Chemistry Topic 1a - Atomic Structure Paper 1:KS4

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

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 (CO_2) is a compound so is not on the periodic table

Molecules can be element or compounds e.g., O_2 is an element and a molecule. H_2O 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?

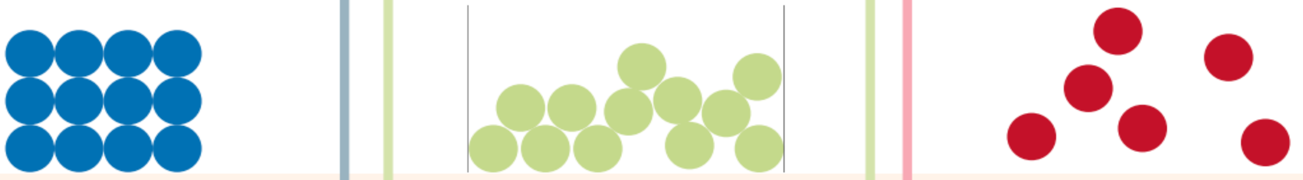
What is the difference between a compound and a mixture?

What is an isotope?

Chemistry Topic 1a - Atomic Structure Paper 1:KS4

States of Matter

Changes of state: melting (solid → liquid); boiling / evaporation (liquid → gas); freezing (liquid → solid); condensation (gas → liquid)

state of matter	solid	liquid	gas
How do the particles move?	particles do not move around, but vibrate on the spot	particles are touching but can slide over each other	particles are spread out far away from each other
arrangement of particles			
can it be compressed?	no, because there is no space between the particles	no, because the particles are touching their neighbours	yes, because there is space between the particles
can it flow?	no, because the particles can't move around	yes, because the particles can slide over each other and move around	yes, because the particles can move around

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)

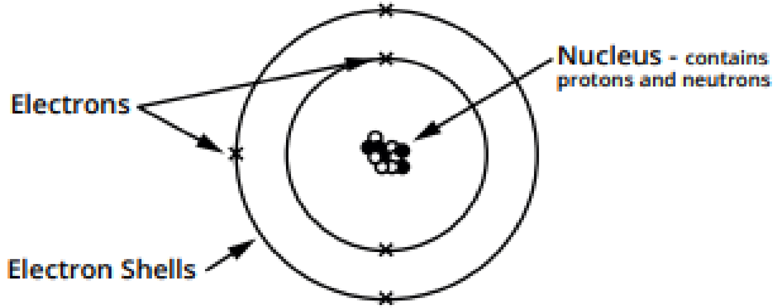
State symbols

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

*this means dissolved in water

Chemistry — Topic 1a - Atomic Structure — Paper 1:KS4

Structure of the atom:

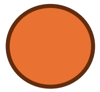


Size of atoms

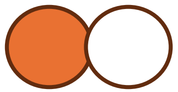
Atoms have a radius of 0.1nm (1×10^{-10}m)

The radius of the nucleus is 1/10 000 of that of an atom (1×10^{-14}m)

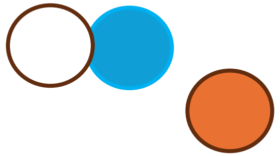
Atoms, elements, compounds and mixtures



Atoms + Elements



Compound



Mixture

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

Isotopes Calculations

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

^{35}Cl 75%

^{37}Cl 25%

e.g.

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

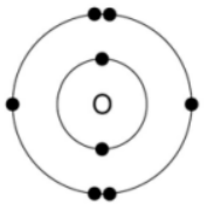
= 35.5

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 or electrons on each shell

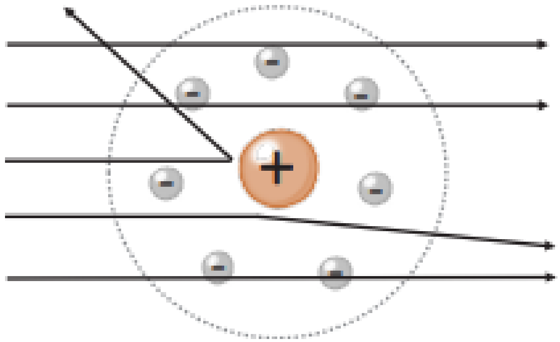


e.g., 2,6

The development of the model of the atom:

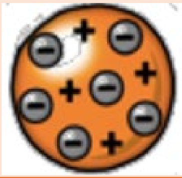
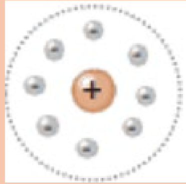
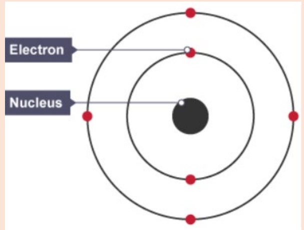
Rutherford's scattering experiment

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

7 ←	**Mass number**	*The sum of the protons and neutrons in the nucleus*	
Li			
3 ←	**Atomic number**	*The number of protons in the atom*	Number of electrons = number of protons

Chemistry Topic 1b - Periodic Table Combined Paper 1:KS4

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

Key Words

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

Key questions

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?

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

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

Chemistry Topic 1b - Periodic Table Combined Paper 1:KS4

Periodic table

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

1	2											3	4	5	6	7	0
							1 **H** hydrogen 1										4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4	Key: relative atomic mass **atomic symbol** name atomic (proton) number										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

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.

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

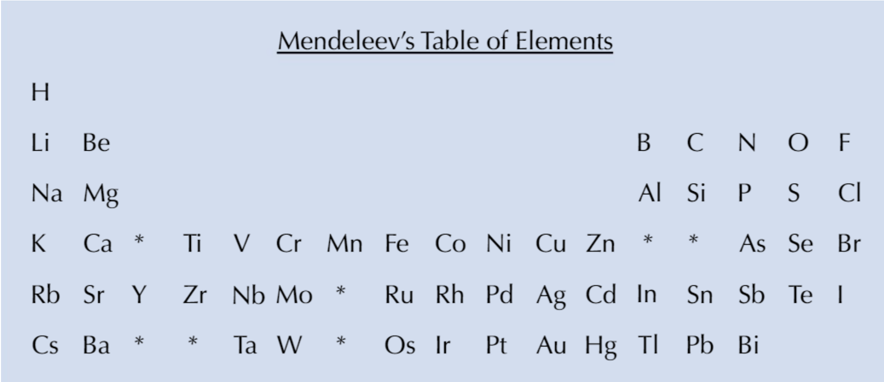
The rows are called **periods**. Elements in the same period have the **same number of electron shells**.

Development of the periodic table

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																
Li	Be											B	C	N	O	F
Na	Mg											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		

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

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

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

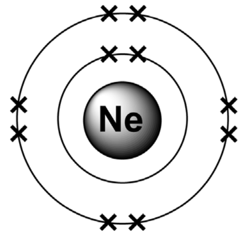
Elements that react to form negative ions are non-metals. They are found towards the right and top of the periodic table.

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

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., Cl_2 and Br_2
- Melting and boiling point increase as you go down the group
- Reactivity decreases as you go down the group

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.

Potassium bromide	+ Chlorine →	Potassium chloride	+ Bromine

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

Group 1 – alkali metals

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

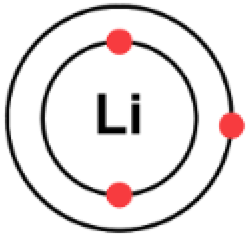
Reactivity

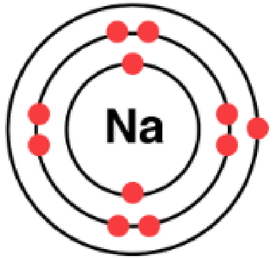
All alkali metals react by losing their outer shell electron

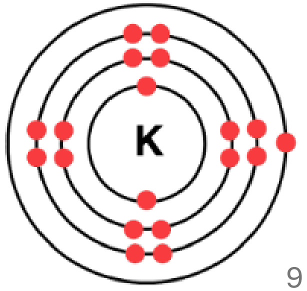
The reactivity of alkali metals increases as you go down the group

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 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

$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + H_{2(g)}$

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

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

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

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

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

Chemistry Topic 1c - Separating Mixtures Paper 1:KS4

Filtration and Crystallisation 1	Filtration and Crystallisation 1	Distillation	Chromatography	Chromatography R_f values	Check 20	Misconceptions

Key Words

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

Misconceptions

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?

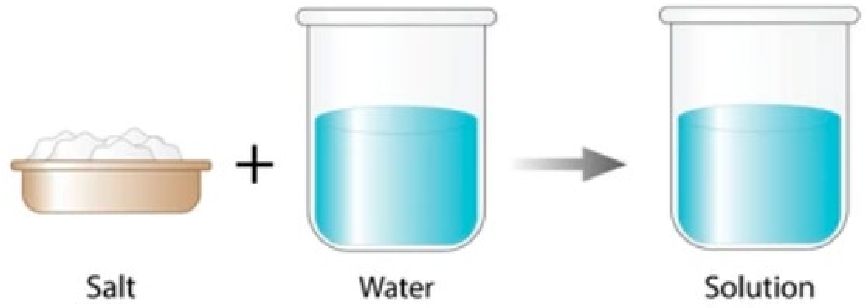
Mixtures

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

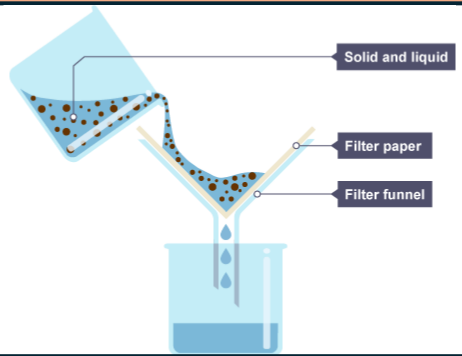


Salt = solute (gets dissolved)
Water = solvent (does the dissolving)
Salty water = solution (mixture of solute and solvent)

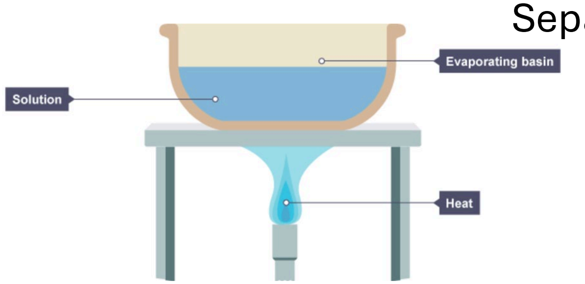
Filtration

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

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



Crystallisation



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

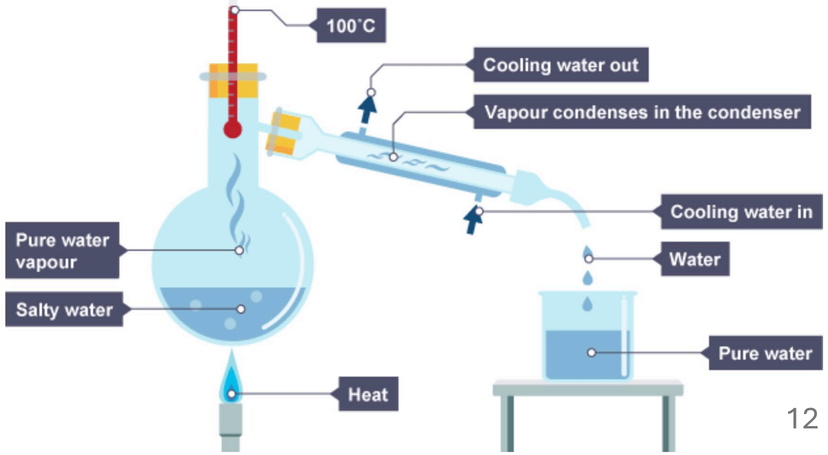
Water evaporates and leaves the solid behind

Distillation

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

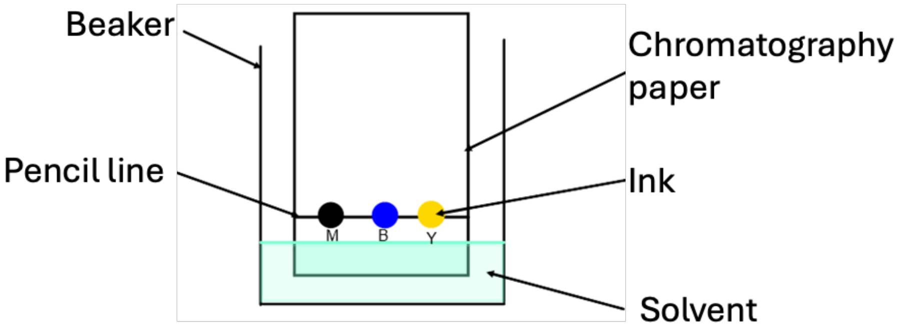
Two processes in distillation:
- Evaporation
- Condensation

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



Chromatography

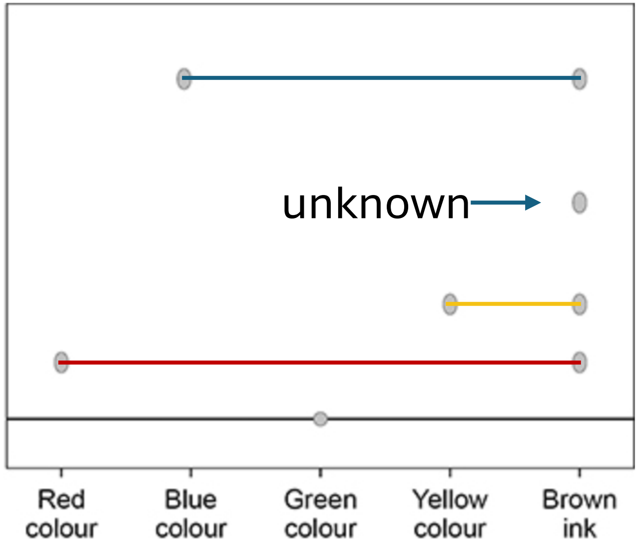
Method used to separate mixtures and give information to help identify substances



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

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 in there is red, yellow and blue and an unknown

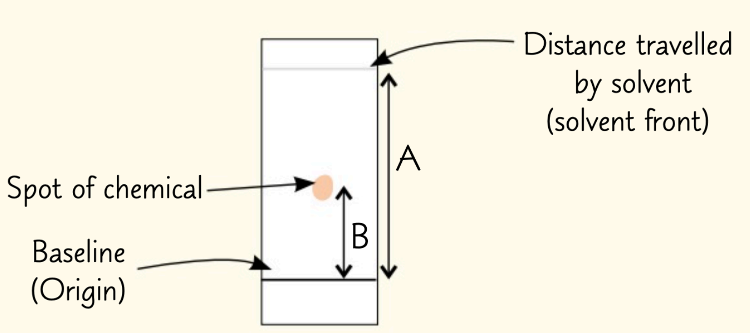
Phases of chromatography

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 mover** and the components move up it.

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

Chemistry Topic 2 - Bonding and Structure Paper 1:KS4

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	Check 20	Misconceptions

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

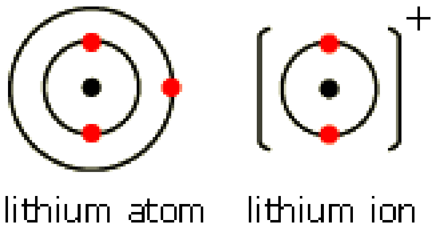
Why do ionic substances have high melting points?

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

Why are alloys harder than pure metals?

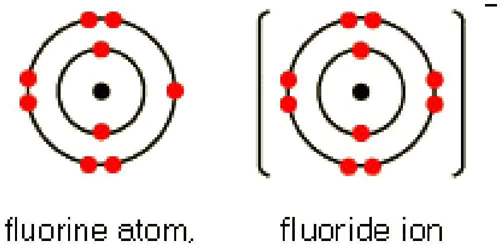
Formation of Ions

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

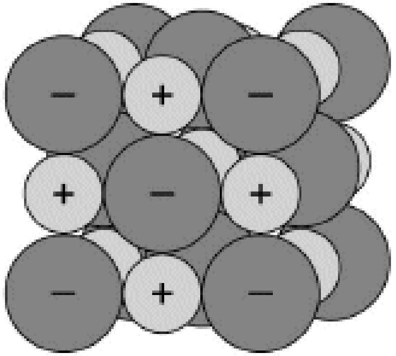
Non-metals gain electrons to form negatively charged ions



Ionic Bonding

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

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

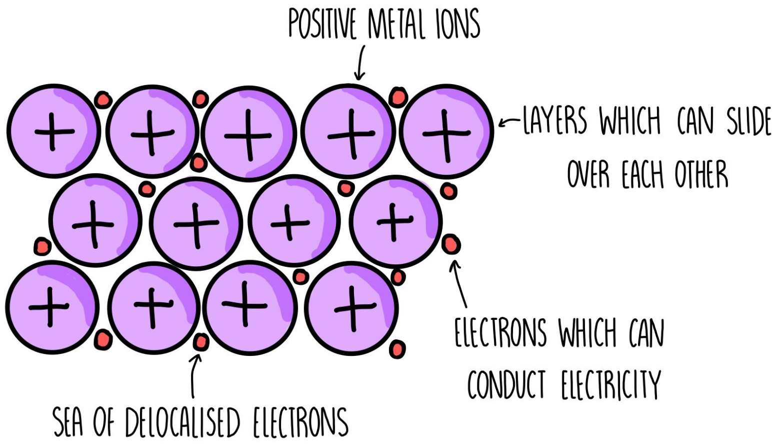


Properties

- 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**

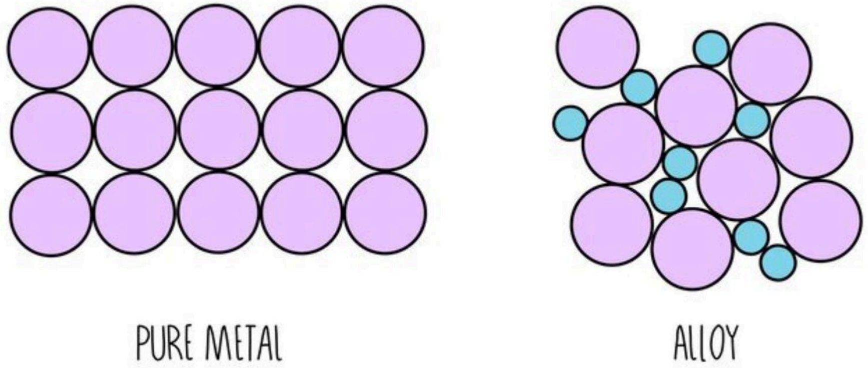


Properties

- 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

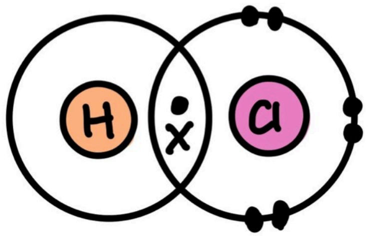
Mixture containing at least one metal



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.

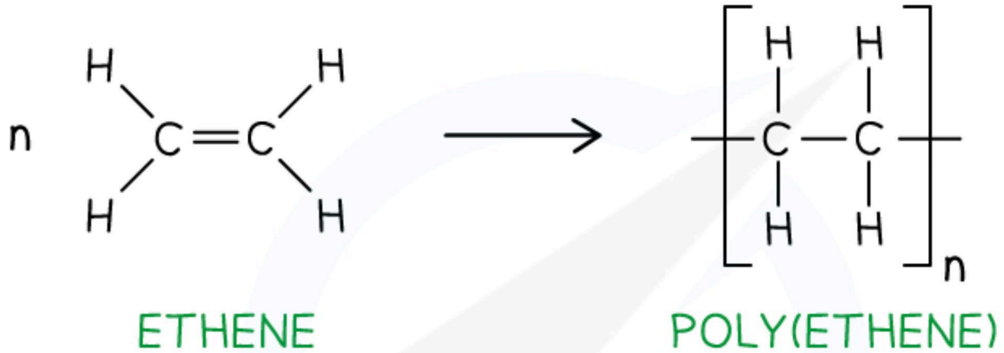


A covalent bond is a shared pair of electrons

Strong covalent bonds are drawn as ——— in diagrams

Polymers

Long chain molecules made up of repeating units called monomers

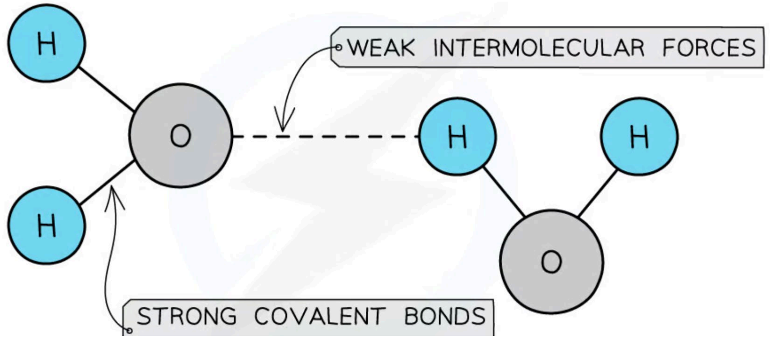


Properties

- 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.

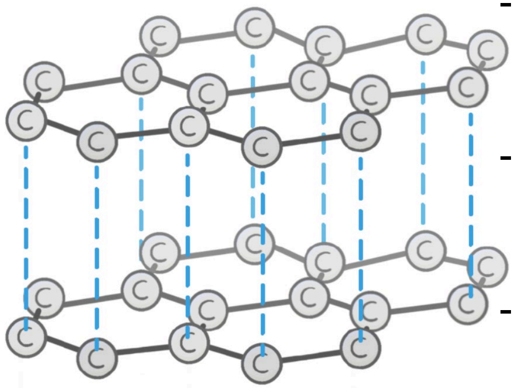
Properties

- 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

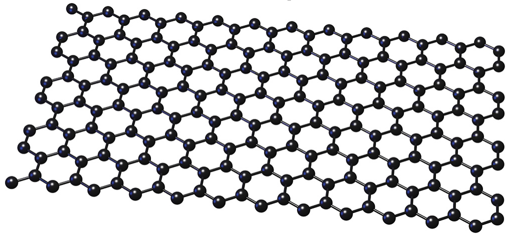
- High melting point → **Each carbon makes 3 strong covalent bonds**, lots of energy is needed to break the bonds
- Soft → Atoms are arranged in layers, the layers are held together by weak intermolecular forces so they can slide
- Conducts electricity → Each carbon makes 3 strong covalent bonds, so 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

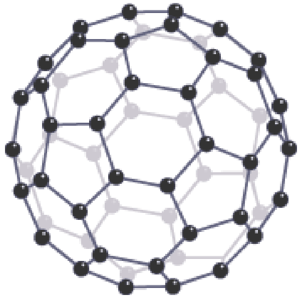
Uses: electronics and making composites



Fullerenes

Structures based on hexagonal rings of carbon

Buckminster fullerene: Spherical molecule made of 60 carbon atoms

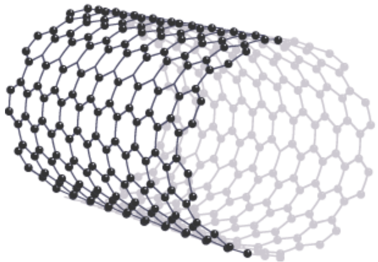


Uses:

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

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

$Mg(NO_3)_2$

Chemistry Topic 3a – Quantitative Chemistry 1 Paper 1:KS4

Balancing Equations	Conservation of mass 1	Conservation of mass 2	RFM	Percentage by mass	Concentration	Check 20	Misconceptions lesson

Key Words

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

Misconceptions

In RFM, bracket. The numbers are multiplied not added. E.g. $Ca(NO_3)_2$ the O is 3x2=6

When calculating % by mass you must work out the total mass of the element e.g. CO_2 for O you would do 16x2 = 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 dm^3 = ÷ by 1000

Key questions

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

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.

$2Na + Cl_2 \rightarrow 2NaCl$

2g 3g 5g

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

$2Na + Cl_2 \rightarrow 2NaCl$

Na = 2	Na = 2
Cl = 2	Cl = 2

Balancing Equations

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

$H_2 + O_2 \rightarrow H_2O$

H = 2	H = 2	**This is unbalanced**
O = 2	O = 1	

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

$2H_2 + O_2 \rightarrow 2H_2O$

H = 4	H = 4	**This is balanced**
O = 2	O = 2	

Reactions and conservation of mass

Conservation of mass is not always shown in some chemical reactions

1. **Where one the reactants is a gas.**

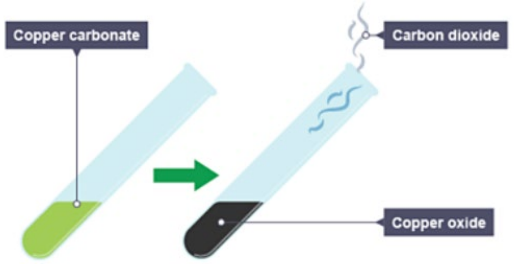
e.g. $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

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**

e.g. $CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$

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



Chemistry Topic 3a – Quantitative Chemistry 1 Paper 1:KS4

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. $MgCl_2$

Mg = 1 x 24 = 24

Cl = 2 x 35.5 = 71

24 + 71 = 95

e.g. $Ca(NO_3)_2$

Ca = 1 x 40 = 40

N = 2 x 14 = 28

O = 6 x 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

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

Steps:

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

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

$MgCl_2$

Mg = 1 x 24 = 24

Cl = 2 x 35.5 = 71

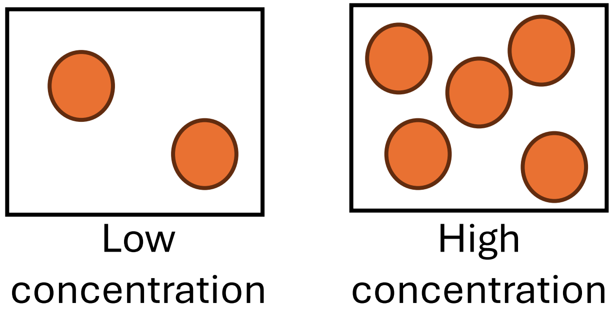
24 + 71 = 95

$$\% \; by \; mass = \frac{71}{95} \; X \; 100$$

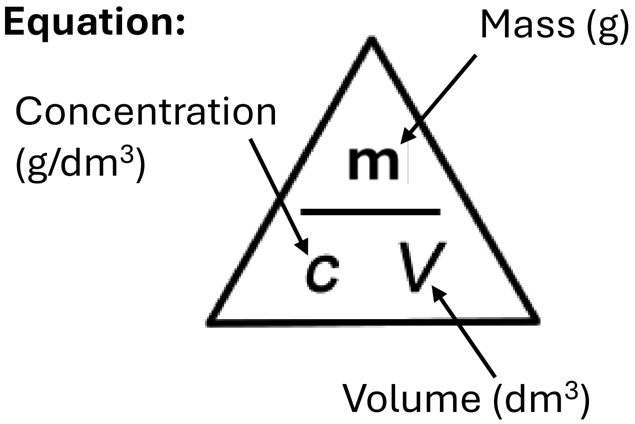
= 74.7%

Concentration

It is the number of particles in a given volume



Equation:



cm^3 to dm^3 = ÷ by 1000

e.g. Calculate the concentration when 5g of NaCl is dissolved in 200cm^3

V = 200cm^3 ÷ 1000 = 0.2dm^3

C = m ÷ v

C = 5 ÷ 0.2 = 25g/dm^3

Chemistry Topic 4a - Chemical Changes Combined F Paper 1:KS4

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

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

Litmus paper cannot be used to show the pH of a substance. The most accurate way to find pH is a pH probe.

When a metal reacts with acid. Hydrogen gas is produced. MASH

Not all metals are extracted using mining. Ores (compound containing metal) are mined, but most metals must be extracted from the ore.

Carbon is not a metal. It is part of the reactivity series as it is used to extract some metals.

Key questions

What are the products of a neutralisation reaction?

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

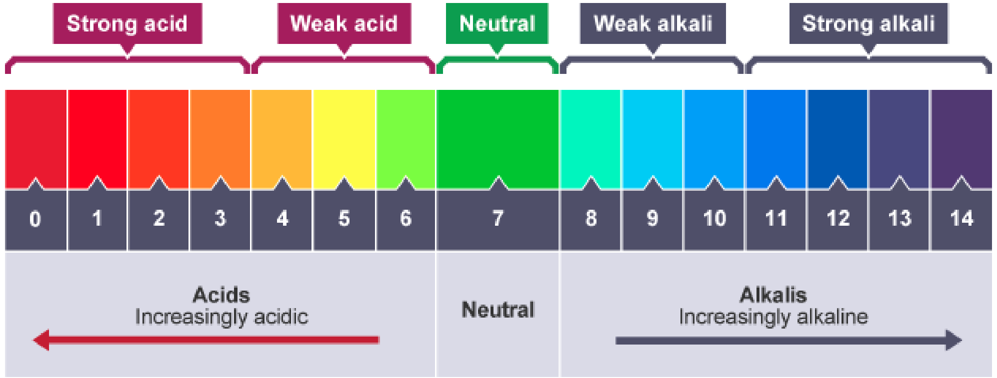
What is a displacement reaction?

What do the terms oxidation and reduction mean?

List the reactivity series?

Chemistry Topic 4a - Chemical Changes Combined F Paper 1:KS4

pH scale



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

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**: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

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**

Reactions of metals

Metal + Water → Salt + Hydrogen

Metal + Acid → Salt + Hydrogen

Metal Oxide + Acid → Salt + Water

Metal Hydroxide + Acid → Salt + Water

Metal Carbonate + Acid → Salt + Water + Carbon dioxide

Naming Salts

Copper oxide + Sulfuric acid → Copper sulfate + Water

From the metal   From the acid

Salt endings:

Hydrochloric acid (HCl) = Chloride

Sulfuric acid (H_2SO_4) = Sulfate

Nitric acid (HNO_3) = Nitrate

Chemistry Topic 4a - Chemical Changes Combined F Paper 1:KS4

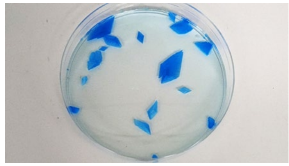
Making Salts

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

→ Extracted using electrolysis because they are more reactive than carbon

Cute – Carbon

Zebra - Zinc
I - Iron
Like - Lead

→ 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

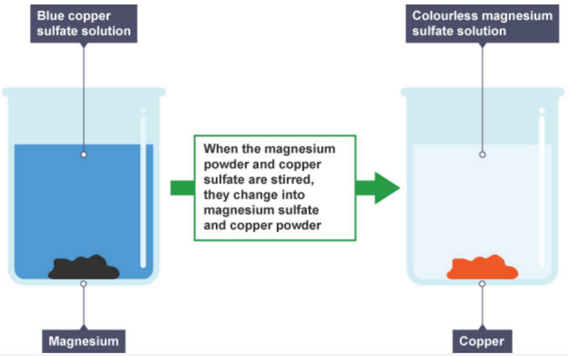
Hilary - Hydrogen

Clinton - Copper
She's – Silver
Great - Gold

→ 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 it more reactive

Observations (what you can see)
- Copper metal at the bottom
- Colour change from blue to colourless

Chemistry — Topic 4b - Chemical Changes F — Paper 1:KS4

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

Key Words

Key Word	Definition
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

Key questions

What is an electrolyte?

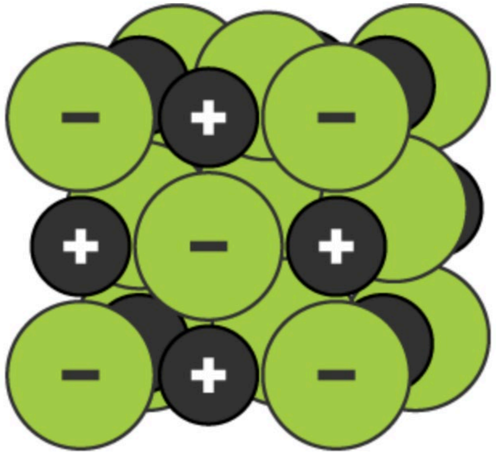
What does PANIC stand for?

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

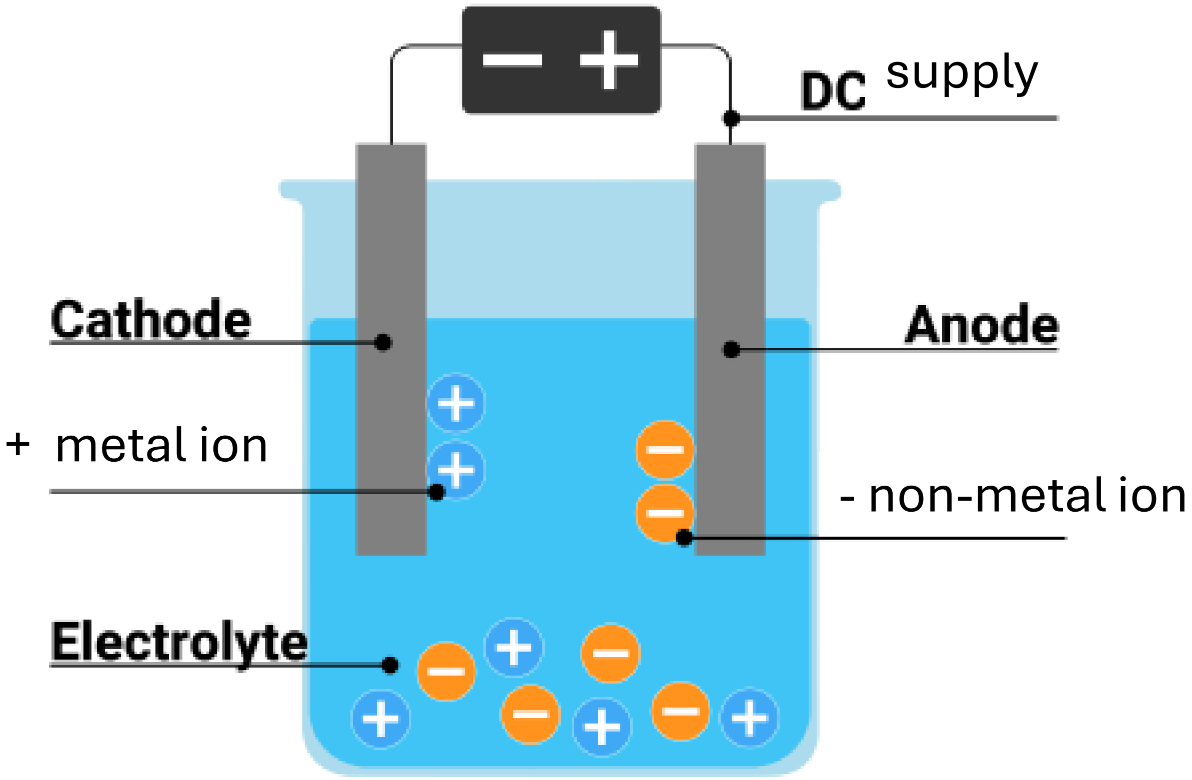
Ionic compounds have high melting points because:
- They have a giant ionic lattice structure
- Oppositely charged ions are held together by strong electrostatic forces
- Lots of energy is needed to break the forces

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 becoming atoms again

- non-metal ions are attracted to the positive electrode (anode), here they are discharged becoming atoms again

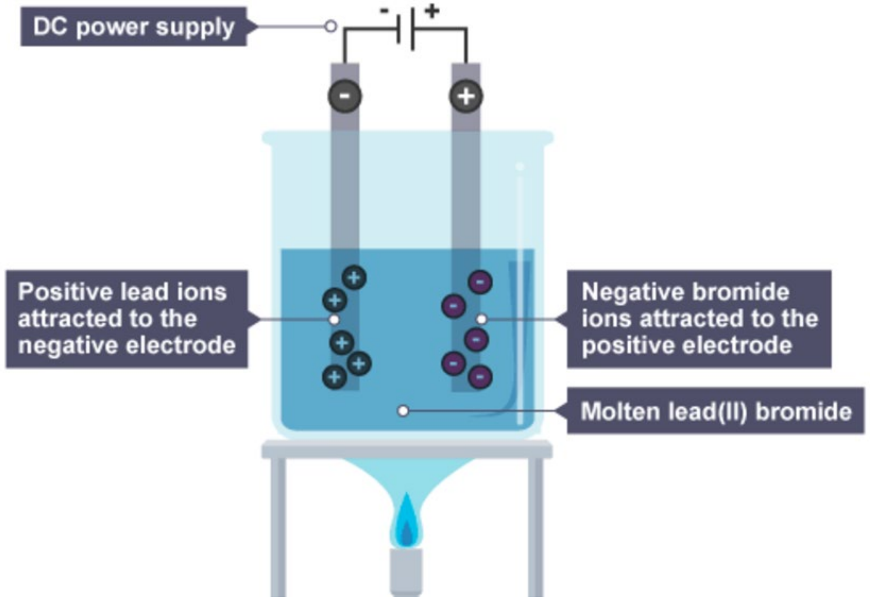
Molten Electrolysis	Rules for Aqueous electrolysis

Molten Electrolysis

Positive metal ions go to the negative electrode as opposite charges attract.

Negative non-metal ions go to the positive electrode as opposite charge attract.

Example: Molten Lead Bromide



Pb^{2+} ions go the negative electrode

Br^- ions go to the positive electrode where they 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

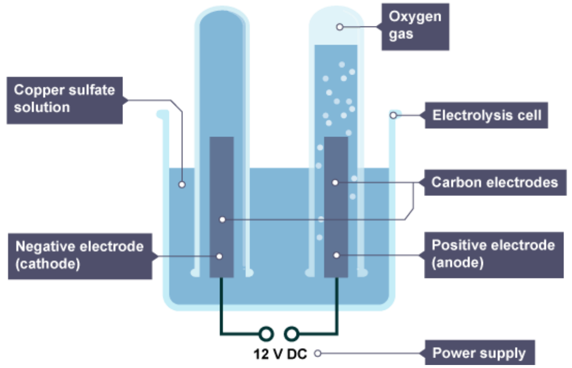
At the + electrode (anode):

- 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 as no halogens are present

At the - electrode (cathode):

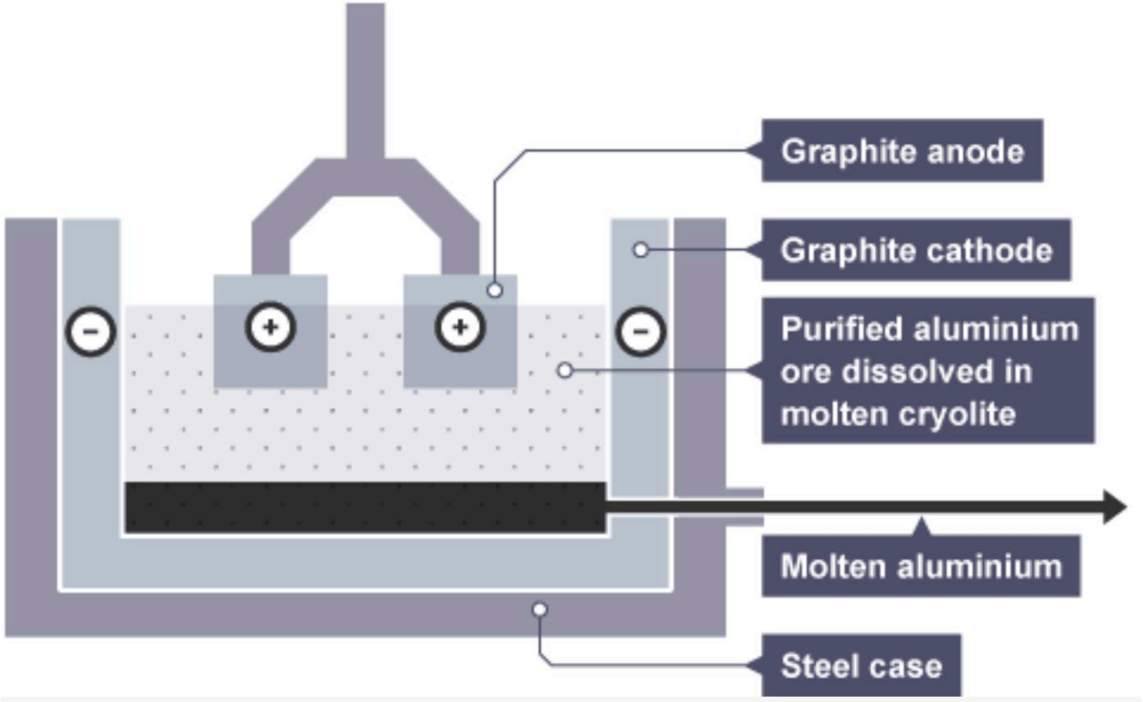
H^+ and Cu^{2+} ions go to the - electrode

Copper is made as it is less reactive than hydrogen

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

At the + electrode (anode):
O^{2-} ions go to the + electrode
Oxygen is made

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

Chemistry Topic 5 - Energy Changes F Paper 1:KS4

Exothermic and Endothermic Reactions	Energy Changes Practical 1	Energy Changes Practical 2	Reaction Profiles	Check 20	Misconceptions

Key Words

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

Misconceptions

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?

Chemistry Topic 5 - Energy Changes F Paper 1:KS4

Exothermic Reactions	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

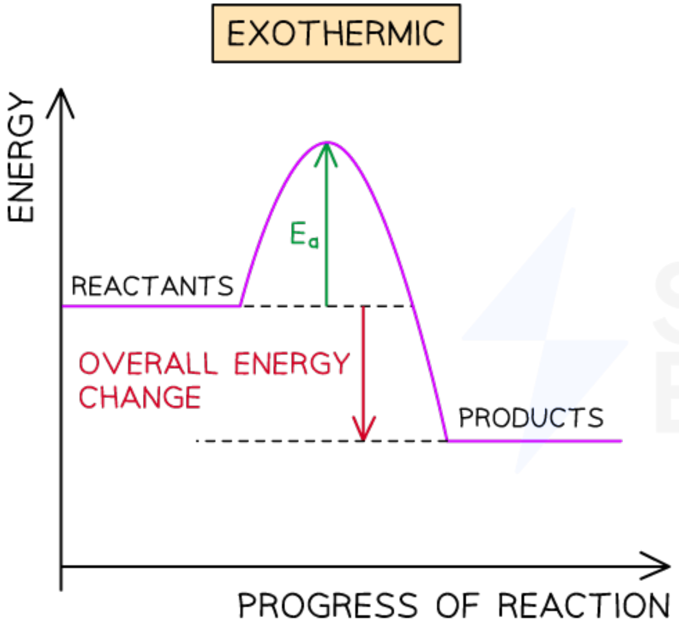
ENDO = ENTERS

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

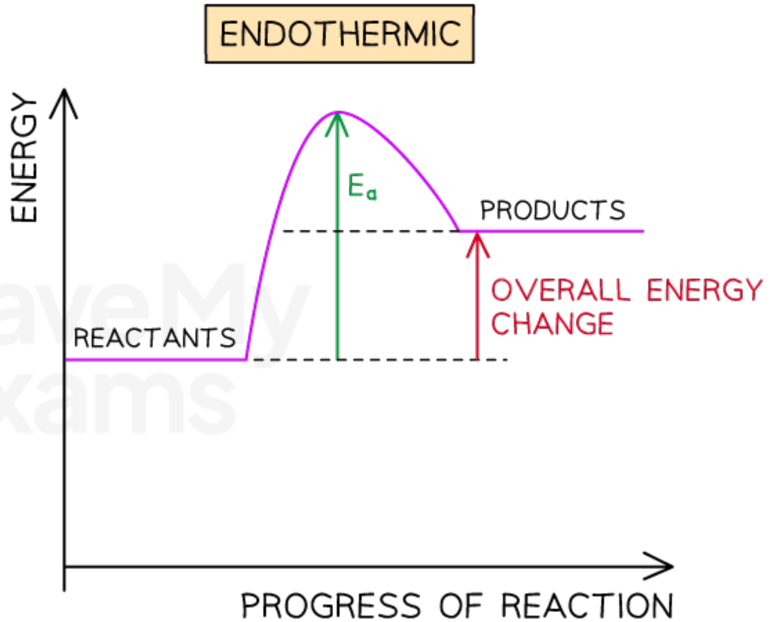
Reaction Profiles



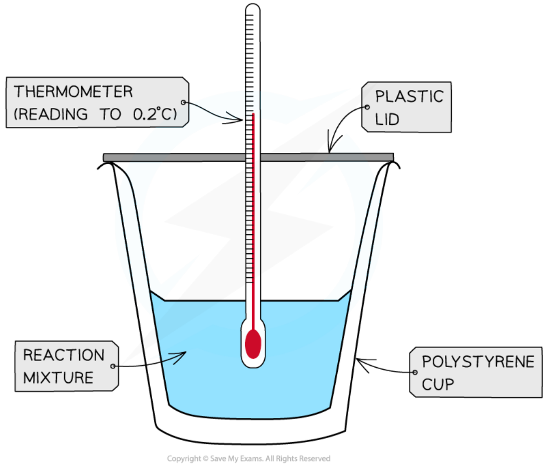
E_a = activation energy

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



Energy Changes Practical

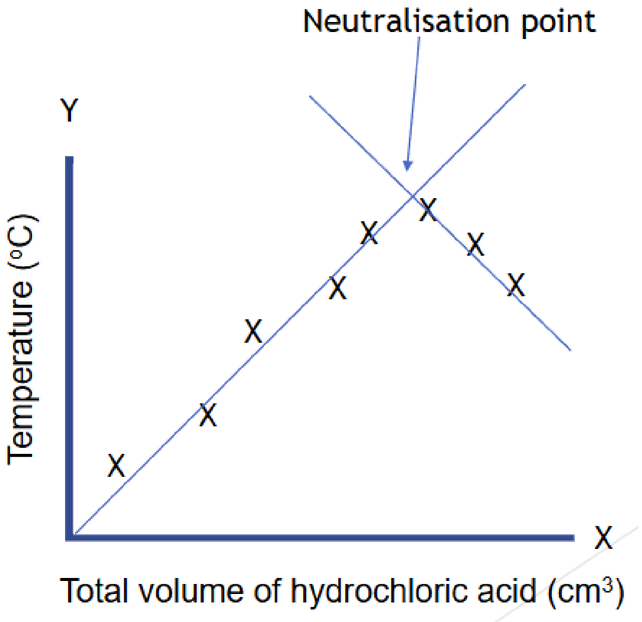


Neutralisation reaction between sodium hydroxide and hydrochloric acid

1. Add 25cm^3 of sodium hydroxide into a polystyrene cup
2. Measure and record the temperature of the acid
3. Add 5cm^3 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 40cm^3 of hydrochloric acid has been added
7. Reat 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