

# Chemistry

## Paper 1 Knowledge Organisers

AQA Combined Science (Trilogy)

NAME: \_\_\_\_\_

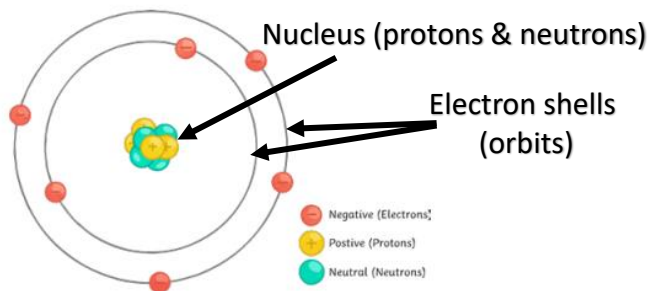
CLASS: \_\_\_\_\_

TEACHER: \_\_\_\_\_

# C1 – Atomic Structure and The Periodic Table

## Atoms

- Made up of **protons, electrons and neutrons.**



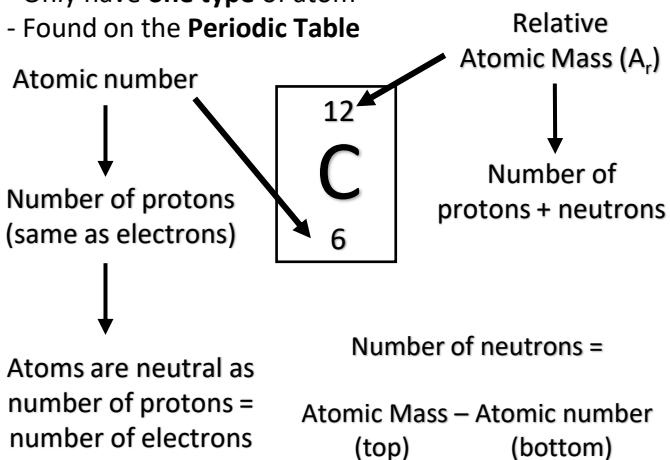
Subatomic particle	Relative Mass	Charge
Proton	1	Positive
Neutron	1	Neutral
Electron	Very small	Negative

Atoms have a radius of about  $0.1\text{nm}$  ( $1 \times 10^{-10}\text{m}$ )

Radius of nucleus = about  $1 \times 10^{-14}\text{m}$

## Elements

- Only have **one type** of atom
- Found on the **Periodic Table**



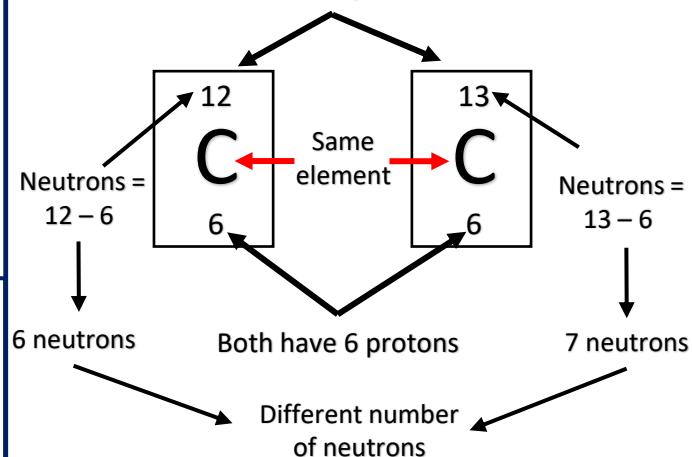
## Compounds

- Two or more elements **chemically combined.**
- Formed by chemical reactions
- For example:  $\text{CO}_2$   $\text{H}_2\text{O}$   $\text{CH}_4$   $\text{HCl}$   $\text{NaCl}$

## Isotopes

**Isotope** = atoms of the **same element** which have the **same number of protons**, but a **different number of neutrons.**

These are isotopes because..



## Chemical Equations

- Shown by using a **word equation.**
- e.g. magnesium + oxygen  $\rightarrow$  magnesium oxide

Left of the arrow = **reactants**

Right of the arrow = **products.**

- Also can be shown by a **symbol equation**
- e.g.  $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

## Mixtures and Separation

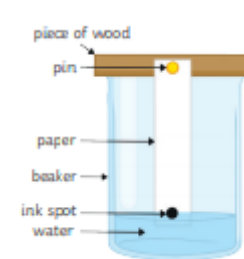
**Mixtures** – two or more elements or compounds **not** chemically joined.

This means the different components of the mixture can be separated by physical methods (below)

E.g. air is a mixture mainly made of nitrogen, oxygen and carbon dioxide.

### Chromatography

to separate out mixtures (usually liquids) (e.g. colours in ink)



### Filtration

To separate insoluble solids from liquids (e.g. sand and water)



### Evaporation

To quickly separate soluble solids from a solution. (e.g. salt and water)



### Crystallisation

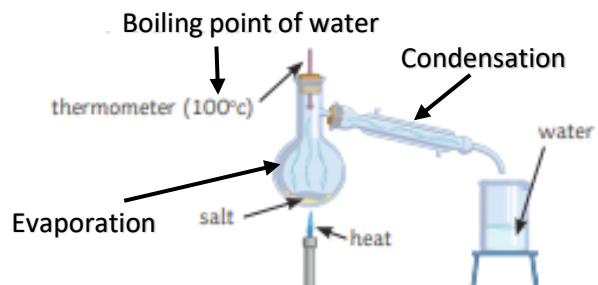
To slowly separate a soluble salt from a solution. (e.g. copper sulfate crystals)



# C1 – Atomic Structure and The Periodic Table

## Distillation

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

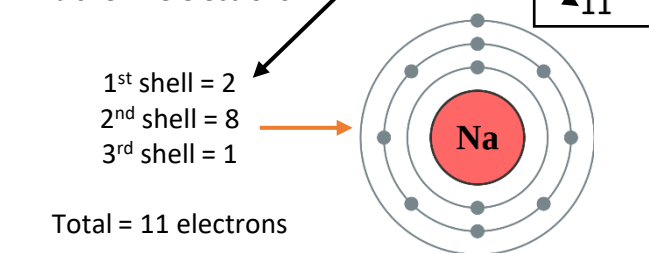


- Liquid is heated to boiling point and evaporates
- Vapours travel up into the condenser
- Condenser has cold water around it.
- Vapours cool and condense (turn back into a liquid).

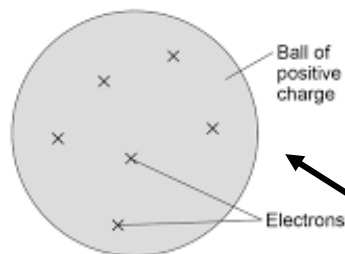
## Electronic Structure

- Electrons are found on shells (orbits) orbiting the nucleus.
- There is a maximum number of electrons allowed on each shell:

First shell = 2 electrons  
 Second shell = 8 electrons  
 Third shell = 8 electrons.



Plum pudding model



### Differences to nuclear model

- Ball of positive charge (no protons)
- No nucleus
- No neutrons
- Evenly distributed mass

Rutherford tested the plum pudding model

## History of the atom

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

What happened?

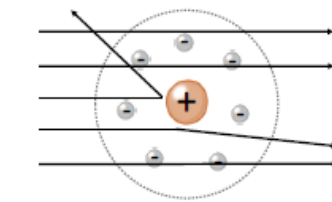
### Rutherford's scattering experiment

alpha particles are positively charged

Fired at gold foil

some alpha particles are deflected/ repelled

most alpha particles passed straight through



Conclusions made

Observation	Conclusion
Most of the particles passed straight through	Most of the atom is empty space
Some were deflected to the sides	The particles had passed close by a positive charge
A very small number were repelled straight back	The alpha particles had approached the nucleus straight on. the tiny number told him that the positive charge is in a very small dense core

# C1 – Atomic Structure and The Periodic Table

## Development of the Periodic Table

### John Newlands – Law of Octaves

- Elements ordered by **atomic weight**.
- Noticed a pattern with every eighth element.
- Some elements placed inappropriately – metals and non-metals grouped together.
- Rejected by other scientists.

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
Co, Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce, La	Zr	Di, Mo	Ro, Ru

John Newlands' Law of Octaves

### Dimitri Mendeleev

- Still ordered by atomic weight
- Left gaps for **undiscovered elements**
- Could predict properties of undiscovered elements.
- Some elements didn't fit pattern – switched them to keep pattern of **similar properties**.

H 1.01	II	III	IV	V	VI	VII						
Li 6.94	Be 9.01	B 10.8	C 12.0	N 14.0	O 16.0	F 19.0						
Na 23.0	Mg 24.3	Al 27.0	Si 28.1	P 31.0	S 32.1	Cl 35.5			VIII			
K 39.1	Ca 40.1		Ti 47.9	V 50.9	Cr 52.0	Mn 54.9	Fe 55.9	Co 58.9	Ni 58.7			
Cu 63.5	Zn 65.4			As 74.9	Se 79.0	Br 79.9						
Rb 85.5	Sr 87.6	Y 88.9	Zr 91.2	Nb 92.9	Mo 95.9		Ru 101	Rh 103	Pd 106			
Ag 108	Cd 112	In 115	Sn 119	Sb 122	Te 128	I 127						
Ce 133	Ba 137	La 139		Ta 181	W 184		Os 194	Ir 192	Pt 195			
Au 197	Hg 201	Tl 204	Pb 207	Bi 209								
		Th 232			U 238							

Dimitri Mendeleev left gaps for undiscovered elements

Eventually, knowledge of isotopes explained why elements could not be ordered by atomic weight.

## The Modern Periodic Table

- Ordered by **atomic (proton) number**.

Columns = groups

Group number = number of electrons in outer shell.

Elements in each group have similar properties.

	1	2		3	4	5	6	7	0										
1	Li	Be		B	C	N	O	F	He										
2	Na	Mg		Al	Si	P	S	Cl	Ar										
3	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
4	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
5	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
6	Fr	Ra	Ac																
7																			

non-metals

metals

Rows = periods

Period number = number of electron shells the atom has.

## Group 1 (alkali metals)

- Similar properties as all have 1 electron in outer shell.
- All lose one electron in reactions to form 1+ ions
- Soft, grey, shiny metals
- Stored in oil as would react with oxygen in air.
- When placed in water they produce an alkali (hence alkali metals) and hydrogen gas

E.g Lithium + water → lithium hydroxide + hydrogen

## Reactivity of Group 1

Li
Na
K
Rb
Cs
Fr

As you go down the group...

- Elements are more reactive because:
- More electron shells
- Outer electron = further from nucleus and more shielded by the other shells
- The electrostatic force of attraction between outer electron and nucleus is weaker
- Easier for outer electron to be lost

## Group 7 (Halogens)

- 7 electrons in outer shell – all react similarly
- All gain one electron when they react to form 1- ions
- Form molecules (e.g. Cl<sub>2</sub>, F<sub>2</sub>)
- Non-metals.
- A more reactive halogen can replace a less reactive halogen in a reaction (**displacement**)

## Reactivity of Group 7

F
Cl
Br
Kr
Xe
Rn

As you go down the group...

- Elements are **less** reactive because:
- More electron shells
- Outer shell is further from nucleus and is **more shielded** by the other shells
- The electrostatic force of attraction between free electron and nucleus is **weaker**
- Harder to attract an electron into the outer shell.

## Group 0 (Noble Gases)

- Full outer shell – unreactive as they don't need to lose or gain any electrons

As you go down...

- Boiling point increases
- More electron shells
- Bigger atoms
- More intermolecular forces
- More energy needed to break forces.

# C2 – Bonding, structure, and the properties of matter

## Formation of Ions

- **Ions** = a charged particle made when atoms lose or gain electrons
- **Positive ion** = atom has lost electrons
- **Negative ion** = atom has gained electrons.

Metals form **positive ions**

Non-metals form negative ions

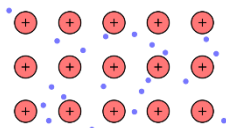
Group	Ions	Example
1	+1	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$
2	+2	$\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$
6	-2	$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$
7	-1	$\text{Br} + \text{e}^- \rightarrow \text{Br}^-$

Lost electrons

Gained electrons

## Metallic Bonding

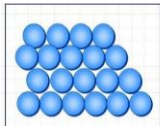
- Happens in **metals only**.
- Positive metal ions surrounded by **sea of delocalised electrons (can move)**.
- Ions tightly packed in rows.
- Strong **electrostatic forces of attraction** between positive ions and negative electrons.



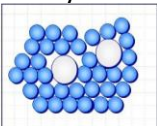
## Alloys

- **Alloys** = mixture of two or more metal atoms
- Pure metals are too soft for many uses.

Pure Metal



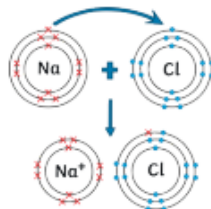
Alloy



- |                   |                         |
|-------------------|-------------------------|
| • Atoms same size | • Different sized atoms |
| • Layers slide    | • Layers cannot slide   |
| • Softer          | • Stronger              |

## Ionic Bonding

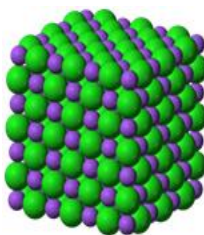
- Between a metal and non-metal.
- Metals give electrons to non-metals so both have a full outer shell.
- **Electrostatic force of attraction** between positive and negative ions.



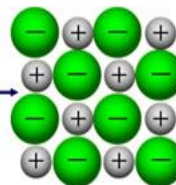
E.g. Sodium loses one electron to become  $\text{Na}^+$ . Chlorine gains one electron to become  $\text{Cl}^-$ . The two ions attract to form sodium chloride.

## Ionic compounds

- Form **giant lattices, as the attraction between ions acts in all directions**



strong electrostatic forces between oppositely charged ions



## Properties of Ionic Compounds

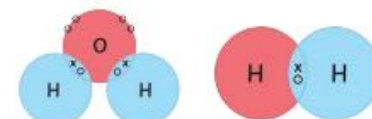
- **High melting point** – lots of energy needed to overcome electrostatic forces.
- **High boiling point**
- **Cannot conduct electricity as solid** – ions cannot move
- **Conducts electricity when molten or dissolved** – ions are free to move.

## Covalent Bonding

- **Covalent bonding** = sharing a pair or pairs of electrons for a full outer shell.
- Between **non-metals only**.

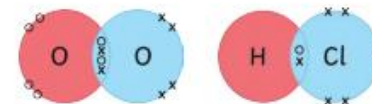
## Dot and cross diagrams

- Show the bonding in simple molecules.
- Uses the outer shell of the atoms
- Crosses and dots used to show electrons
- You should be able to draw the following:



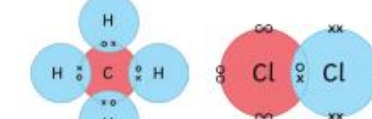
Water ( $\text{H}_2\text{O}$ )

Hydrogen ( $\text{H}_2$ )



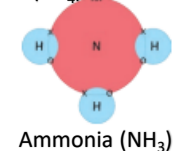
Oxygen ( $\text{O}_2$ )

Hydrogen chloride ( $\text{HCl}$ )



Methane ( $\text{CH}_4$ )

Chlorine ( $\text{Cl}_2$ )



Ammonia ( $\text{NH}_3$ )

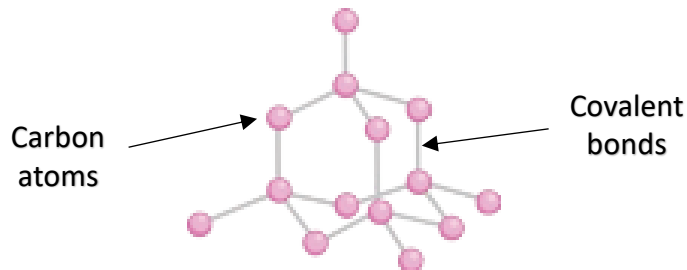
## Simple Covalent Molecules

- Form when all atoms have full outer shells so bonding stops
- Examples are the molecules shown above.
- Have **low melting and boiling points**
- Due to **weak intermolecular forces**
- Do not conduct electricity

## C2 – Bonding, structure, and the properties of matter

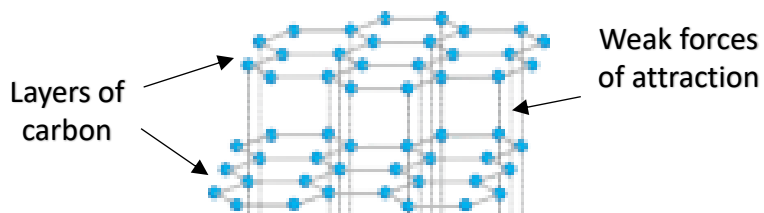
### Giant Covalent Structure – Diamond

- Each carbon atom **covalently** bonded to **four** others.
- Forms a giant structure
- This makes diamond **strong** → a lot of **energy** needed to break lots of strong covalent bonds.
- **Does not conduct electricity** – has no free electrons.



### Giant Covalent Structure – Graphite

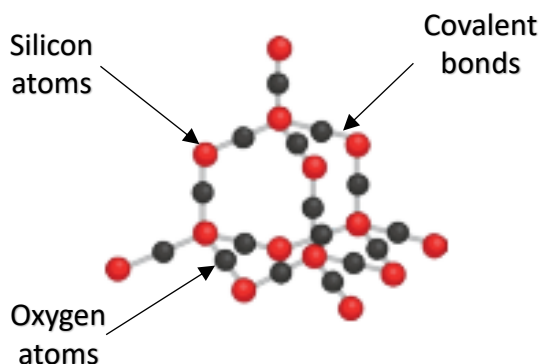
- Layers of **carbon** arranged in **hexagons**.
- Each carbon bonded to **three** other carbons.
- Leaves **one delocalised electron** → moves to carry electrical charge **throughout structure**.



- Layers held together by **weak forces**
- Layers can **slide** over each other easily
- Makes graphite **soft/slippery** → good lubricant.
- Has **high melting point** as has many strong covalent bonds.

### Silicon Dioxide

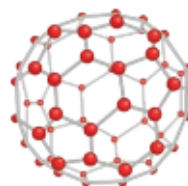
- Similar structure to diamond
- Giant covalent structure.
- Lots of **strong covalent bonds**.
- These require lots of **energy** to break.
- High melting and boiling points.



### Fullerenes and Nanotubes

- Molecules of carbon shaped into hollow tubes or balls.
- Used to **deliver drugs into body**

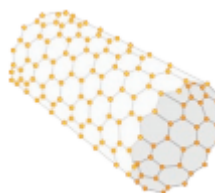
Buckminsterfullerene  
Formula = C<sub>60</sub>



- **Carbon nanotubes** = long narrow tubes
- Can conduct electricity

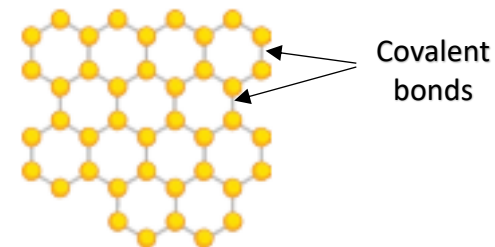
- Can strengthen materials without adding weight.

- Used in electronics and nanotechnology.



### Graphene

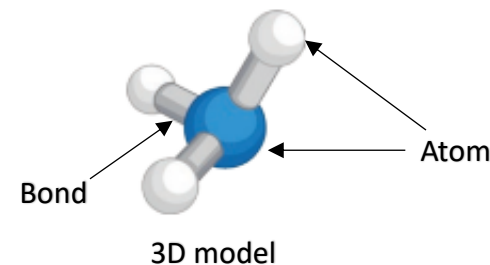
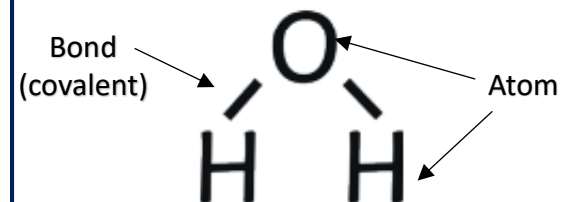
- Graphene = one layer of graphite.
- Very strong → lots of strong covalent bonds.



- Each carbon bonded to three others.
- One **free delocalised electron** → can move to **carry electrical current** throughout the structure.

### Molecular models

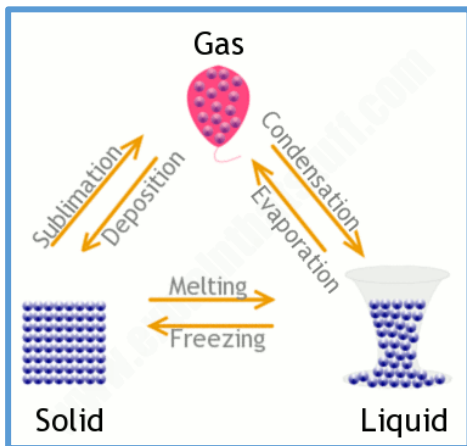
- There are different ways to show a molecule other than dot and cross diagrams.



## C2 – Bonding, structure, and the properties of matter

### States of Matter

- Three states of matter: **solid, liquid & gas.**
- To change state, **energy must be transferred.**



- When heated, particles **gain energy.**
- **Attractive forces** between particles begin breaking when melting or boiling points are reached
- **Amount of energy** needed to change state depends on how strong forces are.

### Gas

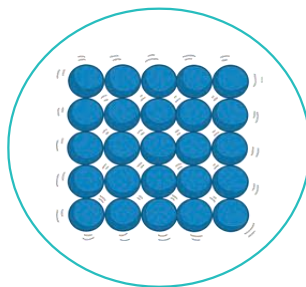
- Randomly arranged.
- Particles **move quickly** – all directions.
- Highest **amount of kinetic energy.**



- Gases **are able to flow** – fill containers
- **Can be compressed** as there is space between particles

### Solid

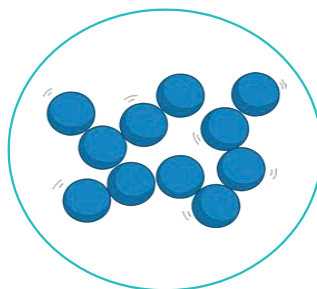
- **Regular** pattern (rows and columns)
- Particles **vibrate** in a **fixed position.**
- Particles have **low amount of kinetic energy.**



- Have a **fixed shape** – cannot flow because of strong forces of attraction between particles
- **Cannot be compressed** – particles close together.

### Liquid

- Particles **randomly** arranged and touching.
- Particles can **move around.**
- **Greater amount of kinetic energy** than solid



- Liquids **able to flow** – take shape of containers.
- **Cannot be compressed** – particles are close together and cannot be pushed closer

### State symbols

- States of matter shown in chemical equations:
- Solid (**s**)
- Liquid (**l**)
- Gas (**g**)
- Aqueous (**aq**)
- **Aqueous** solutions = substance dissolved in water.

### Identifying Physical State of Substances

- If the temperature is **lower** than a substance's melting point – substance is **solid.**
- If the temperature is **between** the melting point and boiling point – substance is **liquid.**
- If the temperature is **higher** than the boiling point – substance is a **gas.**

### Limitations of Particle Model (HT)

- No chemical bonds are shown.
- Particles shown as solid spheres – not the case, particles are mostly empty space like atoms.
- The diagrams don't show any of the forces between particles
- The diagrams are unable to show the movement of the particles.

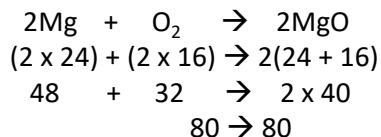
# C3 – Quantitative Chemistry

## Conservation of Mass

- Atoms cannot be created or destroyed during reactions.
- **Mass of reactants = mass of products.**

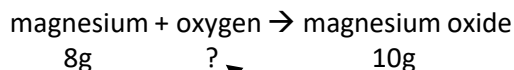
To show mass is conserved in a reaction:

$M_r$  on the left-side must be same as the right side.



## Reacting masses

Use conservation of mass to predict masses:



Both sides need to be equal:  
 $10\text{g} - 8\text{g} = 4\text{g}$  of oxygen

## Percentage Mass

- Percentage mass of an element in a compound

$$\frac{\text{Mass of the element in compound}}{\text{Total mass of compound}} \times 100$$

### Example Question:

Find the percentage mass of oxygen in magnesium oxide (MgO).

$A_r$  of magnesium = 24       $A_r$  of oxygen = 16

$M_r$  of MgO = 24 + 16 = 40

$$\begin{array}{r}
 \% \text{ mass} = \frac{A_r}{M_r} = \frac{16}{40} = 0.4 \times 100 = 40\% \\
 \text{X 100 to make a \%} \quad \quad \quad \text{40\% of the mass of MgO is oxygen}
 \end{array}$$

## Mass Changes

- Mass is always conserved in a reaction.
- Sometimes it may seem like the mass has increased/decreased.
- If a **reactant** is a gas – mass may **increase**.

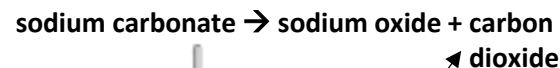


Oxygen is in the air before it combines with magnesium – you cannot find the mass of oxygen on the balance.

It will look like the mass has increased when it is re-weighed at the end.



- If a **product** is a gas and the gas is able to escape the system – mass will **decrease**.

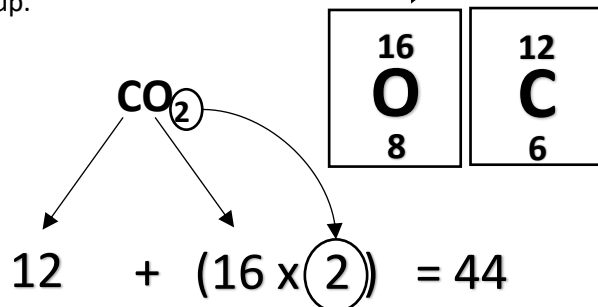


It will look like the mass has decreased as some of the atoms have been given off as gas and have escaped – so cannot be re-weighed.

## Atomic mass ( $A_r$ ) and Relative Formula Mass ( $M_r$ )

- Atomic mass ( $A_r$ ) is the mass number – ie the mass of one atom
- Relative formula mass ( $M_r$ ) = all the **relative atomic masses ( $A_r$ )** of the atoms in a compound or molecule added up.

Example



## The Mole (HT only)

- Avogadro constant –  $6.02 \times 10^{23}$
- One mole contains  $6.02 \times 10^{23}$  atoms or molecules
- The mass, in g, of one mole is the  $A_r$  (if an element) or  $M_r$  if a compound or molecular element

Iron has a  $A_r$  of 56, so 1 mole of iron is 56 g and contains  $6.02 \times 10^{23}$  atoms of iron

Ammonia ( $\text{NH}_3$ ) has an  $M_r$  of 17, so 1 mole of ammonia has a mass of 17g. and contains  $6.02 \times 10^{23}$  molecules of ammonia



# C3 – Quantitative Chemistry

## Concentrations of Solutions

- Concentration = mass of dissolved substance in specific volume (g dm<sup>3</sup>)
- More substance dissolved = more concentrated solution

$$\text{Concentration} = \frac{\text{mass}}{\text{volume}}$$

(g/dm<sup>3</sup>)      (g)      (dm<sup>3</sup>)

Can be rearranged to find mass dissolved:

$$\text{mass} = \text{concentration} \times \text{volume}$$

(g)      (g/dm<sup>3</sup>)      (dm<sup>3</sup>)

$$1000\text{cm}^3 = 1\text{dm}^3$$

$$\text{cm}^3 \rightarrow \text{dm}^3 = \text{divide by } 1000.$$

## Calculating mass in a given volume

If you have a known volume of a solution of known concentration then you can calculate the mass of dissolved solid.

E.g Calculate the mass of dissolved solid in 25cm<sup>3</sup> of a 96g/dm<sup>3</sup> solution

96g/dm<sup>3</sup> means 96g in every 1000cm<sup>3</sup>

Do the same to the other side (÷40)

↓  
2.4g

↓  
25cm<sup>3</sup>

How do we get from 1000 to 25? (÷40)

## Moles and Equations (HT only)

- You can use moles to help you write balanced symbol equations.

### Example Question

18.4g of Sodium reacted with 6.4g of oxygen to give 24.8g sodium oxide. Use the masses to write the balanced equation.

Step	Example
Write the equation for the reaction (unbalanced)	$\text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}$
write down the mass or % given in the question	$18.4 + 6.4 \rightarrow 24.8$
Write the mass of one mole of each element or compound	23    32    62 (e.g $18.4 \div 23$ )
Divide the mass given in question by the mass of one mole	0.8    0.2    0.4
Turn the answers into whole number simple ratio	8        2        4 (cancel down) 4        1        2
Put the numbers into the equation	$4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}$

## Calculating reacting masses (HT)

### Example Question

Calculate the mass of calcium needed to make 11.2g Calcium oxide

Step	Calculation
Write the balanced equation	$2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO}$
Write the masses of each substance	$80 + 32 \rightarrow 112$
Write down the given mass in the question.	11.2
Work out the 'scale' factor (ie what did you have to do to the original number to get to the desired mass	÷ 10
Do the same to the other side	8g

## Limiting Reactants (HT only)

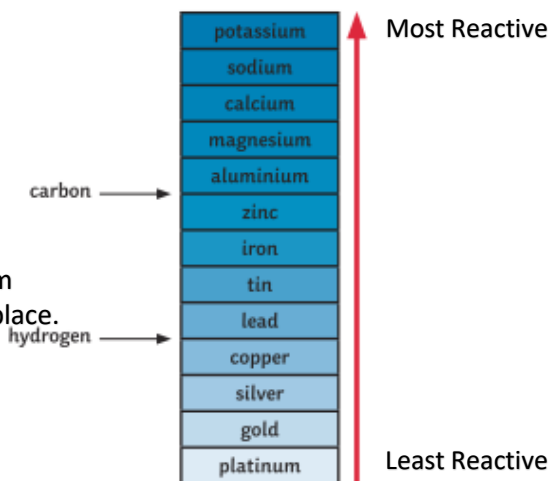
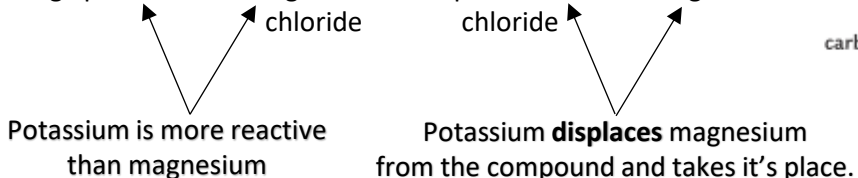
- If one reactant runs out before the other, then the reaction will stop.
- The reactant that runs out first in a reaction is known as the limiting reactant.

# C4 – Chemical Changes

## The Reactivity Series

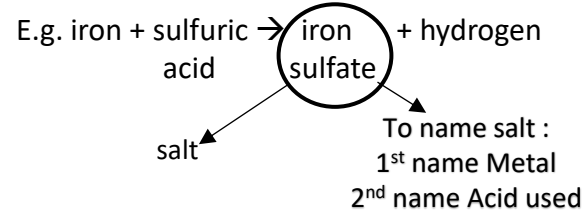
- A more reactive metal will replace a less reactive metal in a compound (**displacement**)

- e.g. potassium + magnesium chloride → potassium chloride + magnesium



## Reactions of acids with metals

- Metal + acid → salt + hydrogen



### Naming Salts

Acid used	Salt produced
Hydrochloric	Chloride
Sulfuric	Sulfate
Nitric	Nitrate

## Extraction of Metals

- Extraction = remove metal from an ore or a compound.

**Ore** = a rock containing enough metal to make extracting metal worthwhile.

### How to extract metals:

**Less reactive than carbon** – reduction with carbon

Reduction = loss of oxygen

e.g. iron oxide + carbon → iron + carbon dioxide

Oxygen has been removed to extract iron.

Carbon and the oxygen removed from the iron react to make carbon dioxide

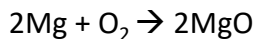
**More reactive than carbon** – electrolysis is used.

- Some metals are found in **native** form (not reacted, so in element form) – usually platinum and gold as **very unreactive**.

## Reaction of metals with oxygen

- Metal + oxygen → metal oxide

e.g. magnesium + oxygen → magnesium oxide



Oxidation reaction as metal gained oxygen

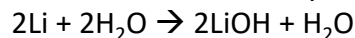
- Oxidation = gaining oxygen
- Reduction = losing oxygen

## Reaction of metals with water

- Most metals don't react well with water
- Group 1 and group 2 react to form alkalis

- Metal + water → metal hydroxide + hydrogen

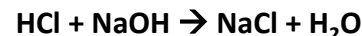
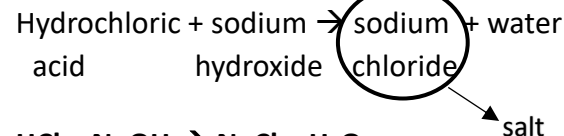
e.g. lithium + water → lithium hydroxide + hydrogen



Metal hydroxides are alkaline

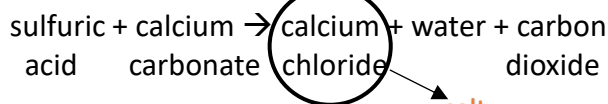
## Reactions of acids with alkalis

- Acid + alkali → salt + water (neutralisation)



## Reactions of acids with carbonates

- Acid + carbonate → salt + water + carbon dioxide



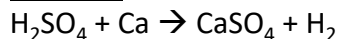
# C4 – Chemical Changes

## Redox Reactions (HT only)

- Redox = reduction and oxidation takes place at same time in a reaction.

- Metal + acid = redox reaction

### Example



Ionic equation:  $2\text{H}^+ + \text{Ca} \rightarrow \text{Ca}^{2+} + \text{H}_2$  Lost 2 electrons (oxidation)

Half equation 1:  $\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$

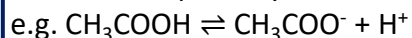
Half equation 2:  $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$  Gained 2 electrons (reduction)

## Strong/Weak Acids (HT only)

**Strong acid** = completely dissociates in a solution  
e.g.  $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$

Examples = nitric acid and sulfuric acid

**Weak acid** = partially dissociates in solution.



$\rightleftharpoons$  = reversible reaction

Hasn't fully turned into ions – only partially

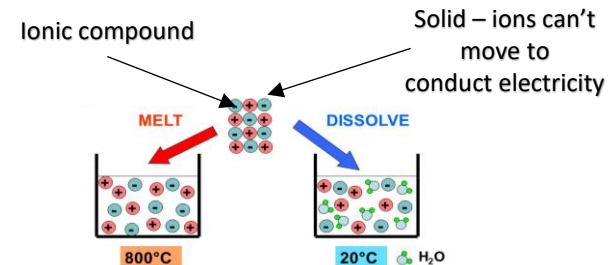
**Concentration** = how much is dissolved in every  $\text{cm}^3$

**Strong/weak** = how well it ionises

As **pH** decreases by 1 unit, **hydrogen ion concentration** of solution increases by factor of 10

## Electrolysis

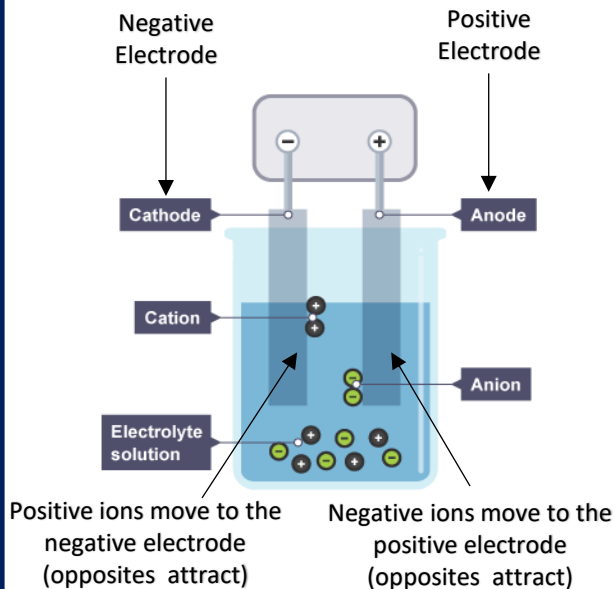
- **Splitting up a compound** using electricity.
- Used to extract metals from compounds, purify metals (eg copper)



- Must be **molten** or **aqueous** (dissolved in water) to allow **ions** to **move** to the electrodes

## The Process of Electrolysis

Two **electrodes** – made of **inert** material (doesn't react)



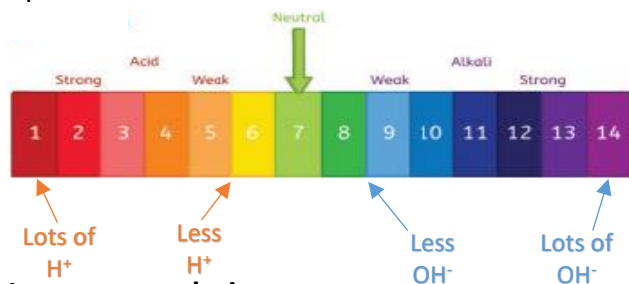
## pH Scale

- Shows how acidic or alkaline solution is.

- pH 1-6 = acid

- pH 7 = neutral

- pH 8-14 = alkali

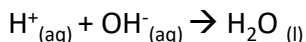


**In aqueous solutions:**

Acids – produce  $\text{H}^+$  ions

Alkalis – produce  $\text{OH}^-$  ions

**In neutralisation reactions:**

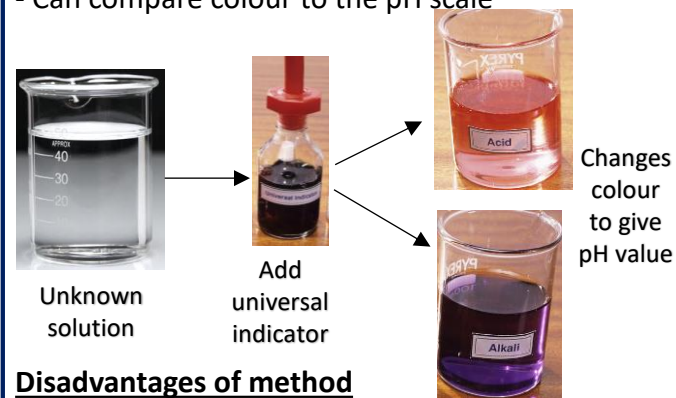


## Measuring pH of a solution

- Can use **universal indicator**

- Gives the solution a colour

- Can compare colour to the pH scale



## Disadvantages of method

- Colour is **subjective** – different people may see different colours

- Doesn't give an exact pH number (could use **pH probe** to make more **accurate**).

## C4 – Chemical Changes – Required Practical – Preparation of soluble salts

### Aim

Prepare a pure, dry sample of a soluble salt from an insoluble **oxide or carbonate**.

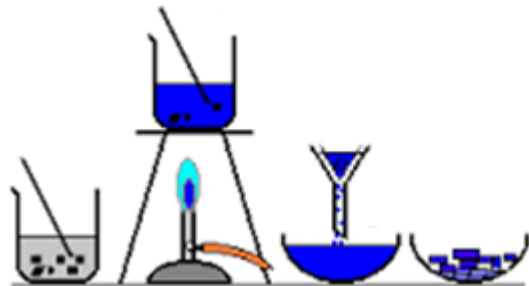
### Equipment

- Beaker
- Measuring cylinder
- Bunsen burner and safety mat
- Filter funnel and filter paper
- Named acid (e.g. hydrochloric acid)
- Metal oxide or carbonate.
- Spatula
- Glass stirring rod

Change method  
depending on reactants in  
the question.

### Method (example copper oxide and sulfuric acid to make copper sulfate)

1. Using measuring cylinder – 20cm<sup>3</sup> **sulfuric acid** → beaker
2. Warm the acid gently (not boiling)
3. Using spatula add **copper oxide** to the acid and stir
4. Keep adding until no more oxide will dissolve (excess).
5. Using a filter funnel and filter paper – filter excess copper oxide.
6. Evaporate some of the filtrate using a water bath.
7. Pour remaining filtrate into an evaporating basin – leave overnight to evaporate water
8. Pat the crystals dry.



### Common questions

**Q1)** Why do you heat the acid before adding the oxide?

**A1)** To speed up the reaction (particles have more energy to react).

**Q2)** Why is the oxide added in excess?

**A2)** To make sure that all the acid has been neutralised.

**Q3)** Why is the solution filtered?

**A3)** Remove any unreacted, excess solid.

**Q4)** Why is the solution left overnight in a warm, dry place?

**A4)** To evaporate excess water, to form crystals (crystallise).

**Q5)** Name 2 safety precautions you should take during this practical.

**A5)** Safety goggles and allow equipment to cool before putting away

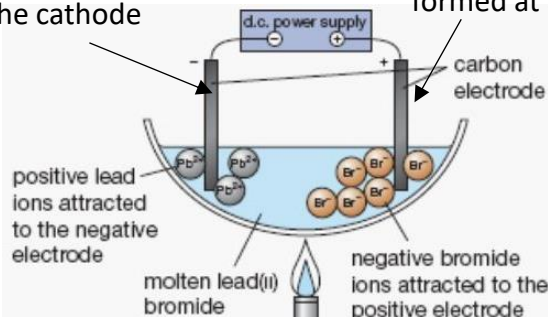
# C4 – Chemical Changes

## Electrolysis of Molten Ionic Compounds

- Molten** = melted so ions can move.
- Metal = produced at **anode**
  - Non-metal = produced at **cathode**

### Example: Lead Bromide - $PbBr_2$

Lead forms at the cathode  
Bromine gas is formed at anode



## Using Electrolysis to Extract Metals

- Used if metal is **too reactive** to be extracted by reduction with carbon.
- Requires **large amount of energy** to melt the compound and produce electrical current. (**expensive**)

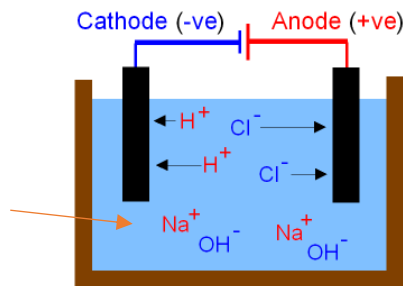
### Example: Aluminium Oxide

- **Cryolite** is added – reduces the melting point (less energy needed – less expensive)
- **Carbon** used as positive electrode – needs to be replaced constantly as **oxygen** will react with it to produce  $CO_2$  – it will degrade.

## Electrolysis of Aqueous Solutions

- Compound is dissolved in water so ions can move.

When aqueous –  $H^+$  and  $OH^-$  (from  $H_2O$ ) are also present along with the two ions from the compound.



- Only **one** ion is discharged at each electrode.

**Anode** – Non-metal or oxygen

**Cathode** – Metal or hydrogen

### Rules

+ ANODE Attracts – ions ('Anions')	- CATHODE Attracts + ions ('Cations')
If – ions are group 7 i.e. chloride $Cl^-$ bromide $Br^-$ iodide $I^-$ Then the groups 7 element is produced as a gas	If + ions (metals) are <b>MORE REACTIVE</b> than hydrogen <b>K, Na, Ca, Mg, Zn, Fe</b> Then <b>HYDROGEN</b> is produced
If – ions are <b>NOT</b> Group 7 Eg sulphate $SO_4^{2-}$ nitrate $NO_3^-$ carbonate $CO_3^{2-}$ <b>OXYGEN</b> is produced.	If + ions (metals) are <b>LESS REACTIVE</b> than hydrogen <b>Cu, Ag, Au</b> Then the <b>METAL</b> is produced

### Examples

Solution	Product at cathode	Product at anode
Potassium chloride	Hydrogen – because K is more reactive than H	Chlorine – as it is a halogen
Copper sulfate	Copper – as copper is less reactive than H	Oxygen – as there is no halogen

## Half-Equations at Electrodes (HT only)

During electrolysis:

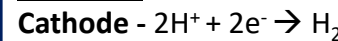
**Cathode** – positive ions **gain** electrons (**reduction**)

**Anode** – negative ions **lose** electrons (**oxidation**)

- Ions become **discharged** (lose their charge) at the electrodes to form the atoms again.

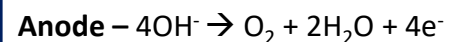
- Reactions at electrodes can be represented by half equations.

### Examples



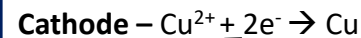
Gained 2 electrons (reduction)

molecules of hydrogen gas produced



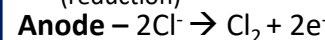
molecules of oxygen produced

Lost electrons (oxidation)



Gained electrons (reduction)

Copper atoms are formed at the cathode



chlorine molecules are formed

Lost electrons (oxidation)

# C5 – Energy Changes

## Exothermic Reactions

- Energy transferred to the surroundings
- Temperature of the reaction mixture **increases**
- This energy is transferred **to** the surroundings

Examples include:

- Hand warmers
- Combustion reactions
- Respiration
- Neutralisation reactions
- Self-heating cans.



Exothermic

## Endothermic Reactions

- Energy absorbed from the surroundings
- Temperature of reaction mixture often **decreases**
- Energy is transferred **from** the surroundings

Examples include:



- Ice packs (injuries)
- Reaction of citric acid and sodium hydrogen carbonate
- Thermal decomposition of calcium carbonate



Endothermic

## Energy change of reactions (HT)

During a reaction:

- Energy is **absorbed** in order to **break** bonds in the reactants 
- Energy is **released** when bonds are **made** in the products. 

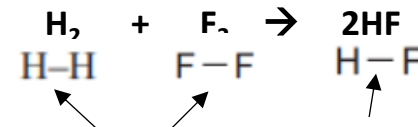
**Bond energy** = the amount of energy that is released when a bond is made or that is needed to break a bond

## Calculating energy changes (HT)

Overall energy change = difference between energy needed to break bonds and the energy released when bonds formed.

To calculate energy change :

Energy change = bonds broken – bonds formed



Bond	Bond Energy / $\text{kJ mol}^{-1}$
F—F	158
H—H	436
H—F	568

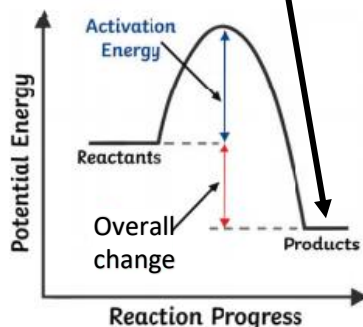
$$\begin{array}{l|l} \text{Bonds broken} = & \text{Bonds formed} \\ 436 + 158 & 2 \times 568 \\ 593 & 1136 \end{array}$$

$$\text{Overall energy change} = 593 - 1136 = -543 \text{ kJ/mol Exothermic}$$

More energy is released in bond making than is required for bond breaking.

## Reaction Profiles – Exothermic

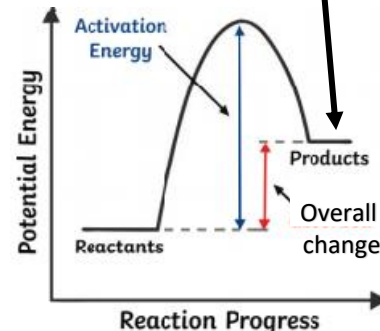
- Energy level diagrams show **difference in energy** between reactants and products.
- Exothermic = Energy of products is **lower than** reactants (energy is released)
- **Activation Energy** = minimum amount of energy needed to start the reaction.
- **Energy change** = the difference in energy between reactants and products.



You may need to draw and label this in the exam!

## Reaction Profiles – Endothermic

- Energy level diagrams show **difference in energy** between reactants and products.
- Endothermic = Energy of products is **higher than** reactants (energy is absorbed)
- **Activation Energy** = minimum amount of energy needed to start the reaction
- **Energy change** = the difference in energy between reactants and products.



You may need to draw and label this in the exam!

## C5 – Energy Changes – Required Practical – Temperature Changes

### Hypothesis

The energy change in the reaction between acid and alkali depends on the volume of alkali added.

### Equipment

- Polystyrene cup and lid
- Thermometer
- 250cm<sup>3</sup> beaker
- Measuring cylinder
- Liquid reactants



### Method (example for hydrochloric acid and sodium hydroxide)

1. Using measuring cylinder to measure 30cm<sup>3</sup> hydrochloric acid and put in polystyrene cup
2. Stand cup inside beaker to make stable.
3. Use a thermometer to measure the temperature of acid and record.
4. Using measuring cylinder – 5cm<sup>3</sup> sodium hydroxide → polystyrene cup
5. Fit the lid and gently stir with thermometer through hole.
6. When reading stops on thermometer, record temperature in table.
7. Repeat, each time adding 5cm<sup>3</sup> more sodium hydroxide up to a maximum of 40cm<sup>3</sup>.
8. Calculate the temperature change on each attempt.
9. Repeat the experiment 3 times and calculate a mean temperature change for each volume of sodium hydroxide.

### Variables

**Independent** – Volume of sodium hydroxide

**Dependent** – Temperature change

**Control** – Volume of hydrochloric acid, concentration of acid, concentration of sodium hydroxide

### Common questions

**Q1)** Why do you use a polystyrene cup and lid?

**A1)** Because polystyrene cups are insulators, which reduces heat loss in the experiment, making the results more accurate.

**Q2)** Why should you calculate the temperature change, instead of just using the final temperature?

**A2)** Because the initial (starting) temperature of the acid may have been different.

**Q3)** Why is it important to stir the mixture?

**A3)** To make sure all of the reactants have reacted and to get a uniform temperature.

**Q4)** Why is the experiment conducted 3 times?

**A4)** So that anomalies can be seen and removed and a mean calculated

### **Energy changes could also be investigated using:**

1. Changing the **mass of metal** added to acid and measuring the **temperature increase**
2. Changing the **type of metal** added to acid and measuring the **temperature increase**
3. Dissolving different **masses of potassium nitrate** into water and observing the **temperature decrease**.







## C10 – Using Resources – Required Practical – Analysis and purification of water

### Analysing the pH of Water Samples

- Test pH of each water sample using pH probe or universal indicator.
- Compare to pH chart if using universal indicator

### Analysis the Mass of Dissolved Solids

1) Measure out 50 cm<sup>3</sup> of water sample using measuring cylinder.

2) Take the mass of evaporating basin using top pan balance.

3) Heat the sample in the evaporating basin gently until all liquid evaporates.

4) Let the evaporating basin cool

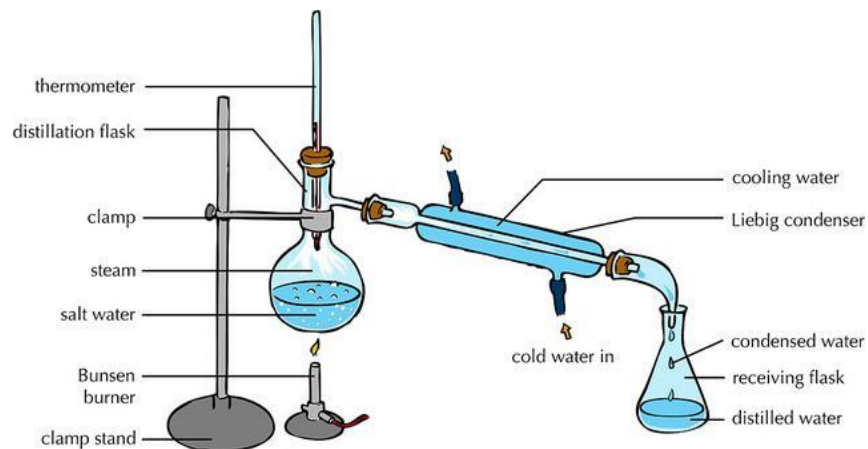
5) Re-take the mass of the evaporating basin.

6) Calculate the mass of the solid left behind by doing: final mass – initial mass.

7) Repeat with different water samples (e.g. rainwater, salt water, spring water)



### Distillation of water Sample



- 1) Set up apparatus as shown in picture with the sample of water in the round bottom flask.
- 2) Heat water sample until it boils gently.
- 3) Water vapour enters the tube at the side (condenser)
- 4) There is cold water surrounding the tube
- 5) The water vapour cools and condenses and collects in the flask.
- 6) The water collected should be **pure**.

## C10 – Using Resources

## Water

### Potable Water

- Water is **essential** for life.
- **Potable water** is water that is safe to drink.
- Potable water is not pure as it contains some dissolved substances.

In the UK – rain water provides water with low levels of dissolved substances that collects in the ground and in lakes and rivers. This is fresh water.

Most potable water is produced by:

- 1) Choosing an appropriate source of fresh water
- 2) Passing the water through filter beds
- 3) Sterilising to kill bacteria

Sterilising agents used for potable water include:

- Chlorine
- Ozone
- Ultraviolet light



### Desalination of Sea Water

- **Potable** water can be made from sea water through desalination.
  - Required a lot of **energy** to **remove salt** in sea water.
- Can be done by:

#### **Distillation**

- Sea water heated until it boils
- Steam is **condensed** to make potable water
- Requires a lot of **energy**

#### **Reverse Osmosis**

- Water put under **high pressure** and passed through **membrane** with tiny holes in.
- Holes allow water through but not salt/ions
- Very **expensive**
- Produces **large volumes** of waste water.

### Waste Water Treatment

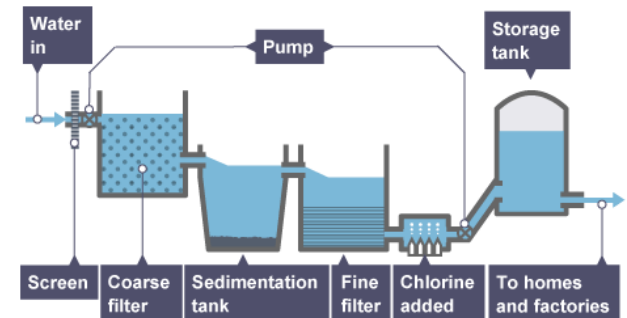
- Waste water needs to be treated before being released back into environment

**Pollutants** can be present in waste water including:

- Human waste contains harmful **bacteria** and nitrogen – can harm aquatic ecosystems.
- Industrial waste can contain **toxic** substances
- Agricultural waste water can contain **fertilisers** or **pesticides** – disrupt ecosystems.

**Sewage treatment involves:**

- 1) Screening and grit removal to remove large particles
- 2) Sedimentation – allows tiny particles to settle – produces sewage sludge and effluent (liquid that remains on the top)
- 3) Sewage sludge is digested anaerobically by specific bacteria
- 4) Effluent is treated with aerobic bacteria to reduce volume of solid waste.



## C10 – Using Resources

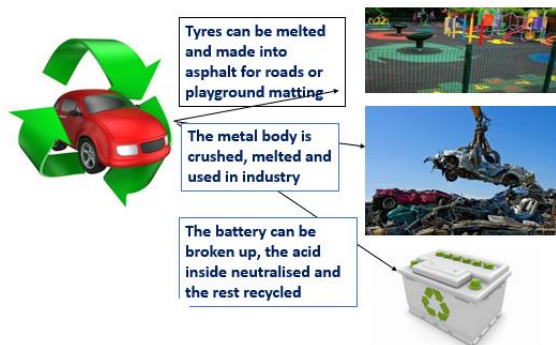
### Reducing the use of resources

Metals, glass, ceramics, building materials and most plastics are produced from limited resources. The energy for the processes involved in making/extracting raw materials also comes from limited resources – e.g. oil. We can reduce the use of limited resources by reducing use, reusing materials and recycling materials at the end of their life.

### Reduce, reuse, recycle.

E.g.

- Glass bottles can be reused.
- Metals can be melted down and recast and so recycled.
- Scrap steel can be added to extracted iron to reduce the amount of iron that has to be extracted in the blast furnace.



### Evaluating methods to reduce, reuse, recycle

Advantages	Disadvantages
Fewer resources such as mines and quarries are needed to extract finite materials	Requires collection and transport of items – involving staff, vehicles and use of fuel
Crude oil does not need to be extracted – avoids high energy costs for fractional distillation etc.	Materials, such as metals, very often have to be separated from other materials first
Less greenhouse gases produced.	Some metals need melting before being reused – energy costs.
Less items in landfill	

### Biological extraction techniques (HT only)

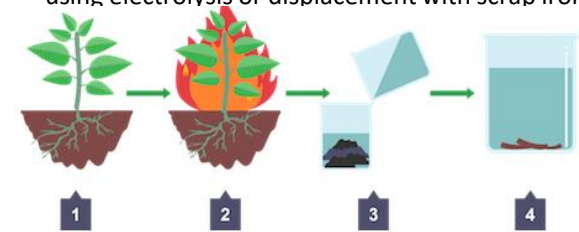
- Earth's supply of metal ores is limited.
- There are fewer sites that give lots of copper (high grade ore sites)
- New ways of extracting from low grade ore sites are:
  - Phytomining
  - Bioleaching

Disadvantage = **slow processes**

Advantage = reduce need for the traditional mining methods of digging, moving and disposing of large amounts of rock.

### Phytomining (HT only)

- 1) Plants are grown on a low-grade ore
- 2) The plants absorb metal ions through their roots
- 3) The plants are harvested and burnt
- 4) Ash left behind contains metal compounds
- 5) Ash is dissolved in acid and copper is extracted using electrolysis or displacement with scrap iron.



### Bioleaching (HT only)

- Uses **bacteria** to produce a solution called **leachate** – contains copper ions.
- The copper can be extracted by using iron to **displace** the copper from the leachate.
- Does not need high temperatures
- Produces **toxic substances** which can damage the environment.
- Iron is cheaper than copper – use of scrap iron is a cost-effective way to produce copper from leachate.
- Can also undergo **electrolysis** to produce copper.

# C10 – Using Resources

## Earth's Resources

We use Earth's resources to provide **warmth, shelter, food and transport.**

E.g.:

- metals from the Earth's crust to build buildings and cars
- Timber and oil to burn for warmth
- Crop plants for food
- Products from crude oil to serve as fuels in cars, trains and planes

**Finite resources** – ones that will run out as they are being used much faster than they can be replaced, e.g. oil

**Renewable resources** – resources that will not run out, e.g. wood, wind etc.

Chemistry plays an important part in finding improvements or alternatives to **current resources.**

Natural	Improved or replaced by....
Wood for furniture	Plastic/polymers
Food crops	Fertilisers/artificially grown foods such as Quorn
Oil for fuel	Ethanol/hydrogen fuel cells
Rubber for tyres	Polymers

## Life-Cycle Assessments (LCA)

- These assess the environmental impact of a product in these stages.

### Stage 1 – extracting raw materials needed to make products.

- Energy cost and effect on habitats of extraction
- Are the raw materials finite/renewable?

### Stage 2 – Manufacturing and packaging product

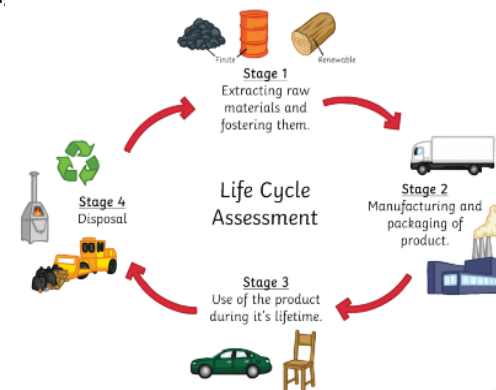
- How much energy and resources are needed?
- What waste products/pollution are released?
- Transportation of goods from factors to user need considering.

### Stage 3 – Use of product during its lifetime

- E.g. a car has a significant impact as needs filled up with petrol which is a finite resource.

### Stage 4 – Disposal at end of product's life.

- 1) Landfill – high environmental impact
- 2) Incineration – burning of product
- 3) Recycling – e.g. batteries contain metals that are harmful to environment – recycling means no new compounds need to be taken out of the ground.



## Example LCA for plastic vs paper bags:

Stage of Life Cycle Assessment	Plastic Bag	Paper Bag
Stage 1 – raw material	Uses finite resource. Process of fractional distillation, cracking and polymerisation all require energy.	Made from trees/recycled paper. Making paper from trees required more energy than recycled paper. Less energy than plastic bags.
Stage 2 – Manufacture	Cheap to make	More expensive to make
Stage 3 – Use	Low environmental impact as can be re-used many times. Much stronger product.	Only be reused a limited number of times – short lifetime.
Stage 4 - disposal	Do not biodegrade easily in landfill.	Paper bags degrade easily in landfill sites.

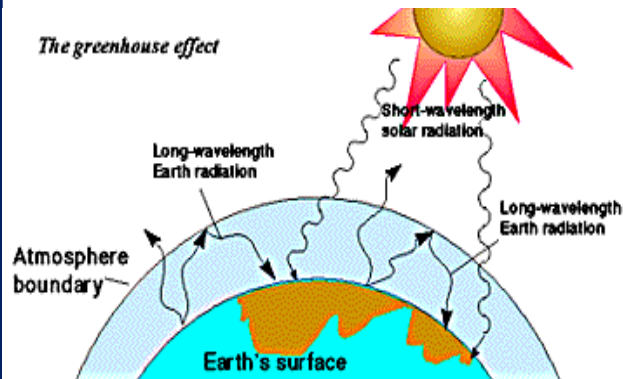
- Different people have different opinions and so depends on who completes the LCA. Bias may be added.
- Some companies may only discuss some of environmental impacts of their product.
- Accurate numerical values should be used where possible – for example to show how much energy has been used.

## C9 – Earth & Atmosphere

### The greenhouse effect

The greenhouse layer is a layer of gases in the atmosphere made of:

- carbon dioxide
- methane
- water vapour



1. Short wavelength infrared radiation from the sun reaches Earth
2. Some energy is absorbed by the Earth
3. Longer wavelength IR is reflected by the Earth
4. Longer wavelength IR cannot get through the greenhouse layer as easily so some is trapped, warming the Earth

**The thicker the layer of gases, the more heat is trapped**

### Global warming

The greenhouse layer is getting thicker, because:

- CO<sub>2</sub> released from fossil fuels to generate electricity
- CO<sub>2</sub> released from fossil fuels in vehicles
- Methane released from cattle
- Methane released from rotting landfill sites

Many scientists believe that human activities are causing the warming of the Earth.

### Potential consequences:

- Melting ice caps
- Loss of habitats for animals and plants
- Damage to coral reefs caused by warmer oceans
- Changes to animal migration patterns
- Extreme weather patterns – more hurricanes, heat waves, droughts, snow and ice
- Difficulty growing crops so reduced food supply

### Carbon footprint



The total amount of CO<sub>2</sub>, CH<sub>4</sub> and water vapour released by of a product or service.

E.g for a concert:

- electricity in performance
- Fossil fuels used by people travelling there
- Plastics used and disposed of in refreshments etc

Carbon footprints can be reduced by recycling, reducing energy use or eating vegetarian diets but this is hard to get people to do.

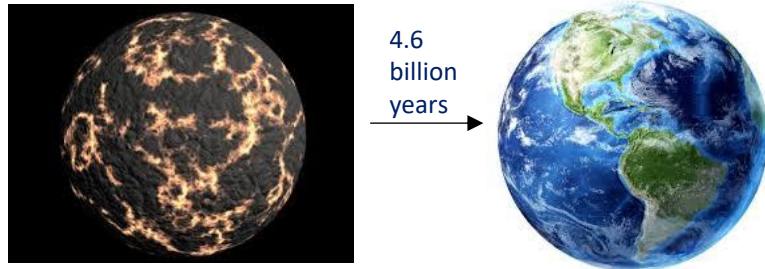
### Pollutants :

Pollutant	Source	Effects
Carbon dioxide	Combustion	Global warming
Carbon monoxide	Incomplete combustion of fuels	Toxic gas, can be fatal
Sulfur dioxide	Traces of sulfur in coal react with oxygen when burned	Acid rain
Nitrogen oxides	Hot engines provide the energy for N <sub>2</sub> to react with O <sub>2</sub>	Acid rain
particulates	Incomplete combustion	Global dimming, breathing problems

# C9 – Earth & Atmosphere

## Early Atmosphere vs modern atmosphere:

Gas	Levels in earth's early atmosphere	Percentage in air today
Nitrogen	None	78
Oxygen	None	21
Others – CO <sub>2</sub> and argon	Very High	1
Water vapour	Very high	Varies – but usually only around 1%
Ammonia	High	None



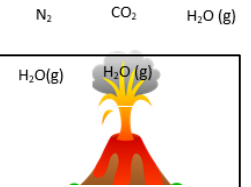
We think that the atmosphere on Earth was once like that of Mars or Venus is today

When Earth was formed it was so hot it was molten on the surface, and the atmosphere was full of toxic gases like methane and ammonia.

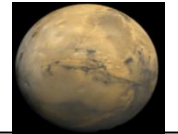
We cannot be sure about exactly what the Earth's early atmosphere as we have no evidence from so long ago

## How did the atmosphere change?

1. Volcanoes released nitrogen, carbon dioxide and water vapour



2. The earth cooled and solidified

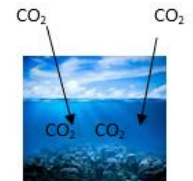


3. Water vapour in the atmosphere condensed and fell as rain



4. Oceans, lakes and rivers formed

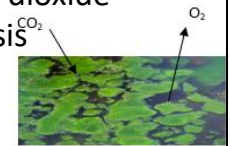
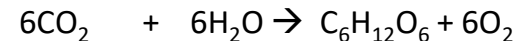
5. Carbon dioxide from the air dissolved in the oceans



6. Some of this reacted to form sedimentary rocks like limestone

7. Algae and then plants evolved, removing carbon dioxide from the air and produced oxygen by photosynthesis

Carbon dioxide + water → glucose + oxygen



8. Many early plants and marine organisms were buried and decayed underground, locking up carbon in fossil fuels like coal (plants) and oil (animals)



## C8 – Chemical Analysis

### Required Practical – Paper Chromatography

Aim: Investigate how paper chromatography can be used to separate and distinguish between coloured substances.

#### Method

- 1) Using a ruler, measure 1cm from bottom of chromatography paper and draw a line across the paper with a **pencil**.
- 2) Using a pipette, drop small spots of each ink onto pencil line (leave a gap so do not merge).
- 3) Pour solvent into a beaker, do not fill solvent above the pencil line on the paper.
- 4) Place chromatograph paper into beaker and allow solvent to move up the paper.
- 5) Remove paper just before solvent reaches top of the paper and leave to dry.
- 6) Calculate  $R_f$  values of all the spots using the equation below:

$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

#### Common questions

**Q1)** Why is a pencil used instead of a pen?

**A1)** Ink in the pen would move up the paper with the substances.

**Q2)** Why do you not fill the solvent above the line?

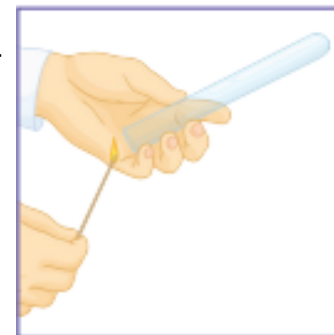
**A2)** Substances would wash off into the solvent instead of rising up the paper

**Q3)** Why might water not work as a solvent?

**A3)** Some substances are **insoluble** in water.

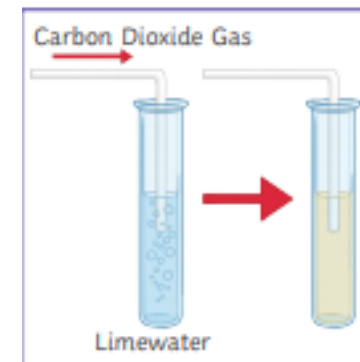
### Identification of the Common Gases

**Test for hydrogen** – Place a **burning** splint at the opening of a test tube. If hydrogen gas is present, it will burn with a **squeaky-pop** sound.



**Test for Oxygen** – Place a **glowing** splint inside a test tube. The splint will **relight** in the presence of oxygen.

**Test for Carbon Dioxide** – Bubble the gas through the lime water – if the gas is carbon dioxide, the limewater turns **cloudy**.



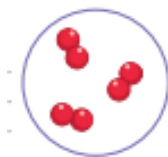
**Test for Chlorine** – **Damp litmus paper** is held over the of gas. If the tube contains chlorine, the litmus paper becomes **bleached** and **turns white**.



## C8 – Chemical Analysis

### Pure substances

Pure = single element or compound – not mixed with any other substance.



### Testing to see if a substance is pure:

- Pure substances have specific melting and boiling points
- Compare your data to a library of known values.

E.g. Water has a boiling point of 100°C, if it is above or below this, it is not pure.

### Formulations

**Formulation** = a mixture that is designed as a useful product.

- Components mixed carefully to get the required **properties**.

Examples of formulations:

- Fuels
- Cleaning agents
- Paints
- Medicines
- Alloys
- Fertilisers
- Food

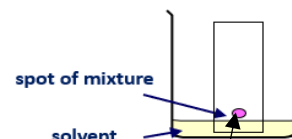


### Chromatography

- Technique used to separate mixtures of **soluble substances**.
- How soluble a substance is determines how far it travels across paper.
- More soluble = travels further (higher up paper)**

### Mobile phase

- **Solvent** is the mobile phase
- The substances dissolve in the solvent
- The solvent then moves through the stationary phase.

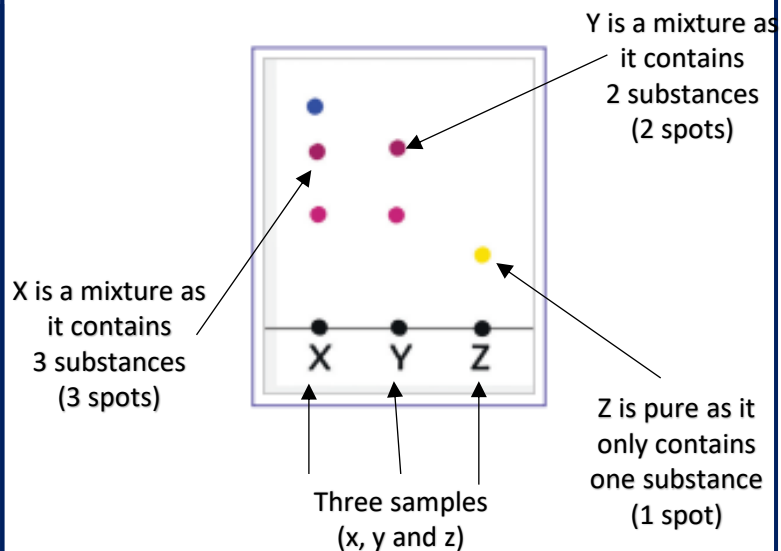


### Stationary phase

- Does not move. The paper is the stationary phase.

**Important** – start line on paper must be drawn in **pencil** as pencil is **insoluble** and **will not run**

The spot and start line must be **above the solvent line** so the colours won't just wash into the solvent in the beaker.



X is a mixture as it contains 3 substances (3 spots)

Y is a mixture as it contains 2 substances (2 spots)

Z is pure as it only contains one substance (1 spot)

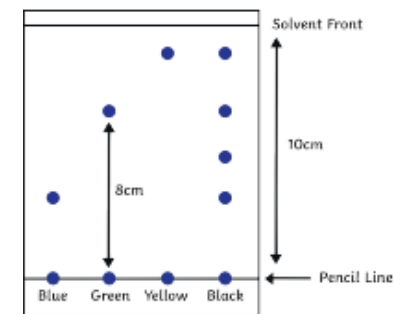
Three samples (x, y and z)

### R<sub>f</sub> Values

This is the ratio of the distance moved by a substance to the distance moved by the compound

$$R_f = \frac{\text{distance travelled by substance}}{\text{distance travelled by solvent}}$$

- Should always be between 0 and 1.
- Each substance has a unique R<sub>f</sub> value.
- Can compare R<sub>f</sub> values to a library of known substances
- Can identify unknown substances.



R<sub>f</sub> value of green:

$$8\text{cm} / 10\text{cm} = 0.8$$

## C7 – Organic Chemistry

### Cracking

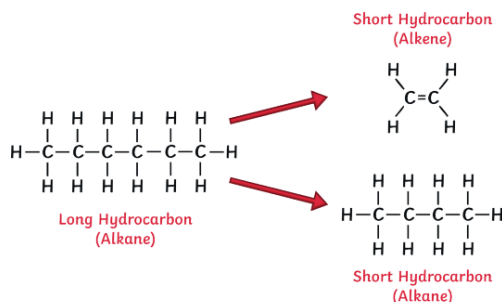
- This is done to solve the problem of having too many long chain hydrocarbons and not enough short ones
- Long hydrocarbons are **broken down** into smaller, more useful hydrocarbons.
- Short chain hydrocarbons are more useful as they are more flammable

**Two types of cracking: catalytic and steam cracking.**

**Catalytic cracking** – needs a **high temperature** and a **catalyst**.

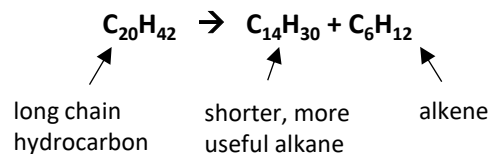
**Steam cracking** – high temperature and steam

- Cracking produces a **short-chain alkane** and an **alkene**.



### Cracking equations

Same number of carbon and hydrogen atoms on both sides of the equation:



### Alkenes

- Alkenes are **unsaturated** hydrocarbons.
- Contain carbon-carbon **double bonds**.

### Test for Alkenes

Use bromine water to test for alkenes.

If an alkene is present, the bromine water turns from orange/brown to colourless.

Alkanes do not react with bromine water.

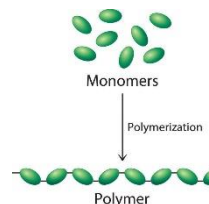


### Uses for alkenes:

- Can be used as fuels
- Can be used as a starting material for other chemicals
- Can be used to make polymers (e.g. plastic)

### Polymers

- Polymers are large molecules made of many repeating units (monomers)
- Alkenes (small molecules) are joined together to make polymers



**Poly(ethene)** – plastic bags/drinks bottles

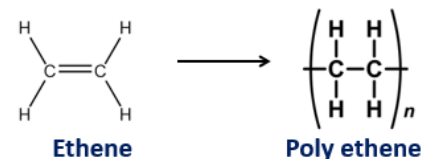
**Poly(propene)** – strong tough plastics

### Drawing and naming polymers

1. Redraw the **monomer given**, but without the double bond. Make sure to copy all other elements exactly.
2. Put brackets around the monomer and extend joining bonds out through the brackets on both sides
3. Add an 'n' at the bottom right of the bracket
4. To name the polymer, you put **poly** in front of the monomer name

E.g.:

Draw and name the polymer made from the monomer ethene:



### Combustion of Hydrocarbons

Combustion means burning.

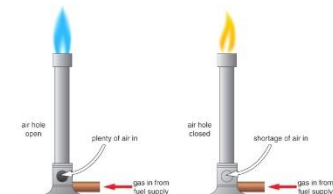
**Complete combustion** - when there is a good supply of **oxygen** for a fuel to burn.

**Fuel + oxygen → carbon dioxide + water**

**Incomplete combustion** - not enough oxygen

Products are **carbon monoxide** and water.

**Carbon monoxide = poisonous gas**



# C7 – Organic Chemistry

## Crude oil

Crude oil = a mixture of **hydrocarbons**.

- It is a **non-renewable resource (fossil fuel)**
- Made from remains of dead sea creatures **compressed** over millions of years

**Hydrocarbons** - molecules containing **hydrogen** and **carbon only**.

Two types of hydrocarbons are **alkanes** and **alkenes**.  
The hydrocarbons in crude oil are mostly alkanes.

## Alkanes

- Alkanes = **saturated** hydrocarbons.
- Held together by **single covalent bonds**.
- General formula =  $C_nH_{2n+2}$
- Have different boiling points – longer the chain, higher the boiling point

**You need to remember the names, and formulas of the first 4 alkanes.**













Name of Alkane	Structural Formula	Molecular Formula
methane	$\begin{array}{c} \text{H} \\   \\ \text{H}-\text{C}-\text{H} \\   \\ \text{H} \end{array}$	$CH_4$
ethane	$\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H}-\text{C}-\text{C}-\text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$	$C_2H_6$
propane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\   \quad   \quad   \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$	$C_3H_8$
butane	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\   \quad   \quad   \quad   \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\   \quad   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$	$C_4H_{10}$

## Fractional Distillation

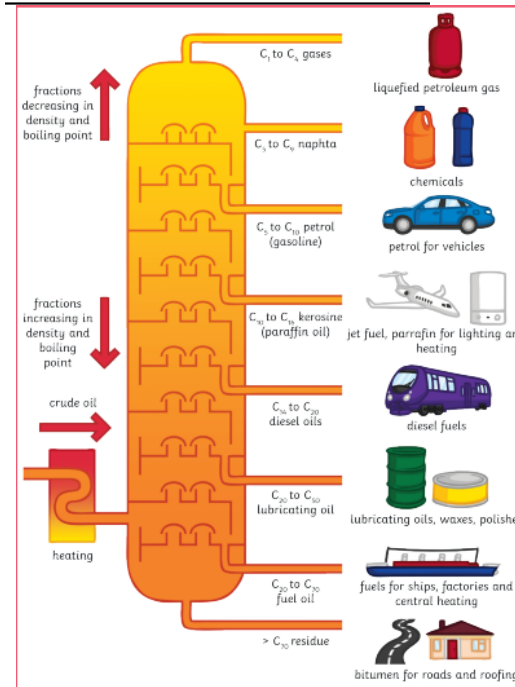
- Used to **separate** the mixtures of hydrocarbons in **crude oil**.

### Steps in Fractional Distillation

1. Crude oil enters **fractioning column** and is heated to boiling point so the hydrocarbons evaporate.
2. It is **cooler** at the **top** of the fractionating column and **hotter** at the **bottom**.
3. Vapours rise up the column and, as they rise, they cool
4. The different hydrocarbons condense at different **boiling points**
5. The different 'fractions' have different **properties**

Short-Chain Molecules	Increasing Chain Length	Long-Chain Molecules
		
As chain length increases, the <b>boiling point</b> of the hydrocarbon chains also increases.		
thin		thick
		
		
<b>Flammability</b> is a measure of how easily a substance burns.		

## Uses of the different fractions



## Supply and demand

Product	Supply in tonnes	Demand in tonnes
petrol	100	300
diesel	200	100
heating oil	250	50

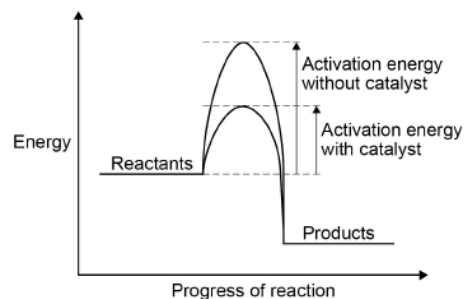
After fractional distillation, we find:

- we have more of the long chain hydrocarbons than we need
- There are not enough short chain hydrocarbons.
- Short chain are more useful as they are more flammable so can be used as fuels.

## C6 – Rate and extent of chemical change

### Catalysts

- Catalysts are substances that speed up chemical reactions without themselves being used up.
- They provide a different pathway for the reaction with a lower activation energy.
- Different reactions require different catalysts.



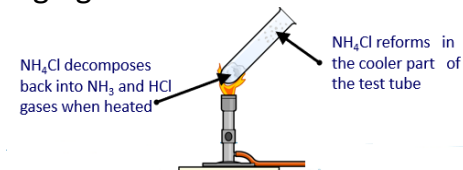
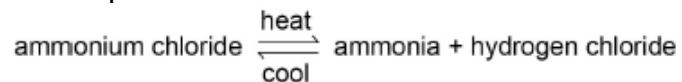
### Reversible reactions

These are reactions in which the products can react to produce the original reactants

They are represented by the symbol  $\rightleftharpoons$

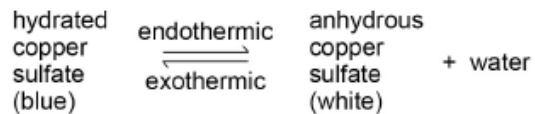
The direction of the reaction can be changed by changing the conditions

For example:



If a reaction is exothermic in one direction, it is endothermic in the opposite direction.

**The same amount of energy** is transferred in each case.



When a reversible reaction takes place in sealed apparatus, then a point occurs when the forward and backward reactions occur at the same rate. This is **equilibrium**

### The effect of changing conditions on equilibrium (HT)

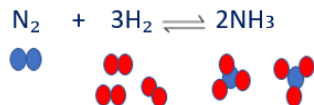
If a system is at equilibrium and a change is made to the conditions, then the system responds to counteract the change.

E.g. – if the temperature is increased, then the system will respond by increasing the rate of the endothermic reaction, to bring the temperature back down

If the concentration of the reactants is increased, then equilibrium will shift right and more products will be made.

In gaseous reactions, a change in pressure will result in equilibrium shifting to the side that restores the pressure.

E.g. :



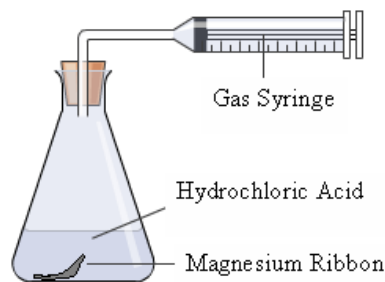
In this reaction, there are 4 moles of gas on the reactants side and only 2 on the product side

If the pressure is increased, equilibrium will shift right as there are fewer moles on the products side, and this will decrease the pressure.

## C6 – Required practical – the effect of concentration on rate of reaction

### Experiment 1

Using volume of gas collected over time as a measure of the rate



**Independent variable:** concentration of HCl

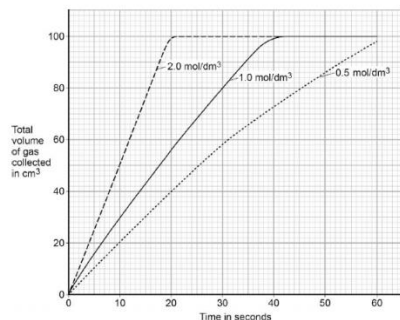
**Dependent variable :** Volume of gas produced / min

**Control variables :** volume of HCl, mass of Mg, temperature of acid

#### Method

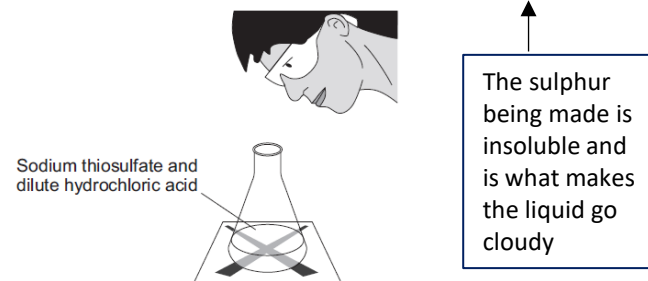
1. Measure 20cm<sup>3</sup> 0.5M HCl into a conical flask.
2. Insert 2 x 2cm pieces of Mg and attach a gas syringe
3. Start a stopwatch and measure the volume of gas collected every 20 seconds until the reaction is over.
4. Repeat using different concentrations of HCl.

An increase in the concentration leads to an increase in the rate of the reaction, but the same volume of product overall



### Experiment 2

Investigating the effect of changing the concentration of HCl on the rate of reaction



**Independent variable:** concentration of HCl

**Dependent variable :** Time taken for the cross to disappear

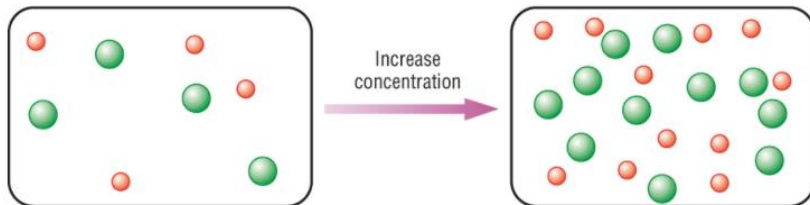
**Control variables :** volume of HCl, volume of sodium thiosulphate, temperature of both solutions, concentration of sodium thiosulphate

#### Method

1. Use a measuring cylinder to put 10 cm<sup>3</sup> sodium thiosulfate solution into the conical flask.
2. Put the conical flask on the black cross.
3. Put 10 cm<sup>3</sup> of 0.5M hydrochloric acid into the 10 cm<sup>3</sup> measuring cylinder.
4. Put this acid into the flask. At the same time swirl the flask gently and start the stopwatch.
5. Look down through the top of the flask. Stop the stopwatch when you can no longer see the cross. Record the time.
6. Repeat steps 1-5 using different concentrations of HCl – 1M, 1.5M, 2M and 2.5M

## C6 – Rate and extent of chemical change

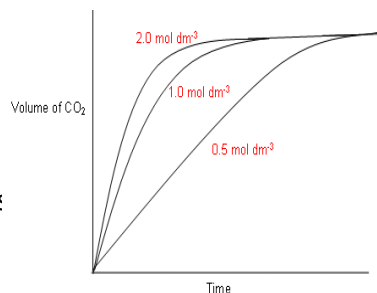
### The effect of concentration



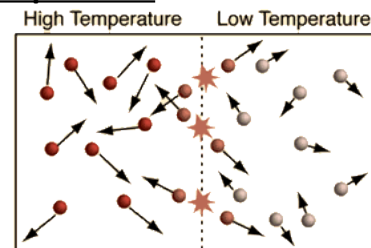
Concentration means number of particles per  $\text{cm}^3$

Increasing the concentration of any of the reactants increases the rate of the reaction

This is because there are more particles per  $\text{cm}^3$  so there are **more frequent collisions**, increasing the rate.

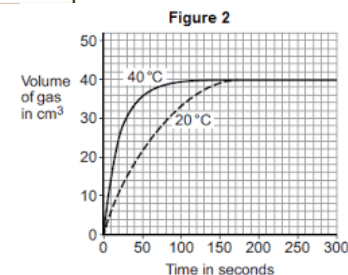


### The effect of temperature



Increasing the temperature of the reactants increases the rate of the reaction.

This is because the particles have more kinetic energy and therefore move faster, so there are **more frequent collisions**, increasing the rate.

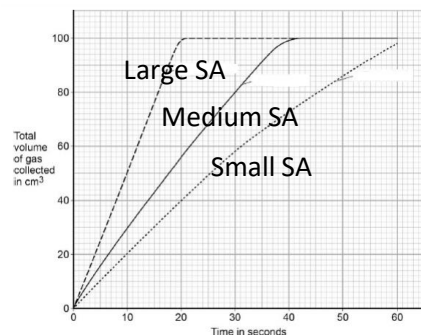


### The effect of surface area

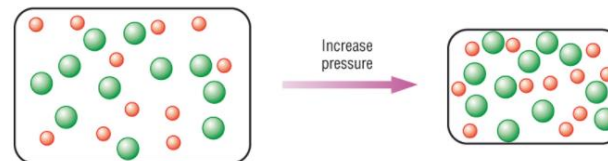


The smaller the pieces of a solid, the higher the surface area. Increasing the surface area of solid reactants increases the rate of reaction.

This is because there is a greater area available for collisions to occur so there are **more frequent collisions**, increasing the rate.

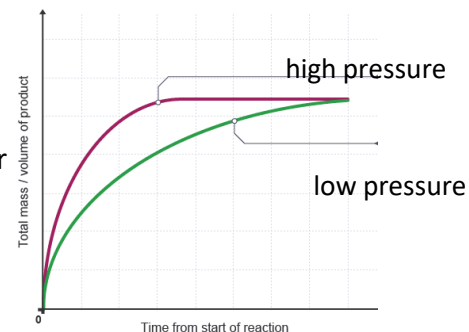


### The effect of pressure



Increasing the pressure of gaseous reactions increases the rate of the reaction.

This is because the same number of particles are now in a smaller volume, so there are **more frequent collisions**, increasing the rate.



**In all cases, the overall amount of product is the SAME, the end point of the reaction is just reached faster**

## C6 – Rate and extent of chemical change

### Rate of reaction.

Measuring the rate of anything always involves a **measurement of time**

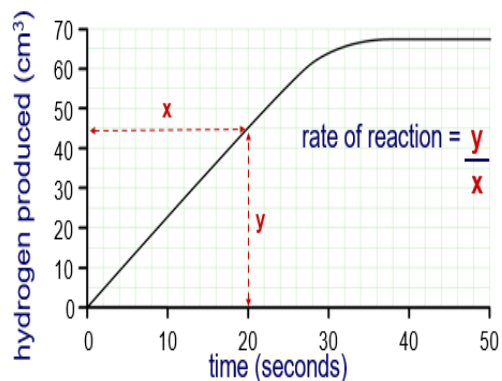
The rate of a chemical reaction can be found using:

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

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

Quantities for reactants or products are measured in **mass in g** or by **volume in cm<sup>3</sup>**

Rate calculations can be done from tables of data or graphs:



Volume of hydrogen produced = 45cm<sup>3</sup>

Time taken = 20 seconds

Rate =  $\frac{45}{20}$  cm<sup>3</sup>/s

20 s

rate = 2.25 cm<sup>3</sup>/s

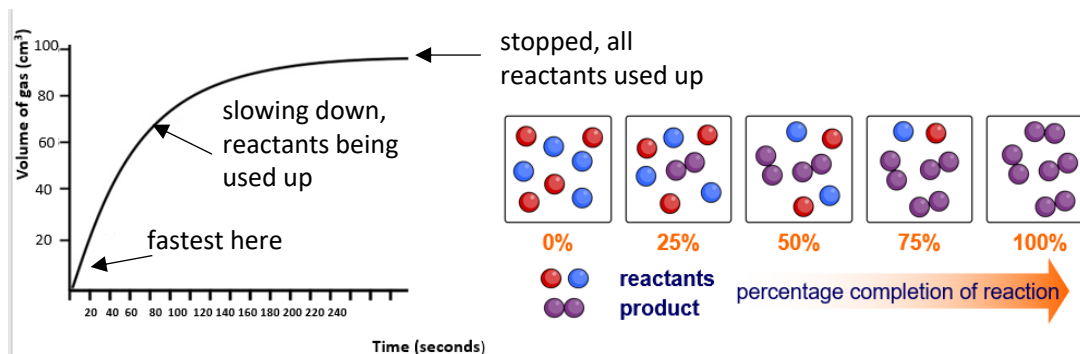
### The progression of a chemical reaction

For a reaction to take place, reactant particles have to collide.

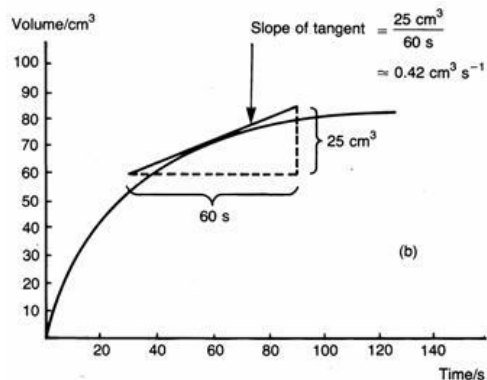
The rate of a reaction depends on the **frequency of collisions** and the **energy with which the particles collide**.

The minimum amount of energy needed to start a reaction is called the **activation energy**.

A reaction is always **fastest at the beginning** and slows down over time as the reactants get used up and the frequency of collisions decreases.



### Using a tangent to calculate rate (HT)



- Draw a line along the point you're interested in. The line should touch the curve at the point given.
- Make a triangle. Try to make the angles either side of the line equal.
- Measure the change in volume and change in time
- Calculate the gradient
- Use units from the axes to determine the units for rate

# Chemistry

## Paper 2 Knowledge

# Organisers

AQA Combined Science (Trilogy)

NAME: \_\_\_\_\_

CLASS: \_\_\_\_\_

TEACHER: \_\_\_\_\_