, the alkali metals have some unusual properties:

- Soft (they can be cut with a knife)
- Relatively low melting points
- Low densities

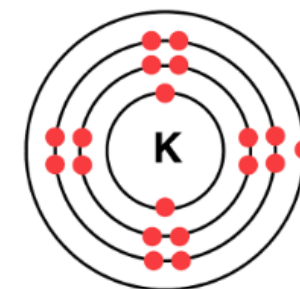
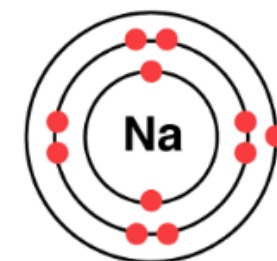
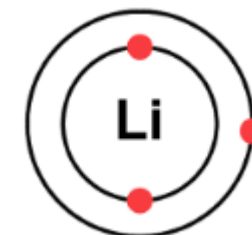
Reactivity

All alkali metals react by losing their outer shell electron

The reactivity of alkali metals increases as you go down the group

This is because:

- Atoms get bigger (there are more electron shells)
- So, the outer shell electron gets further away from the nucleus
- This means the forces (electrostatic) of attraction between the nucleus and the outer shell electron get weaker
- This means it becomes easier to lose the outer shell electron



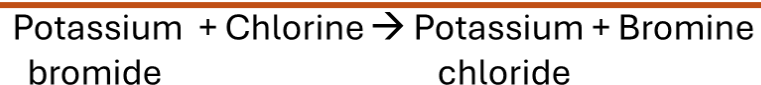
Reactions

- Halogens will react with non-metals to form simple covalent compounds
- Halogens will react with metals to form ionic compound

Displacement Reactions

Halogens undergo displacement reactions.

→ Where more reactive elements replaces a less reactive elements in a compound.



e.g., chlorine
displaces
bromine as it is
more reactive

Reactions of group 1 – alkali metals

Reactions with water

They are called alkali metals because they react with water to form an alkaline solution (release OH⁻ ions)

Alkali metals react with water to form a metal hydroxide and hydrogen

e.g., Sodium + water → sodium hydroxide + hydrogen



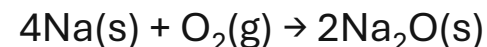
Element	Observations when added to water
Li	Floats, fizzes, moves on surface of water, slowly reacts until it disappears
Na	Floats, fizzes rapidly, moves on surface of water, melts to form a ball and quickly becomes smaller until it disappears
K	Burns violently, violet flame, quickly melts to form a ball and disappears rapidly, often with a small explosion

Reactions with Oxygen

The *alkali metals* burn vigorously when heated and placed in a gas jar of oxygen.

The oxide forms as white smoke.

e.g., sodium + oxygen → sodium oxide



The *reactivity* of the group 1 elements increases down the group, so their reactions with oxygen get more vigorous.

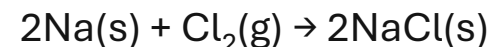
Reactions with chlorine

The alkali metals react vigorously with chlorine. The products of the reactions are chlorides.

At room temperature the chlorides are white solids.

They dissolve in water to form colourless solutions.

e.g., sodium + chlorine → sodium chloride



The reactions with chlorine get more vigorous going down the group.

Transition Metals

You need to be able to compare group 1 and transition metals

Transition metals	Group 1
High melting points	Relatively low melting points
High densities	Low densities
Greater strength	Less strength
Greater hardness	Relatively soft
Conduct electricity and heat	Conduct electricity and heat
Shiny when cut	Shiny when cut
Form colourful compounds	Form white compounds
Form multiple charged ions (e.g. Fe^{2+} , Fe^{3+})	Form +1 ions
They (and their compounds) can be used as catalysts	
Less reactive (slow or no reaction with cold water, oxygen. Some react with halogens)	Really reactive (react vigorously with water and halogens and react quickly with oxygen)

Filtration and Crystallisation 1	Filtration and Crystallisation 1	Distillation	Chromatography	Chromatography R_f values	Check 20	Misconceptions
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Key Words	Misconceptions
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Key Word	Definition
Soluble	Can dissolve in water
Insoluble	Cannot dissolve in water
Solute	Solid that gets dissolved
Solvent	The liquid that does the dissolving
Solution	The solute and solvent mixed
Filtration	Method used to separate an insoluble solid from a solution
Crystallisation	Method used to separate a soluble solid from a solution
Distillation	Method used to separate a mixture containing two substances with different boiling points
Chromatography	Method used to separate e.g., coloured mixtures e.g., ink, food colouring

Salt and water can be separated by filtration. This is incorrect as filtration only works on insoluble solids. Salt is soluble

When a solute is dissolved in solvent, it is no longer there. This is incorrect, the solute breaks into smaller pieces until it can't be seen

Compounds can be separated easily. This is incorrect and the elements are chemically bonded together. Mixtures can be separated as the components are not chemically bonded together.

Key questions

Why should the start line in chromatography be drawn in pencil?

Why can mixtures be separated?

How can sand, salt and water be separated?

How can you calculate a R_f value?

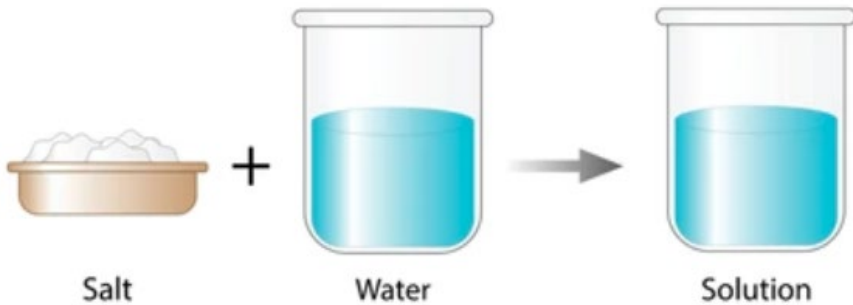
Mixtures

Two or more different elements or compounds, NOT chemically bonded together

As mixtures are not chemically bonded, they can be separated. This is not through a chemical reaction but a physical process.

Solute, solvent and solution

Some substances are soluble – they dissolve in water



Salt = solute (gets dissolved)

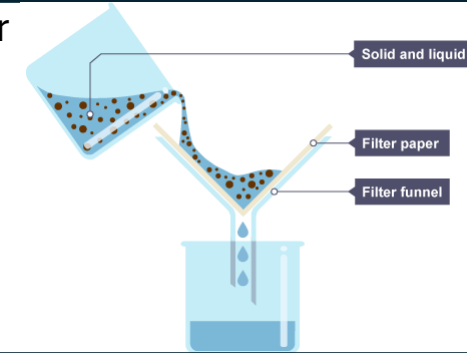
Water = solvent (does the dissolving)

Salty water = solution (mixture of solute and solvent)

Filtration

Separates insoluble solids from a liquid e.g., sand and water

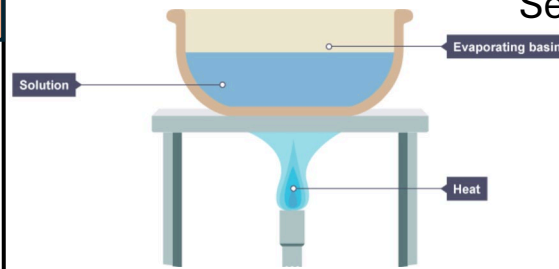
Filter paper traps large particles but lets smaller particles, e.g., water, through.



Crystallisation

Separates soluble solids from a liquid e.g., salt and water

Water evaporates and leaves the solid behind



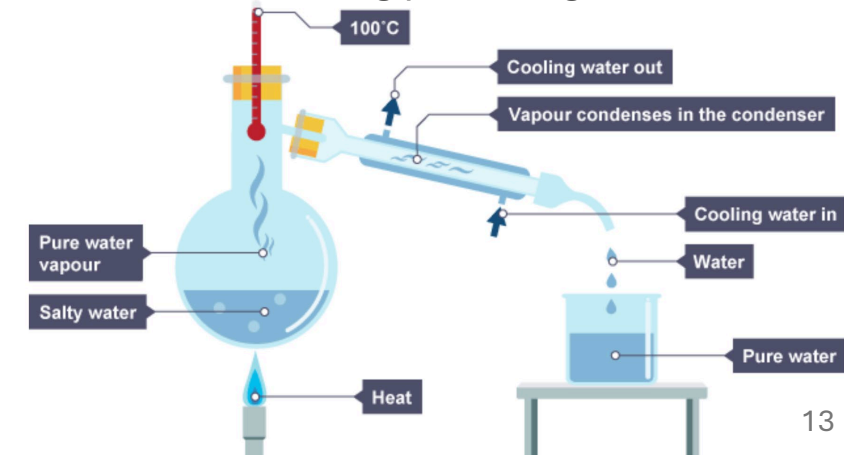
Distillation

Separates mixtures where components have different boiling points e.g., ink and water

Two processes in distillation:

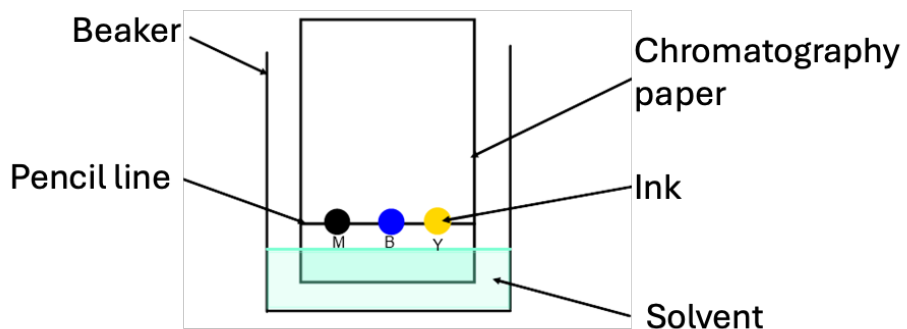
- Evaporation
- Condensation

If there is more than 2 substances, fractional distillation can be used



Chromatography

Method used to separate mixtures and give information to help identify substances



Start line is drawn in pencil as it is insoluble. If it was drawn in pen, it would run.

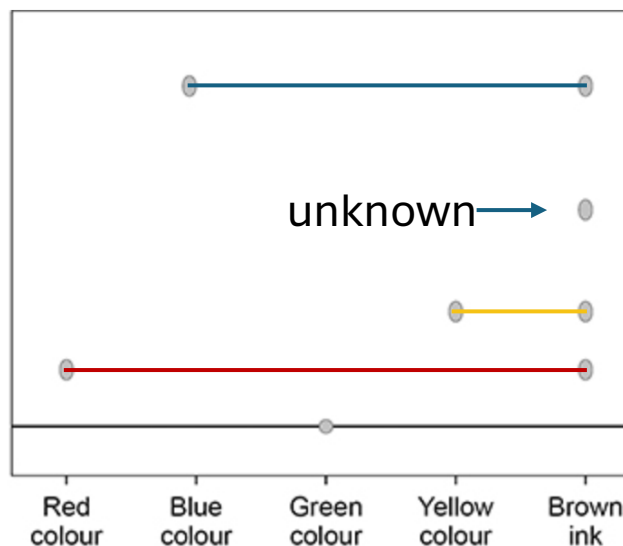
Solvent must be below the start line. This stops the substances washing off into the solvent

Phases of chromatography

Mobile phase – in paper chromatography, this is the solvent. It **moves** and carries the components up the paper

Stationary phase – in paper chromatography, this is the paper. It **doesn't move** and the components move up it.

Chromatography separates substances based on their solubility. The more soluble, the higher up the paper it travels e.g. blue is most soluble, green is insoluble as it doesn't move off the start line



Understanding chromatograms

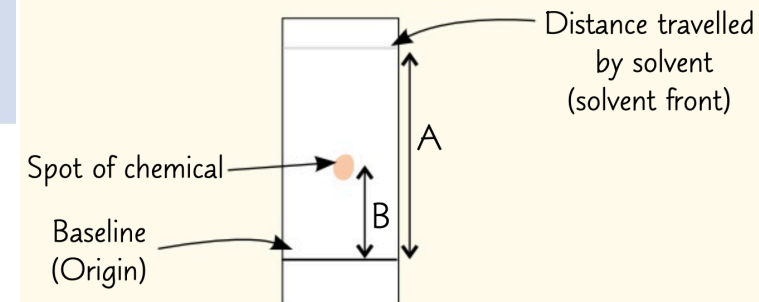
- **1 spot** in a vertical line means it is a **pure** substance e.g., red, blue, yellow
- **More than 1 spot** in a vertical line means a **mixture** e.g., brown ink
- To work out what is in the mixture, look at spots that match e.g., in brown ink there is red, yellow and blue and an unknown

 R_f Values

$$R_f \text{ value} = \frac{\text{distance travelled by substance (B)}}{\text{distance travelled by solvent (A)}}$$

Hints:

- Always measure to middle of the spot
- Answers should be between 0 and 1.
- Check the questions for the units to measure in e.g., mm or cm



Formation of Ions	Ionic Formula 1	Ionic Formula 2	Ionic Bonding	Properties of Ionic Bonding	Covalent Bonding	Properties of Simple Molecules	Polymers	Giant Covalent Structures	Graphene and Fullerenes	Metallic Bonding	Alloys	Nanoparticles	Check 20	Misconceptions
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Key Words

Key Word	Definition
Ion	Atoms that have lost or gained electrons (charged particles)
Ionic bonding	Occurs between a metal and a non-metal
Covalent bonding	Occurs in non-metallic elements and compounds
Metallic Bonding	Occurs in metallic elements and alloys
Monomer	Single molecule
Polymer	Long chain molecules made up of repeating units (monomer)
Alloys	Mixture containing at least one metal
Delocalised electrons	Electrons that are free to move, they are no longer on shells

Misconceptions

When talking about melting and boiling point, we refer to breaking chemical bonds. Wrong – unless it is diamond and graphite, then we refer to breaking forces

Solid ionic substances cannot conduct electricity as they don't have delocalised electrons. Wrong. It is because the ions cannot move and carry a charge.

All covalent substances are giant. Wrong. The only giant covalent structures are diamond, graphite and silicon dioxide. Every other covalent structure is a simple molecule.

Key questions

Why are most covalent structures gas or liquid at room temperature?

Why do ionic substances have high melting points?

Why are alloys harder than pure metals?

Formation of Ions

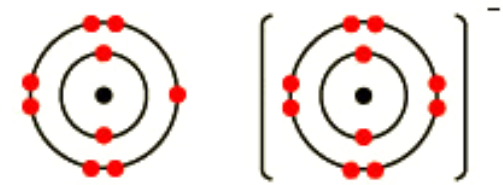
When metals and non-metals react, they lose or gain electrons to get full outer shells of electrons. This makes them ions.



Metals lose electrons to form positively charged ions

lithium atom lithium ion

Non-metals gain electrons to form negatively charged ions

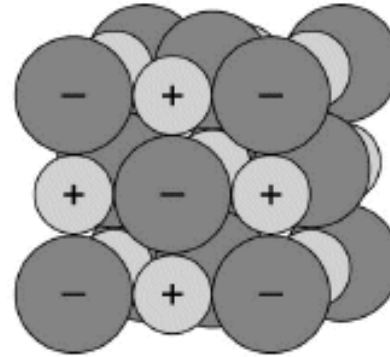


fluorine atom, fluoride ion

Ionic Bonding

Positive and negative ions attract to form a **giant ionic lattice**

The oppositely charged ions are held together by **strong electrostatic forces**

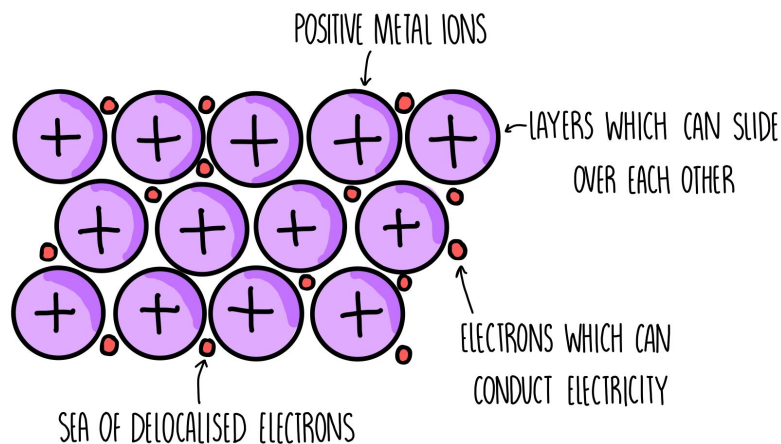


Properties

- High melting and boiling points → giant structure, strong electrostatic forces between oppositely charged ions in all directions which require lots of energy to break
- Conduct electricity when molten or dissolved in water → ions can move and carry a charges

Metallic Bonding

Metals are made of **positive metal ions** and **delocalised electrons**. They have a **giant lattice structure**, held together by **strong electrostatic forces**

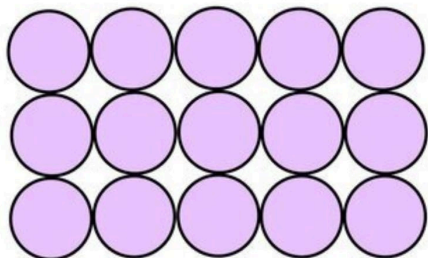


Properties

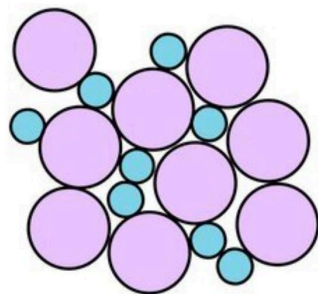
- Malleable → ions are in layers so they can slide
- High melting point → giant lattice structure, held together by strong electrostatic forces, lots of energy needed to break the forces
- Conduct electricity → delocalised electrons which can move through the whole structure and carry a charge
- Conduct heat → delocalised electrons which can move through the whole structure and transfer thermal energy

Alloys

Mixture containing at least one metal



PURE METAL

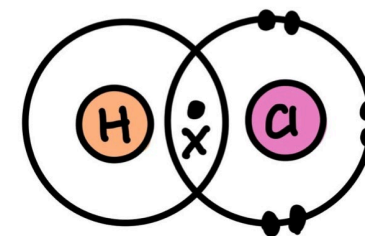


ALLOY

Alloys are harder than pure metals because different sized atoms, distort the layers so they cannot slide

Covalent bonding

When non-metals react with non-metals, they form compounds by sharing electrons, so all the atoms involved have full outer shells.



A covalent bond is a shared pair of electrons

Strong covalent bonds are drawn as — in diagrams

Polymers

Long chain molecules made up of repeating units called monomers



ETHENE

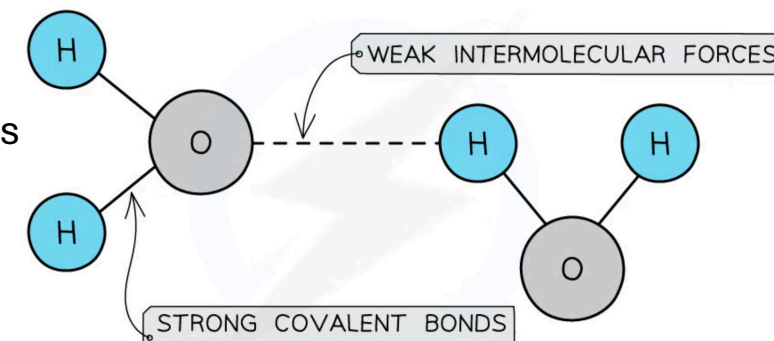
POLY(ETHENE)

Properties

- Solid at room temperature → **strong intermolecular forces** between the polymer chains, lots of energy is needed to overcome these forces (strong covalent bonds between atoms are NOT broken)

Simple molecules

Most covalently bonded substances are simple molecules. They are gases or liquids at room temperature



Properties

- Low melting and boiling points → simple molecules, held together by **weak intermolecular forces** which require little energy to break (the strong covalent bonds between atoms are NOT broken)
- Do not conduct electricity → no delocalised electrons

Diamond

Is a giant covalent structure made of carbon.



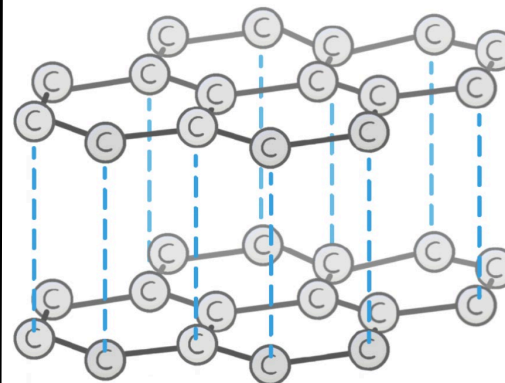
Properties

- High melting point → **Each carbon makes 4 strong covalent bonds**, lots of energy is needed to break the bonds
- Strong → each carbon makes 4 strong covalent bonds)
- Do not conduct electricity → no delocalised electrons

Silicon dioxide has a similar structure and properties to diamond, but is made of silicon and oxygen

Graphite

Is a giant covalent structure made of carbon.



Properties

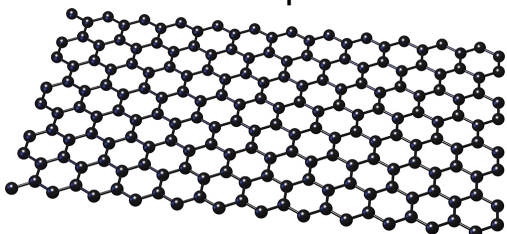
- High melting point → **Each carbon makes 3 strong covalent bonds**, lots of energy is needed to break the bonds
- Soft → Atoms are arranged in layers, the layers are held together by weak intermolecular forces so they can slide
- Conducts electricity → Each carbon makes 3 strong covalent bonds, so there is 1 delocalised electron per carbon atom which can move through the whole structure and carry a charge

Graphene

Graphene – a single layer of graphite

Properties: high melting point, strong and conduct electricity

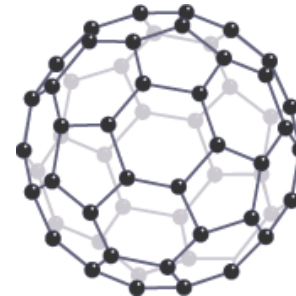
Uses: electronics and making composites



Fullerenes

Structures based on hexagonal rings of carbon

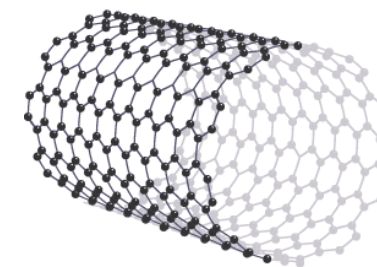
Buckminster fullerene: Spherical molecule made of 60 carbon atoms



Uses:

Lubricant → spherical so will roll
Medicinal drug delivery → hollow so things can be put inside

Nanotubes: Tubes made of carbon



Uses:

Rackets → high tensile strength
Electronics → contains delocalised electrons so can conduct electricity

Nanoparticles

Nanoparticles are structures with a size between 1-100nm. Most nanoparticles are made up of a few hundred atoms.

Properties of nanoparticles can vary to the same material in bulk due to:

- The size of nanoparticles compared to the same material in bulk
- The large SA:V ratio of nanoparticles

Uses:

- medical treatments
- cosmetics, deodorants and sunscreens
- electronics
- catalysts

Example – Zinc oxide blocks ultraviolet light, so is used in sunscreens. Bulk zinc oxide is white, but nanoparticles of zinc oxide are invisible on the skin. Due to the larger SA:V, less is needed to be effective therefore it can also save money.



Issues with nanoparticles

The small size of nanoparticles could make it possible to breathe them in, or for them to pass into cells. Once inside the body, they might catalyse reactions that are harmful. *Toxic* substances could bind to them because of their large surface area to volume ratios. Modern nanoparticulate materials have only become common recently, so it is difficult for scientists to determine their risks.

Surface area : Volume ratio

Surface area = (w x h) x number of sides

Volume = w x l x h

	
<p>Surface area</p> <p>1 side = $10 \times 10 = 100 \text{ cm}^2$ 6 sides = $6 \times 100 = 600 \text{ cm}^2$</p>	<p>Surface area</p> <p>1 side = $1 \times 1 = 1 \text{ cm}^2$ 6 sides = $6 \times 1 = 6 \text{ cm}^2$</p>
<p>Volume</p> <p>Volume = $10 \times 10 \times 10 = 1000 \text{ cm}^3$</p>	<p>Volume</p> <p>volume = $1 \times 1 \times 1 = 1 \text{ cm}^3$</p>
<p>Surface area : Volume ratio</p> <p>600:1000 6:10 0.6:1</p>	<p>Surface area : Volume ratio</p> <p>6:1</p>

As the side of cube decreases by a factor of 10 the surface area to volume ratio increases by a factor of 10.

Ionic Formula

Compounds have no overall charge. When writing ionic formula, you need to have the same number of positive and negative ions

Example 1: Sodium chloride

Sodium = Na^+

Chlorine = Cl^-

The charges cancel out

NaCl

Example 2: Magnesium chloride

Magnesium = Mg^{+2}

Chlorine = Cl^-

As the Mg is +2 we need x2 of the Cl ions to cancel this out

MgCl_2

Example 3: Magnesium nitrate

Magnesium = Mg^{+2}

Nitrate = NO_3^-

As the Mg is +2 we need x2 NO_3 but because it contains more than one element it will need brackets

$\text{Mg}(\text{NO}_3)_2$



Balancing Equations	Conservation of mass 1	Conservation of mass 2	RFM	Percentage by mass	Concentration	Moles	Balancing Equations using Moles	Reacting Masses 1	Reacting Masses 2	Limiting Reactants	Check 20	Misconceptions lesson
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Key Words						Misconceptions					
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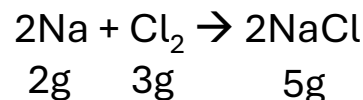
Key Word	Definition
Conservation of mass	No atoms are lost or gained during a chemical reaction
RFM	Total mass of all of the atoms in a compound
Concentration	Number of atoms in a given volume

Key questions			
In RFM, bracket. The numbers are multiplied not added. E.g. $\text{Ca}(\text{NO}_3)_2$ the O is $3 \times 2 = 6$	When calculating % by mass you must work out the total mass of the element e.g. CO_2 for O you would do $16 \times 2 = 32$	When balancing equations, you can only add big numbers to the front. Adding little numbers changes the chemicals	For concentration, volume must be converted into dm^3 . cm^3 to $\text{dm}^3 = \div$ by 1000

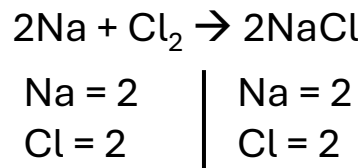
What is conservation of mass?	When might reactions not show conservation of mass?	How do you calculate RFM?	How do you calculate percentage by mass?
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Conservation of Mass

No atoms are lost or gained during chemical reactions. This means the mass of the reactants should equal the mass of the products.

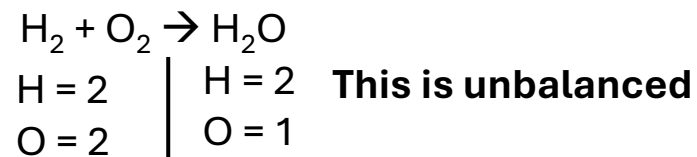


Balanced symbol equations show conservation of mass. There is the same number of atoms on each side.



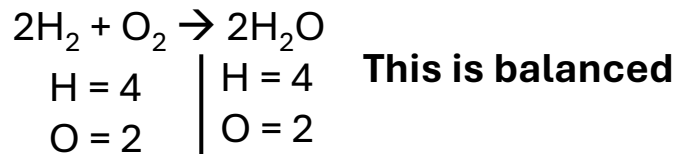
Balancing Equations

There should be the same number of atoms, of each element on each side of the equation



We can balance equations by putting big numbers at the front of the molecules to balance them.

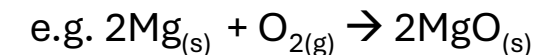
Hint: if in doubt
put a 2 at the
front



Reactions and conservation of mass

Conservation of mass is not always shown in some chemical reactions

1. Where one of the reactants is a gas.



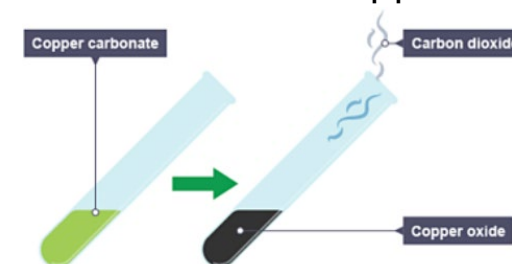
As the oxygen is a gas you cannot measure the mass of it at the beginning, but if you measure the mass of the Mg then the mass of the MgO, the difference will be the mass of oxygen. The mass of this reaction would appear to increase.



2. When one of the products is a gas



As CO_2 is a gas, it can escape the container. This means the mass of this reaction would appear to decrease



RFM

Total mass of all of the atoms in a compound

Steps:

1. Work out how many atoms of each element
2. Multiply these by the mass numbers from the periodic table
3. Add them together

e.g. MgCl_2

$$\text{Mg} = 1 \times 24 = 24$$

$$\text{Cl} = 2 \times 35.5 = 71$$

$$24 + 71 = 95$$

e.g. $\text{Ca}(\text{NO}_3)_2$

$$\text{Ca} = 1 \times 40 = 40$$

$$\text{N} = 2 \times 14 = 28$$

$$\text{O} = 6 \times 16 = 96$$

$$40 + 28 + 96 = 164$$

Percentage by Mass

This is the percentage of mass that makes up a compound

It is calculated using the equation below

$$\% \text{ by mass} = \frac{\text{Total Mass of element in compound}}{\text{RFM}} \times 100$$

Steps:

1. Work out the RFM
2. Put numbers into equation

e.g. Calculate the % by mass of **Cl** in **MgCl_2**

MgCl_2

$$\text{Mg} = 1 \times 24 = 24$$

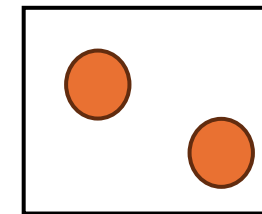
$$\text{Cl} = 2 \times 35.5 = 71$$

$$24 + 71 = 95$$

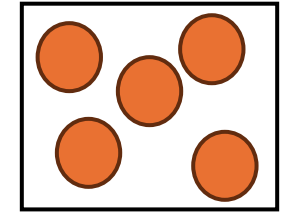
$$\% \text{ by mass} = \frac{71}{95} \times 100 = 74.7\%$$

Concentration

It is the number of particles in a given volume



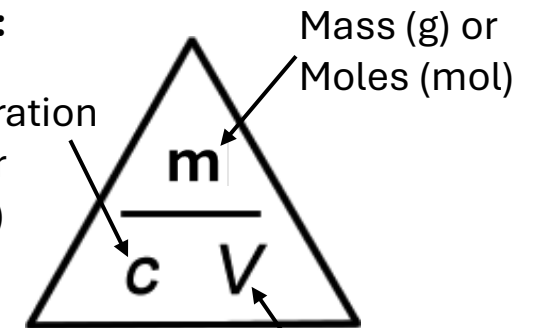
Low concentration



High concentration

Equation:

Concentration
(g/dm^3 or
 mol/dm^3)



Volume (dm^3)

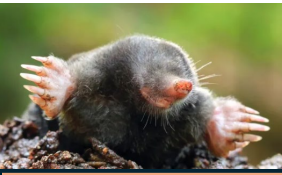
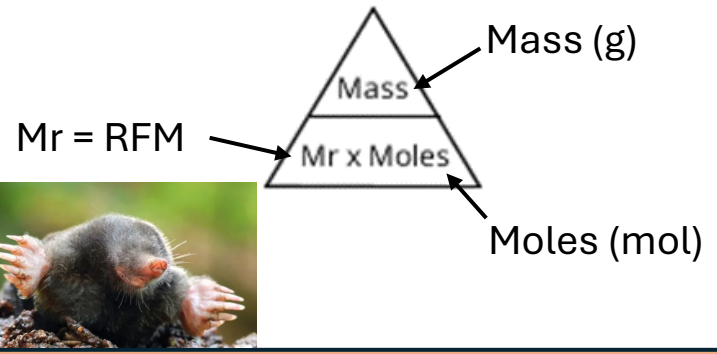
cm^3 to $\text{dm}^3 = \div$ by 1000

To convert from g/dm^3 to $\text{mol/dm}^3 \div$ by Mr

To convert from mol/dm^3 to $\text{g/dm}^3 \times$ by Mr

Moles

A chemical unit



Avogadro's Constant

$$6.02 \times 10^{23}$$

This is the number of atoms, molecules or ions in 1 mole of any substance

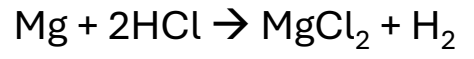
e.g. to work out how many molecules of O_2 in 2 moles you would do:
 $2 \times 6.02 \times 10^{23} =$



Reacting Masses

Working out the mass of a reactant/ product when given a mass

e.g. What mass of hydrogen is made when 20g of HCl reacts with hydrochloric acid?



Sort the information into the table

	Mass	Mr	Moles
HCl	20	36.5	
H ₂		2	

Check ratio of moles using equation (moles of H₂ is half of HCl)

	Mass	Mr	Moles
HCl	20	36.5	0.548
H ₂		2	0.274

Leave number on calculator during each step

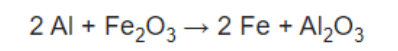
	Mass	Mr	Moles
HCl	20	36.5	0.548
H ₂	0.548g	2	0.274

Limiting Reactant

The reactant which limits the amount of a product you can make

A mixture contains 1.00 kg of aluminium and 3.00 kg of iron oxide.

The equation for the reaction is:



Show that aluminium is the limiting reactant.

Relative atomic masses (A_r): O = 16 Al = 27 Fe = 56

Steps:

1. Convert any masses in Kg to g
2. Calculate moles of each reactant
3. Work out how many moles of Fe₂O₃ you would need to use all the moles of Al
4. Work out how many moles of Al you would need to use all the moles of Fe₂O₃
5. Which one doesn't work (e.g. which ones don't you have enough moles of)?
6. That is your limiting reactant.

Atom Economy	Percentage Yield	Volume of Gases	Titration Method	Titration Calculations 1	Titration Calculations 2	Titration compulsory practical	Check 20	Misconceptions
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Key Words

Key Word	Definition
Atom Economy	Percentage of the starting atoms that become useful products
Percentage Yield	Calculates how what percentage of the theoretical yield is actually made
Concordant	Within 0.1cm ³

Misconceptions

Universal indicator cannot be used in titrations because the colour changes are not distinct enough

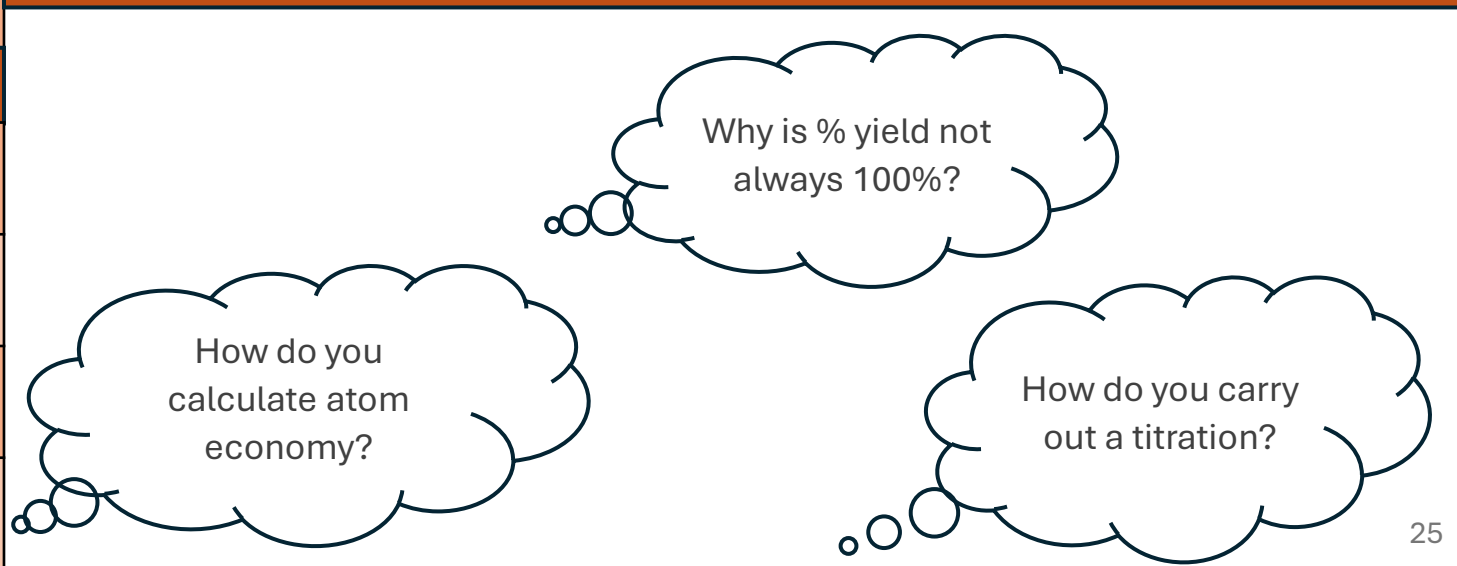
When calculating % atom economy, the big numbers in front of the molecules need to be factored into the RFM

Volumes must be converted from cm³ to dm³. To do this divide by 1000

Equations

Atom Economy	$= \frac{\text{RFM of desired product}}{\text{Total RFM of reactants}} \times 100$
Percentage Yield	$= \frac{\text{Actual Mass}}{\text{Theoretical mass}} \times 100$
Volumes of Gas	$= \text{moles of gas} \times 24$
Concentration	$= \text{mass or moles} \div \text{volume}$

Key questions



Why is % yield not always 100%?

How do you calculate atom economy?

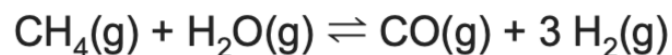
How do you carry out a titration?

Atom Economy

Atom economy measure the percentage of atoms in a reaction that become the useful product

$$\frac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula masses of all reactants from equation}} \times 100$$

It is important when doing this calculation, to include the big numbers when working out the RFMs



Example: % atom economy for the formation of hydrogen

$$\text{RFM of desired: } 6 \times 1 = 6$$

$$\text{RFM of reactants: } 16 + 18 = 34$$

$$(6 \div 34) \times 100 = 17.6\%$$

Equations where there is only one product, show 100% atom economy

High atom economies are important in industry because:

- Reduces waste
- More sustainable
- Economic reasons
- To produce a higher percentage of useful products

Percentage Yield

The amount of a product obtained is known as the yield. When compared with the maximum theoretical amount as a percentage, it is called the percentage yield.

$$\% \text{ Yield} = \frac{\text{Mass of product actually made}}{\text{Maximum theoretical mass of product}} \times 100$$

Example:

- (d) 60.0 kg of aluminium oxide produces a maximum of 31.8 kg of aluminium.

In an extraction process only 28.4 kg of aluminium is produced from 60.0 kg of aluminium oxide.

Calculate the percentage yield.

$$\frac{28.4}{31.8} \times 100 = 89.3\%$$

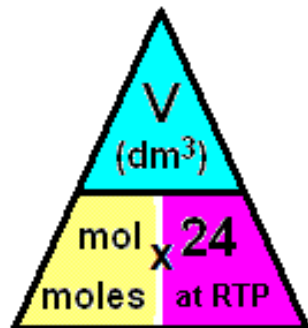
Reasons % yield may not be 100%

- the reaction may not go to completion because it is reversible
- some of the product may be lost when it is separated from the reaction mixture
- Unexpected reactions

Volumes of Gases

One mole of any gas at room temperature and pressure occupies a volume of 24dm^3

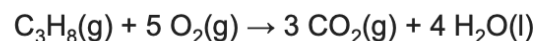
We can use the following equation to calculate the volume of a gas when given the moles



Lots of questions about volumes of gas, don't rely on using the equation above, but rely on you understanding balanced symbol equations showing you ratios (look at big numbers)

Example:

The equation for the reaction is:



3.60 dm^3 carbon dioxide is produced when a sample of propane is burned in 7.25 dm^3 oxygen.

Calculate the volume of unreacted oxygen.

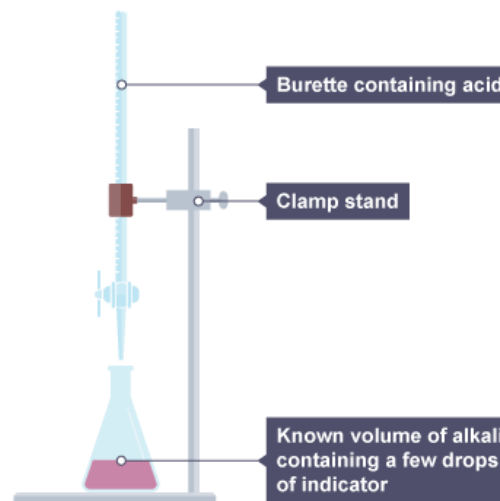
Ratio of $\text{CO}_2:\text{O}_2 = 3:5$

Volume of oxygen needed = $(3.6 \div 3) \times 5 = 6\text{dm}^3$

Excess oxygen = $7.25 - 6 = 1.25\text{dm}^3$

Titration Practical

The volumes of acid and alkali solutions that react with each other can be measured by titration using a suitable indicator.



Indicators

Phenolphthalein – **Pink in alkalis** and **colourless in acids**

Methyl orange – **yellow in alkalis** and **red in acids**

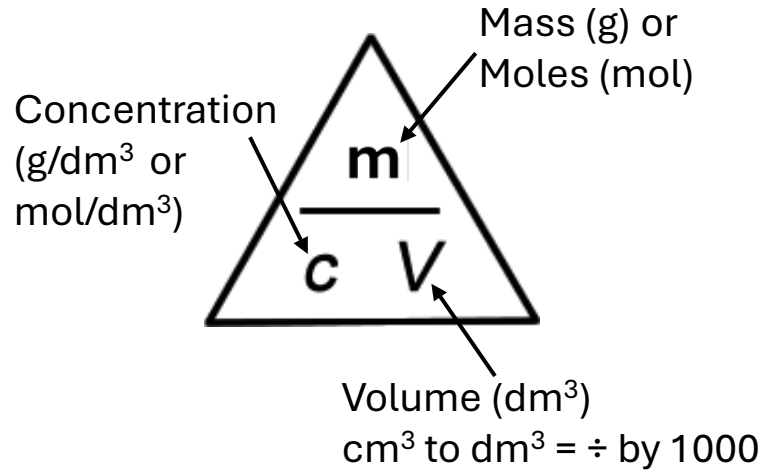
Cannot use universal indicator as the colour change is not distinct enough

A student does a titration to find the volume of sulfuric acid that reacts with 25 cm^3 of potassium hydroxide solution.

1. Measure out 25cm^3 of potassium hydroxide using a pipette and add to a conical flask
2. Add 4 drops of an indicator to the flask and place in a white tile
3. Fill the burette with sulfuric acid
4. Add the acid to the flask dropwise, swirling the flask constantly
5. Stop the burette then the indicator changes colour and record the volume
6. Repeat the titration until concordant results are collected (0.1cm^3)

Titration Calculations

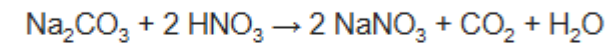
These calculations use the equation:



To convert from g/dm³ to mol/dm³ ÷ by Mr
 To convert from mol/dm³ to g/dm³ x by Mr

A student does a titration using sodium carbonate solution and nitric acid.

The equation for the reaction is:



25.0 cm³ of 0.124 mol/dm³ sodium carbonate solution is neutralised by 23.6 cm³ of nitric acid.

Calculate the concentration of the nitric acid.

Add in information from the question first

	M	C	V
Na ₂ CO ₃	M = 0.124 x 0.025 = 0.0031mol	0.124mol/dm ³	25cm ³ 0.025dm ³
2HNO ₃	M = 0.0031 x 2 = 0.0062mol	C = 0.0062 ÷ 0.0236 = 0.263mol/dm ³	23.6cm ³ 0.0236dm ³

Note: Blue arrows in the original image indicate the flow of information: 1. Convert volumes (25 cm³ to 0.025 dm³); 2. Calculate moles of Na₂CO₃ (0.124 mol/dm³ x 0.025 dm³); 3. Calculate moles of 2HNO₃ (0.0031 mol x 2); 4. Calculate concentration of HNO₃ (0.0062 mol ÷ 0.0236 dm³).

Neutralisation and pH	Naming Salts	Making Salts 1	Making Salts 2	Reactivity of Metals	Displacement Reactions	Extraction of Metals	Ionic Equations	Check 20	Misconceptions lesson
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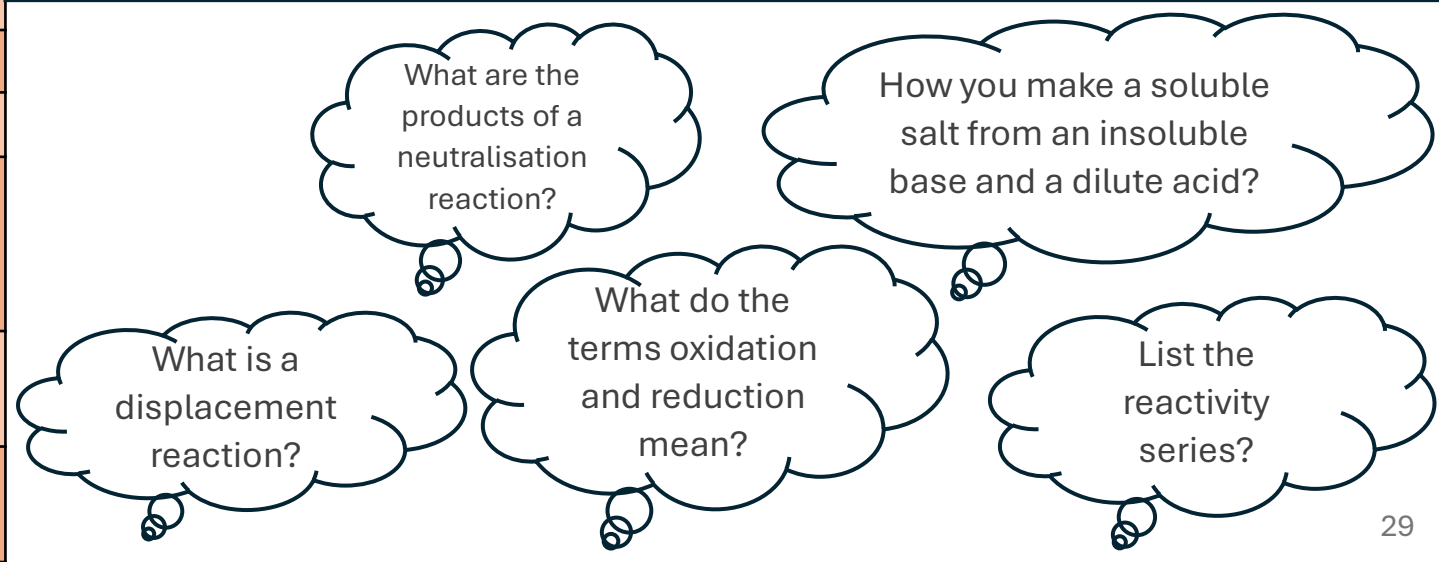
Key Words

Key Word	Definition
Acid	Ionise in water to release H ⁺ ions
Alkali	Ionise in water to release OH ⁻ ions
Neutralisation	When an acid and alkali react to form a salt and water
Soluble	Can dissolve
Insoluble	Cannot dissolve
Insoluble base	Metal oxide or Metal carbonate
Displacement reaction	When a more reactive element replaces a less reactive element in a compound
Oxidation	Gaining oxygen
Reduction	Losing/ removing oxygen

Misconceptions

<p>Litmus paper cannot be used to show the pH of a substance. The most accurate way to find pH is a pH probe.</p>	<p>When a metal reacts with acid. Hydrogen gas is produced. MASH</p>	<p>Not all metals are extracted using mining. Ores (compound containing metal) are mined, but most metals must be extracted from the ore.</p>	<p>Carbon is not a metal. It is part of the reactivity series as it is used to extract some metals.</p>
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Key questions



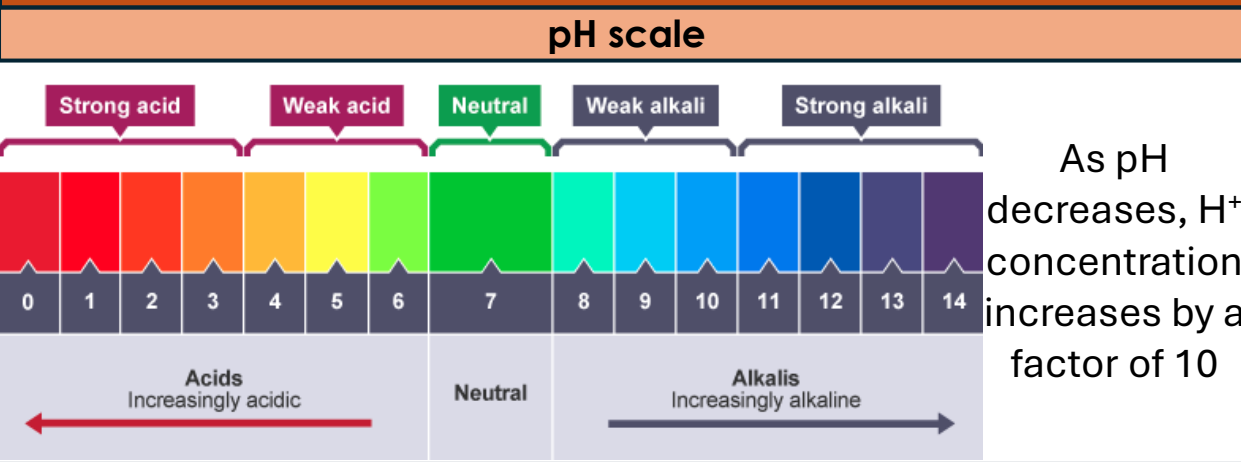
What are the products of a neutralisation reaction?

How you make a soluble salt from an insoluble base and a dilute acid?

What is a displacement reaction?

What do the terms oxidation and reduction mean?

List the reactivity series?



Acids, bases and alkalis

- **Acids** ionise in water to release **H⁺ ions**
- A base is a substance that reacts with acids in neutralisation reactions. An alkali is a soluble base.
- **Alkalis** ionise in water to release **OH⁻ ions**

Strong and Weak Acids

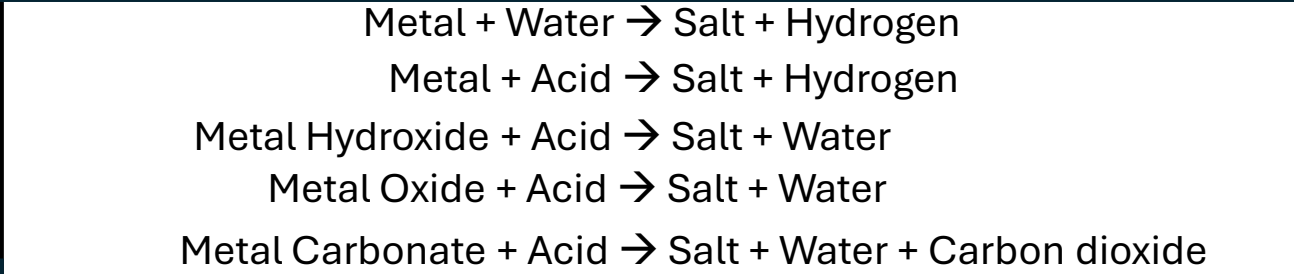
Strong acids fully ionise in water to release H⁺

Weak acids partially ionise in water to release H⁺

Measuring pH

- Universal indicator**
- This changes colour in solution which can then be matched to the pH scale to show the pH of the solution
- pH probes**
- This most accurate way to measure pH as it gives you a number

Reactions of metals



Neutralisation

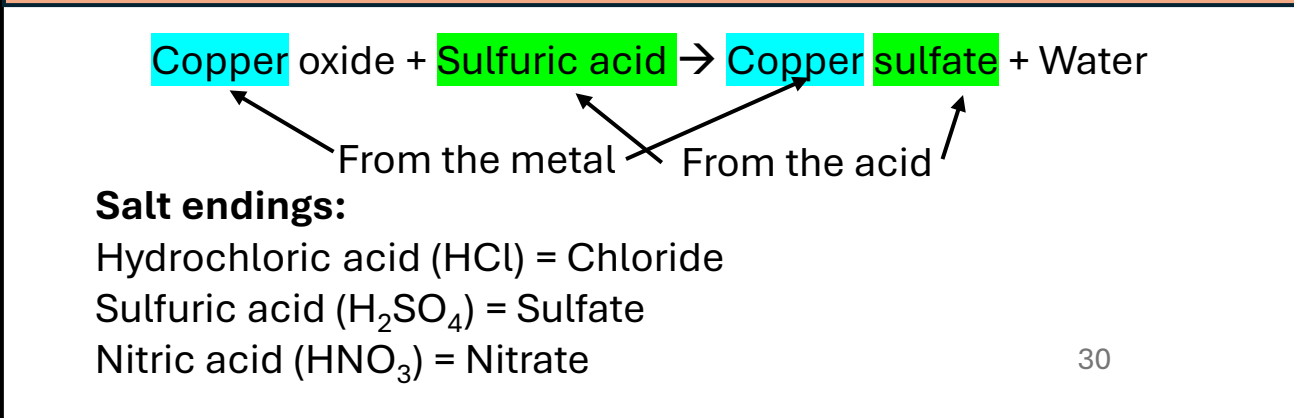
This is when an acid and alkali react

Acid + Alkali → Salt + Water

The water is made by hydrogen ions from acids reacting with hydroxide ions from alkalis.

Shown by the following **ionic equation**: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

Naming Salts



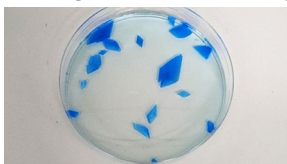
Making Salts

Making a pure dry sample of a soluble salt from an insoluble base and a dilute acid.

Hint: look at the question to work out the chemicals

Method for making pure, dry sample of copper sulfate

1. Wear goggles because acid is corrosive
2. Heat sulfuric acid using a water bath (to speed up the reaction)
3. Add copper oxide until it is in excess (to make sure all the acid has reacted)
4. Filter out the excess copper oxide using a funnel and filter paper
5. Heat remaining solution in an evaporating basin over a water bath until volume has halved
6. Pour solution into a crystallising dish and leave to cool and crystallise
7. Pat crystals dry with filter paper



Reactivity Series and extracting metals

Please - Potassium

Stop - Sodium

Liam - Lithium

Calling - Calcium

Me - Magnesium

A - Aluminium

Cute - Carbon

Zebra - Zinc

I - Iron

Like - Lead

Hilary - Hydrogen

Clinton - Copper

She's - Silver

Great - Gold

Extracted using electrolysis because they are more reactive than carbon

Extracted using carbon reduction. Carbon removes oxygen from the ore (reduction)

e.g.

Iron oxide + carbon → carbon dioxide + iron

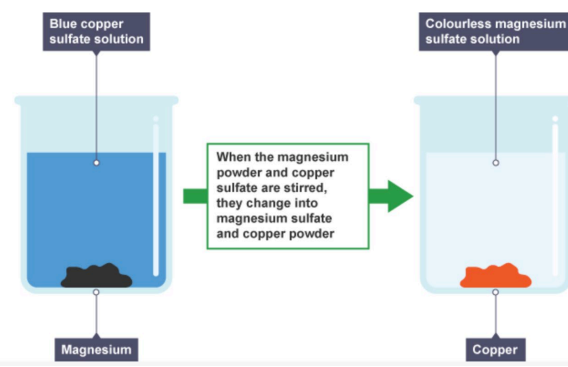
Iron is reduced – loses oxygen

Carbon is oxidised – gains oxygen

Native metals – found in the ground as pure metals

Displacement Reactions

When a more reactive element replaces a less reactive element in a compound e.g., Copper sulfate + magnesium → magnesium sulfate + copper



Magnesium displaces the copper as it is more reactive

Observations (what you can see)

- Copper metal at the bottom
- Colour change from blue to colourless

Oxidation and reduction - electrons

Oxidation and reduction can happen in terms of electrons

O xidation	R eduction
I s	I s
L oss	G ain

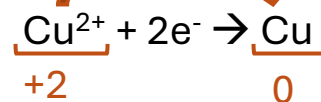
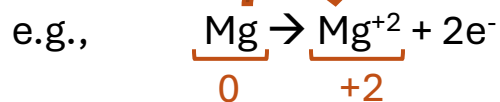
Half-equations

We can turn ionic equations into two half-equations

Half-equations show oxidation and reduction

Lost electrons = oxidation

Gained electrons = reduction



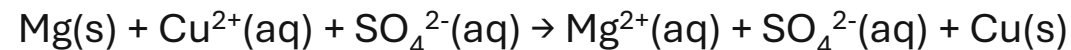
Always put the electrons on the most positively charged side. **Elements have no charge**

Ionic Equations

Ionic equations can be written for displacement reactions. It shows what happens in terms of ions

e.g., Magnesium + copper sulfate → magnesium sulfate + copper

Can be written as:



Sulfate ions, SO_4^{2-} , appear on both sides of the equation, but they do not take part in the reaction (spectator ions), so we remove them to leave us with:



Redox Reactions

This is where oxidation and reduction happen in the same reaction. Displacement reactions are redox reactions. A species is always oxidised and the other reduced

Process of Electrolysis	Half Equations	Electrolysis of Molten Ionic Compounds	Electrolysis of Aqueous Compounds	Electrolysis Compulsory Practical	Electrolysis to Extract Metals	Check 20	Misconceptions lesson
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Key Words

Key questions

Key Word	Definition
Oxidation	Losing Electrons
Reduction	Gaining Electrons
Electrolysis	Breaking down ionic compounds using and electric current
Electrolyte	Molten or aqueous ionic compound
Anode	Positive electrode
Cathode	Negative electrode
PANIC	Positive anode, negative is cathode
Aqueous	Dissolves in water
Ions	Charged particles formed when atoms have lost or gained electrons
Discharged	When something loses its charge and becomes neutral

What is an electrolyte?

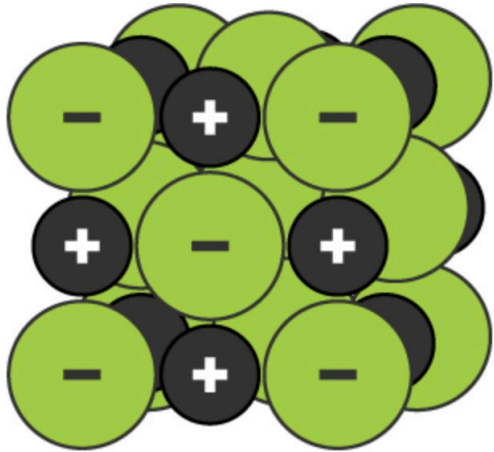
What does PANIC stand for?

Why does the + electrode have to be continually replaced during the extraction of aluminium?

What ions are in water?

What are the rules for aqueous electrolysis?

Ionic Compounds - Recap



Ionic compounds are made of positive metal ions (lost electrons) and negative non-metal ions (gained electrons), in a giant ionic lattice structure held together by strong electrostatic forces

Ionic compounds have high melting points because:

- They have a giant ionic lattice structure
- Oppositely charged ions are held together by strong electrostatic forces
- Lots of energy is needed to break the forces

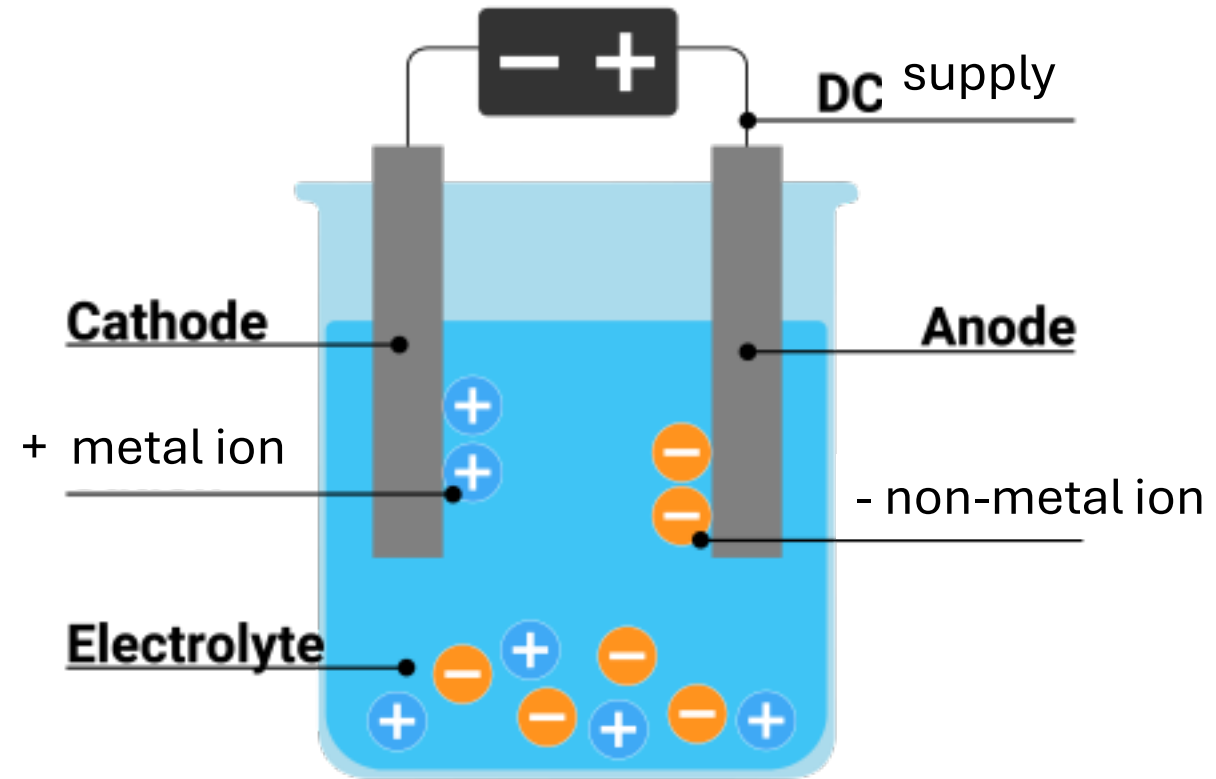
Conductivity

As a solid, ionic compounds cannot conduct electricity as the ions cannot move and carry a charge.

So, it can conduct electricity we must either melt it (make it molten) or dissolve it in water (make it aqueous)

Process of Electrolysis

PANIC – Positive Anode, Negative is Cathode



Key thing to remember; opposite charges attract

+ metal ions are attracted to the negative electrode (cathode) here they are discharged by gaining electrons (reduction)

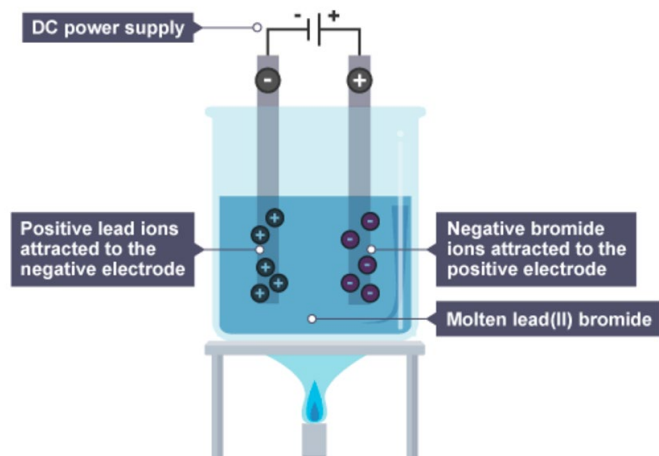
- non-metal ions are attracted to the positive electrode (anode), here they are discharged by losing electrons (oxidation)

Molten Electrolysis

Positive metal ions go to the negative electrode as opposite charges attract. Here they gain electrons to become atoms (reduced)

Negative non-metal ions go to the positive electrode as opposite charge attract. Here they lose electrons to become atoms (oxidised)

Example: Molten Lead Bromide



Pb^{2+} ions go to the negative electrode where they gain two electrons (reduced), to become lead atoms



Br^- ions go to the positive electrode where they lose electrons (oxidised) to become Br atoms, which pair up to form Br_2 molecules



Rules for Aqueous electrolysis

H^+ ions and OH^- ions from the water are also present

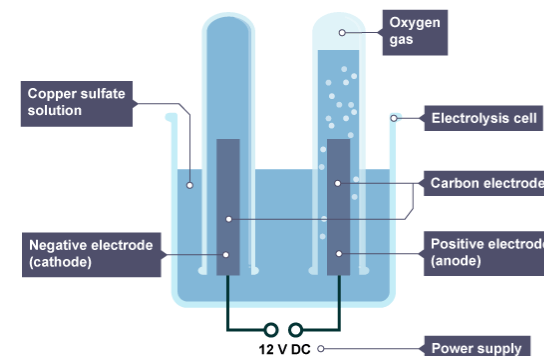
At the + electrode (anode):

- If a halogen is present (Cl^- , Br^- , I^-) then this is made
- If no halogen is present O_2 is made instead)

At the - electrode (cathode):

- The least reactive thing is made

Example: Aqueous Copper Sulfate



During this the solution would become less blue, and copper metal would be seen at the - electrode

At the + electrode (anode):

OH^- and SO_4^{2-} ions go to the + electrode

Oxygen is made: $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$ (oxidation)

At the - electrode (cathode):

H^+ and Cu^{2+} ions go to the - electrode

Copper is made: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (reduction)

Oxidation and reduction - electrons

Oxidation and reduction can happen in terms of electrons

Oxidation	Reduction
Is	Is
Loss	Gain

Half-equations

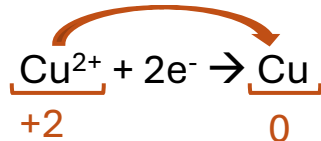
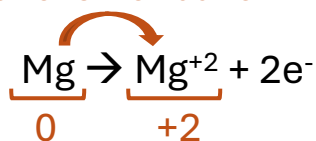
We can turn ionic equations into two half-equations

Half-equations show oxidation and reduction

Lost electrons = oxidation

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e.g.,

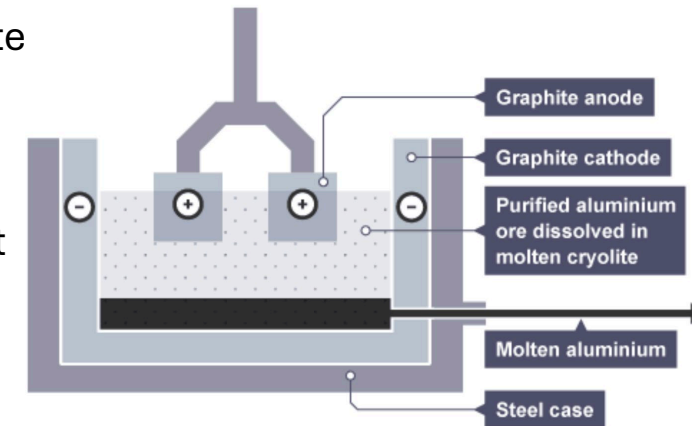


Always put the electrons on the most positively charged side. **Elements have no charge**

Extracting Aluminium

Aluminium oxide is extracted from bauxite

Aluminium oxide is dissolved in cryolite. Cryolite has a lower melting point than aluminium oxide, so reduces the amount energy needed, and the cost



At the - electrode (cathode):

Al^{3+} ions go to the - electrode

Aluminium is made: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$

At the + electrode (anode):

O^{2-} ions go to the + electrode

Oxygen is made: $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$

Oxygen made at the + electrode reacts with the carbon of the electrode to make CO_2 . This means electrode wears away and has to be continually replaced

Exothermic and Endothermic Reactions	Energy Changes Practical 1	Energy Changes Practical 2	Reaction Profiles	Bond Energy Calculations 1	Bond Energy Calculations 2	Cells and Batteries	Fuel Cells	Check 20	Misconceptions
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Key Words	Misconceptions
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Key Word	Definition
Exothermic reaction	Transfers energy to the surroundings so the temperature of the surroundings increases
Endothermic	Reaction that takes in energy from the surroundings so the temperature of the surroundings decreases
Activation energy	Minimum amount of energy the particles must have to react
Overall Energy Change	Difference in energy between the reactants and products
Catalyst	Increase the rate of reaction by providing an alternative pathways for the reaction that has a lower activation energy
Neutralisation	Reaction between an acid and an alkali/base to make a salt and water

Radiators are not an example of an exothermic reaction. They contain hot water, the thermal energy is transferred to the surroundings through convection

When drawing a reaction profile, the activation energy is from the reactants level to the peak of the curve.

Thermal decomposition is an example of an endothermic reaction. Thermal energy is taken IN to the reaction

Key questions

Give two examples of endothermic reactions

What is an exothermic reaction?

What is activation energy?

Exothermic Reactions

EXO = EXITS

An exothermic reaction is one that transfers energy to the surroundings, so the temperature of the surroundings increases

Examples of exothermic reactions: combustion, oxidation reactions and neutralisation

Everyday uses of exothermic reactions: self-heating cans and hand warmers



Exothermic Reactions

ENDO = ENTERS

An endothermic reaction is one that takes in energy from the surroundings, so the temperature of the surroundings decreases

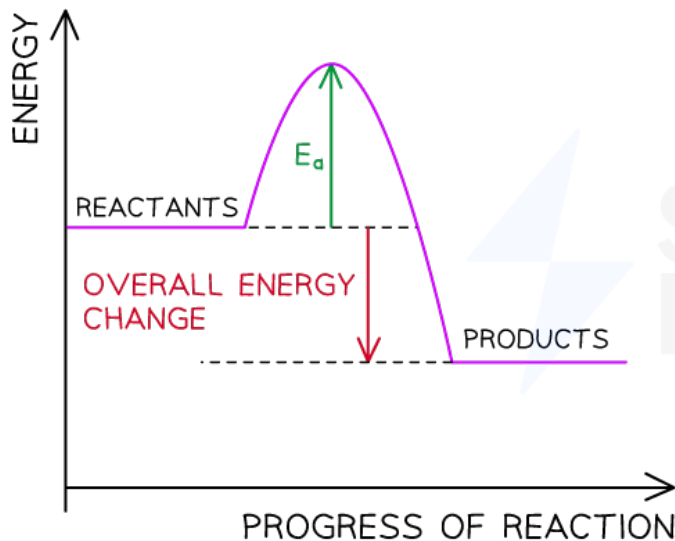
Examples of endothermic reactions: thermal decomposition, reaction of citric acid and sodium hydrogencarbonate

Everyday uses of endothermic reactions: sports injury packs



Reaction Profiles

EXOTHERMIC

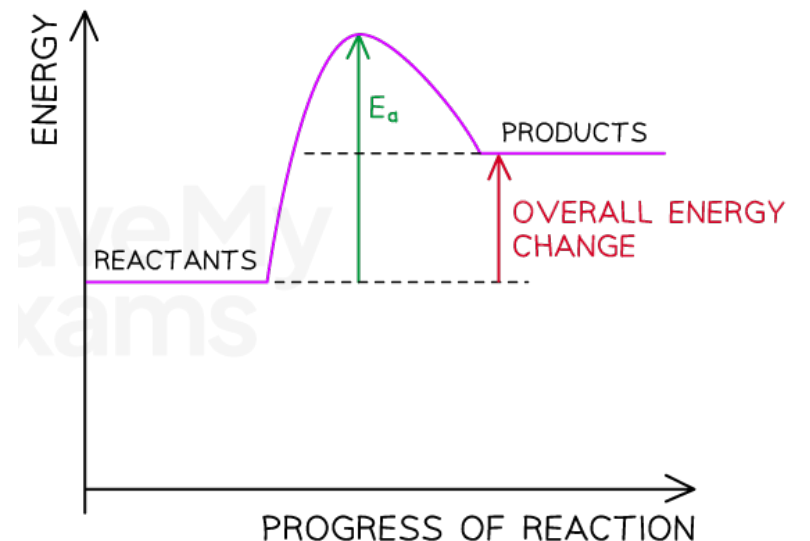


E_a = activation energy

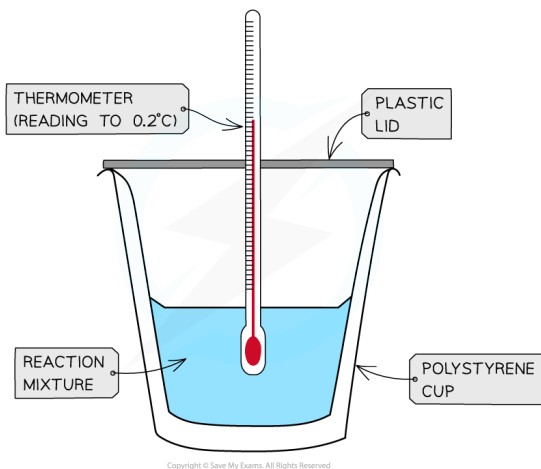
This is the minimum amount of energy the particles must have to react

Overall energy change is the difference in energy between the reactants and products

ENDOTHERMIC



Energy Changes Practical

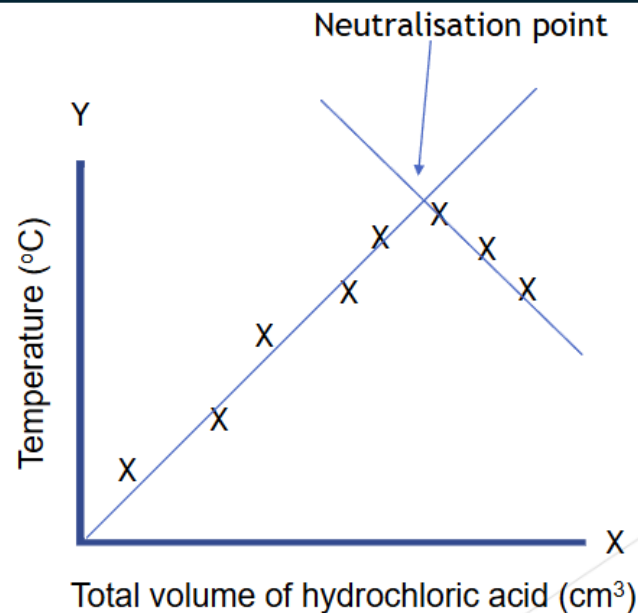


Neutralisation reaction
between sodium hydroxide
and hydrochloric acid

1. Add 25cm³ of sodium hydroxide into a polystyrene cup
2. Measure and record the temperature of the acid
3. Add 5cm³ of hydrochloric acid to the polystyrene cup
4. Stir and replace the lid
5. When the thermometer stops changing, record the temperature
6. Repeat steps 3-5 until 40cm³ of hydrochloric acid has been added
7. Repeat the practical 2 more times
8. Calculate a mean

Plot results onto a graph

Energy Changes Practical - Results



Temperature increases as the reaction is exothermic. The temperature then decreases as the reaction is complete, all the NaOH has reacted. HCl is being added in excess, no further reaction takes place.

Two lines of best fit are drawn, where these cross shows us the volume of HCl needed to neutralise the sodium hydroxide

Variables

IV – volume of hydrochloric acid

DV – Temperature change

CV – concentration and volume of NaOH, concentration of HCl

Reasons for errors:

- Did not stir the chemicals properly
- Did not wait for temperature to stop changing before recording

Ways to make results more accurate:

- Use a polystyrene cup (reduces energy loss or gain)
- Use a lid (reduces energy loss or gain)
- Use a digital thermometer

Bond Energy Calculations

Bond energy calculations are used to find the overall energy change of a reaction.

Energy change (ΔH) = reactants - products

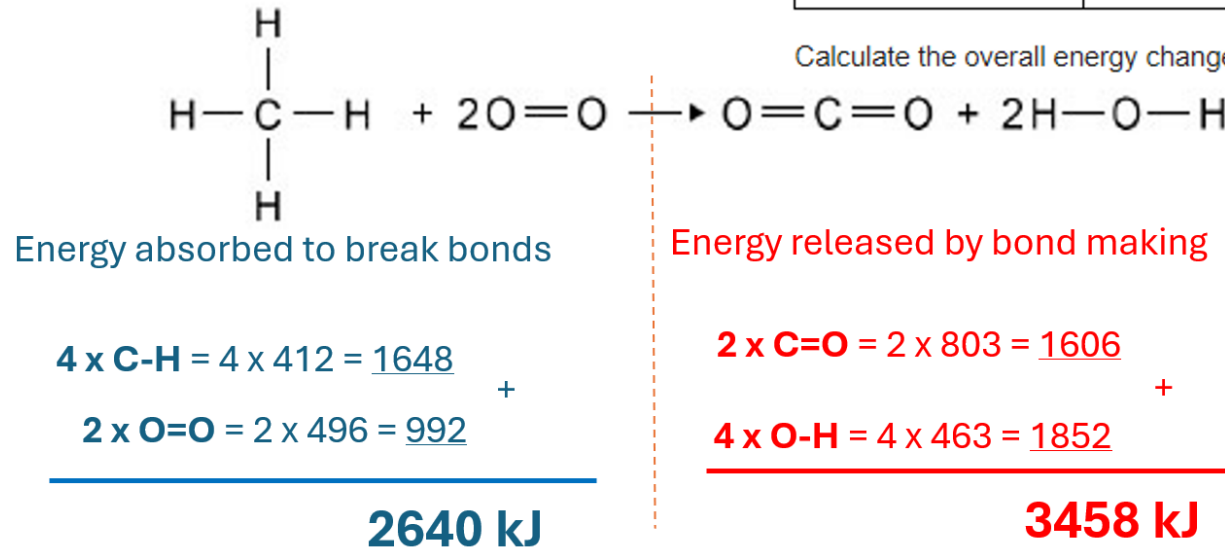
If ΔH is + the reaction is endothermic (more energy is absorbed for bond breaking than released in bond making)

If ΔH is - the reaction is exothermic (more energy is released through bond making than absorbed for bond breaking)

The table shows the bond energies.

Bond	C-H	O=O	C=O	O-H
Bond dissociation energy in kJ per mole	412	496	803	463

Calculate the overall energy change for the combustion of one mole of methane.



$$\text{Reactants} - \text{Products} = 2640 - 3458 = \underline{\underline{-818 \text{ kJ}}}$$

Exothermic – energy released from bond making is higher than energy absorbed for bond breaking

Cells

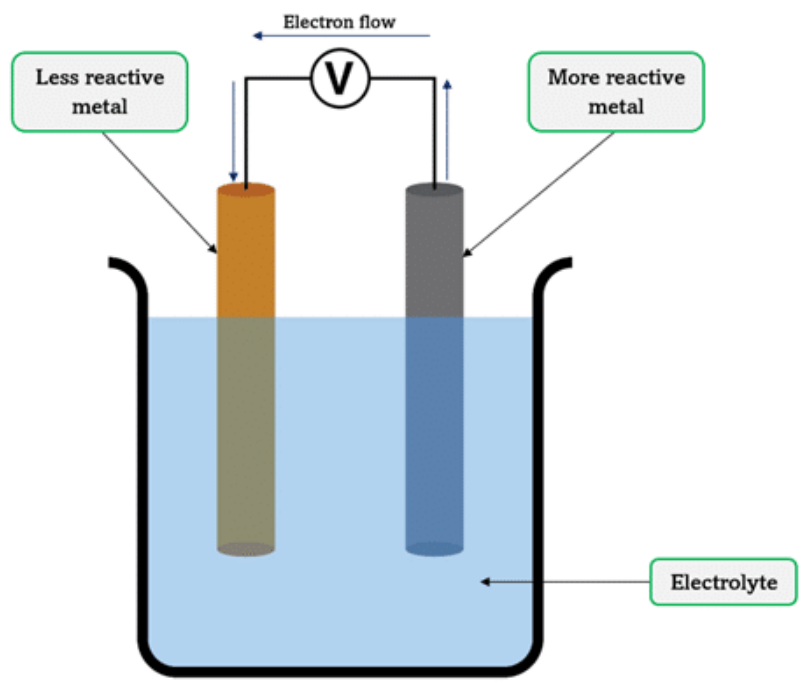
Cells contain chemicals which react to produce electricity

A simple cell is made of:

- Two different metal electrodes
- An electrolyte

Voltage is dependent on:

- The electrolyte (type and concentration)
- The electrodes (type – the bigger the difference in reactivity, the higher the voltage produced)



Batteries

Batteries are made of two or more cells connected in series, the voltage equals the voltage of the cells added together

Batteries can be rechargeable or non-rechargeable



Rechargeable cells and batteries can be recharged as the chemical reactions are reversed when external current is applied.

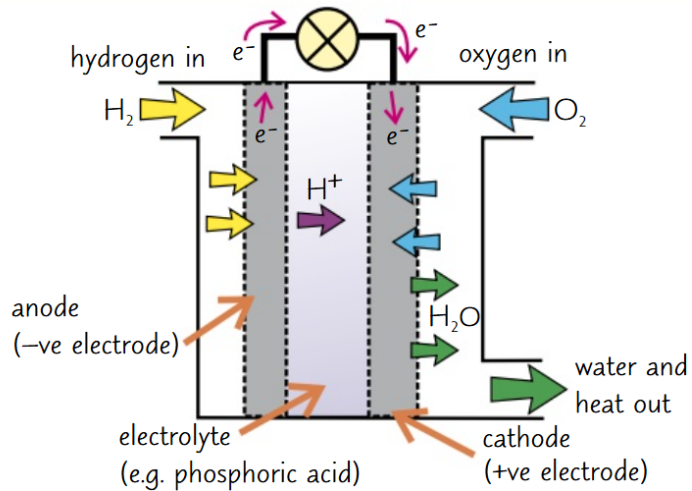
Non-rechargeable cells and batteries cannot be recharged as the chemical reactions stop when one of the reactants has been used up.
 Alkaline batteries are non-rechargeable.



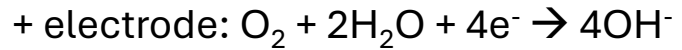
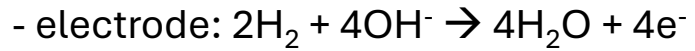
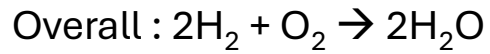
Hydrogen Fuel Cells

Fuel cells are supplied with an external source of fuel (e.g., hydrogen) and oxygen from the air

The fuel is oxidised electrochemically within the fuel cell to produce a potential difference



Equations:



Fuel cells Vs Batteries

	Hydrogen-oxygen fuel cells	Rechargeable Batteries
Pollutants	Water and Heat	Don't produce many pollutant but polluting to dispose of due to being made of highly toxic metal compounds
Cost	Cheaper to make than batteries	Need replacing eventually as you can only recharge them a certain number of times. Expensive to make
Use	Don't need replacing once installed	Store less energy so have to be recharged more often which can take a long time
Practicality	<ul style="list-style-type: none"> - Gases take up large spaces - hydrogen is also highly explosive so difficult to store and transport - Hydrogen is widely available at the moment - Difficult to make hydrogen (made from wither fossil fuels of electrolysis of water which again requires energy usually from burning fossil fuels) - Cars need to be specially adapted to burn hydrogen 	<ul style="list-style-type: none"> - Batteries only take up a small space. - Can't be used for long journeys. Charging stations would have to be places all over the country. - Electric cars are expensive at the moment. - Batteries won't last forever which is another cost for owners. - Cars need to be adapted to run off batteries but you can already get these and charge them at home