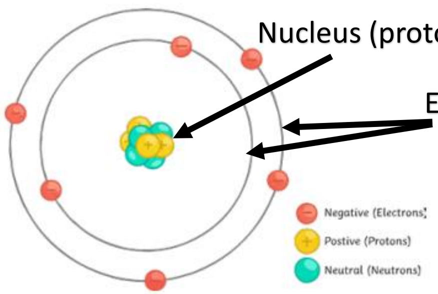
Chemistry
Paper 1 Knowledge Organisers + Questions

AQA Combined Science (Trilogy)

C1 – Atomic Structure and The Periodic Table

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



Nucleus (protons & neutrons)

Electron shells (orbits)

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

Atoms have a radius of about 0.1nm ($1x10^{-10}$ m)

Radius of nucleus = about $1x10^{-14}$ m

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

Atomic number → Number of protons (same as electrons) → Atoms are neutral as number of protons = number of electrons

12
C
6

Relative Atomic Mass (A_r) → Number of protons + neutrons

Number of neutrons =

Atomic Mass – Atomic number
(top) (bottom)

Compounds
- Two or more elements **chemically combined**.
- Formed by chemical reactions
- For example: CO_2 H_2O CH_4 HCl 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..

12 C 6	Same element	13 C 6

Neutrons = 12 – 6 → 6 neutrons

Neutrons = 13 – 6 → 7 neutrons

Both have 6 protons

Different number of neutrons

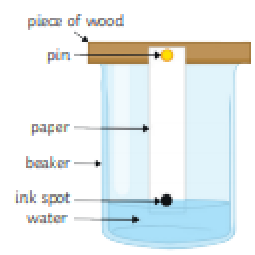
Chemical Equations
- Shown by using a **word equation**.
e.g. magnesium + oxygen → magnesium oxide

Left of the arrow = **reactants**
Right of the arrow = **products.**

- Also can be shown by a **symbol equation**
e.g. $2Mg + O_2 \rightarrow 2MgO$

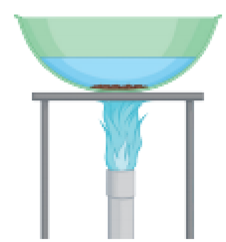
Mixtures and Separation
Mixtures – two or more elements or compounds **not** chemically joined.
This means the different components of the mixture can be separated by physical methods (below)
E.g. air is a mixture mainly made of nitrogen, oxygen and carbon dioxide.

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

1. Name the three subatomic particles.
2. Which two subatomic particles are found in the nucleus of an atom?
3. What is the mass of a proton?
4. What is the radius of an atom?
5. What is the radius of the nucleus of an atom?

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1. Where are elements found?
2. What does the relative atomic mass of an element show?
3. What does the atomic number show?
4. How do you calculate the amount of neutrons?

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1. Define the word compound.
2. Give three examples of compounds.

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1. What is an isotope?

2. Why are the two elements below isotopes? (use the numbers of **subatomic particles**)

$^{14}_{7}\mathrm{N}$ $^{15}_{7}\mathrm{N}$

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1. Where do you find the reactants in a chemical reaction?
2. Where do you find the products in a chemical reaction?

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1. Is air an element, compound or mixture? Why?
2. What is chromatography used to separate?
3. What can be separated using filtration?
4. Give an example of a mixture that can be separated using filtration.
5. What is evaporation used to separate?
6. Give an example of a mixture that can be separated using evaporation.

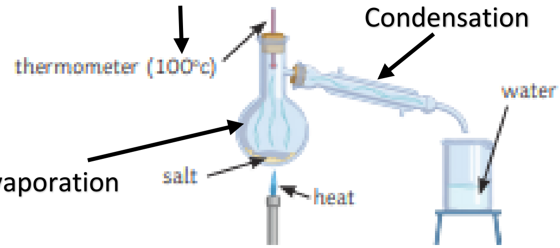
C1 – Atomic Structure and The Periodic Table

Distillation

Simple distillation – separating a liquid from a solution.

Boiling point of water

Condensation



Evaporation

- 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

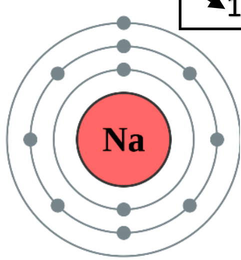
- Electrons are found on shells (orbits) orbiting the nucleus.
- There is a maximum number of electrons allowed on each shell:

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

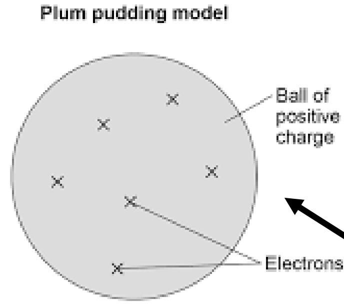
11 electrons

23
Na
11

1st shell = 2
2nd shell = 8
3rd shell = 1



Total = 11 electrons



History of the atom

Scientist	Time	Discovery
John Dalton	Start of the 19th 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.

Differences to nuclear model
- Ball of positive charge (no protons)
- No nucleus
- No neutrons
- Evenly distributed mass

Rutherford tested the plum pudding model

Rutherford's scattering experiment

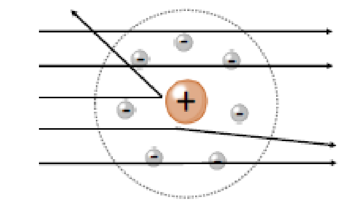
What happened ?

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

1. What two changes of state occur in distillation?
2. What temperature would the thermometer show when distilling salt and water?
3. Why does the water vapour condense in the condenser?

1. Who suggested the plum pudding model?
2. State three differences between the nuclear model and the plum pudding model.
3. What did Niels Bohr discover?
4. What did James Chadwick discover?
5. Put the particles into order of discovery:
   proton electron neutron

1. Where are electrons found?
2. How many electrons can be placed in the first, second and third shells?
3. Which number on the element shows the number of electrons?

1. Who conducted the scattering experiment?
2. What was fired at gold leaf during the scattering experiment?
3. Only a tiny number of the alpha particles were deflected, what did this show about the atom?
4. Some particles went straight through, what did this show about the atom?

C1 – Atomic Structure and The Periodic Table

Development of the Periodic Table

John Newlands – Law of Octaves

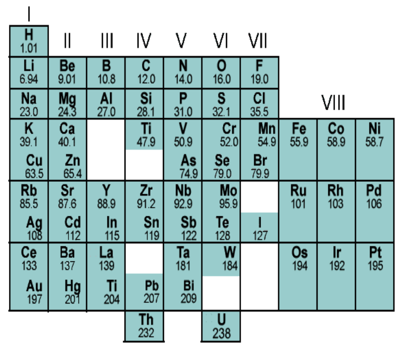
- Elements ordered by **atomic weight**.
- Noticed a pattern with every eighth element.
- Some elements placed inappropriately – metals and non-metals grouped together.
- Rejected by other scientists.

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**.



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**.

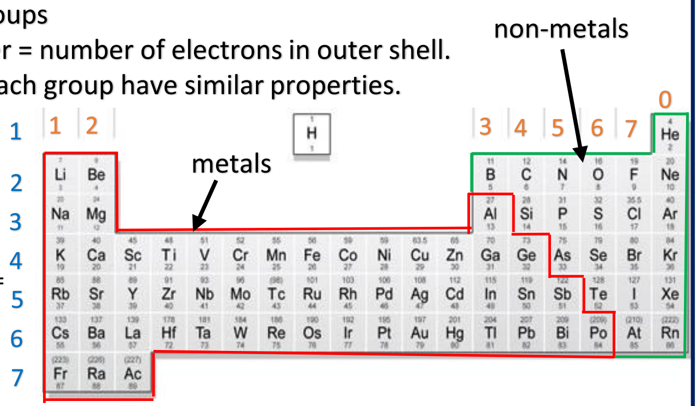
Columns = groups

Group number = number of electrons in outer shell.

Elements in each group have similar properties.

Rows = periods

Period number = number of electron shells the atom has.



Group 0 (Noble Gases)

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

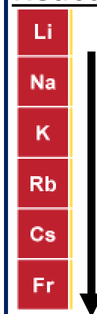
As you go down…
- Boiling point increases
- More electron shells
- Bigger atoms
- More intermolecular forces
- More energy needed to break forces.

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

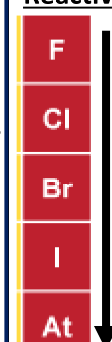


As you go down the group…
- Elements are more reactive because:
- More electron shells
- Outer electron = further from nucleus and more shielded by the other shells
- The electrostatic force of attraction between outer electron and nucleus is weaker
- Easier for outer electron to be lost

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_2, F_2)
  - Non-metals.
  - A more reactive halogen can replace a less reactive halogen in a reaction **(displacement)**

Reactivity of Group 7



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.

C1 – Atomic Structure and The Periodic Table

1. Who created the 'Law of Octaves'?
2. How were the elements ordered in old versions of the periodic table?
3. How did Dimitri Mendeleev order his elements?
4. Why did Mendeleev leave gaps in his periodic table?
5. The knowledge of what eventually explained why elements could not be ordered by atomic weight?

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1. How are elements ordered in the modern periodic table?
2. Groups are rows or columns?
3. What does group number show?
4. What does period number show?

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1. What are elements in group 0 known as?
2. Why are these elements unreactive?
3. What happens to boiling point as you go down group 0?

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1. State 2 properties of Group 1 metals.
2. Why are they known as the alkali metals?
3. Are they reactive or unreactive?
4. As you go down the group, what happens to the reactivity of elements?
5. Explain your answer to Q4.

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1. How many electrons do the halogens have in the outer shell?
2. What type of element are they?
3. State the trend in reactivity as you go down group 7.
4. Explain your answer to Q4.

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

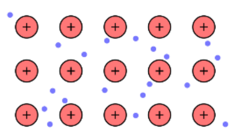
Group	Ions	Example
1	+1	$Li \rightarrow Li^{+} + e^{-}$
2	+2	$Ca \rightarrow Ca^{2+} + 2e^{-}$
6	-2	$O + 2e^{-} \rightarrow O^{2-}$
7	-1	$Br + e^{-} \rightarrow Br^{-}$

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.

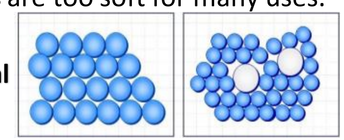


- Strong **electrostatic forces of attraction** between positive ions and negative electrons.

Alloys

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

Pure Metal

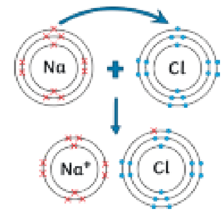


Alloy

Pure Metal	Alloy
• Atoms same size	• Different sized atoms
• Layers slide	• Layers cannot slide
• Softer	• Stronger

Ionic Bonding

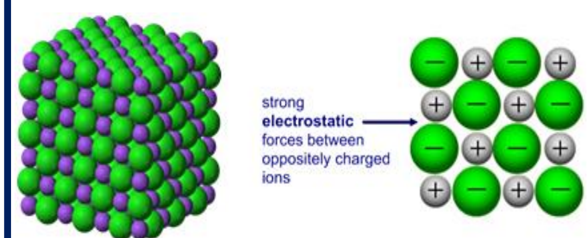
- Between a metal and non-metal.
- Metals give electrons to non-metals so both have a full outer shell.
- **Electrostatic force of attraction** between positive and negative ions.



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

Ionic compounds

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



Properties of Ionic Compounds

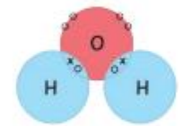
- **High melting point** – lots of energy needed to overcome electrostatic forces.
- **High boiling point**
- **Cannot conduct electricity as solid** – ions cannot move
- **Conducts electricity when molten or dissolved** – ions are free to move.

Covalent Bonding

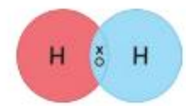
- **Covalent bonding** = sharing a pair or pairs of electrons for a full outer shell.
- Between **non-metals only.**

Dot and cross diagrams

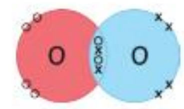
- Show the bonding in simple molecules.
- Uses the outer shell of the atoms
- Crosses and dots used to show electrons
- You should be able to draw the following:



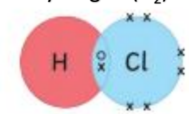
Water (H_2O)



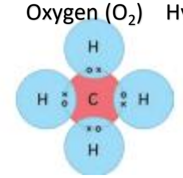
Hydrogen (H_2)



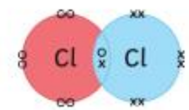
Oxygen (O_2)



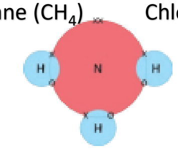
Hydrogen chloride (HCl)



Methane (CH_4)



Chlorine (Cl_2)



Ammonia (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

1. What is an ion?
2. What happens to form a positive ion?
3. What happens to form a negative ion?
4. What type of ions are formed by:
    1. metals
    2. non-metals

1. What are metal ions surrounded by?
2. Name the type of attraction between the electrons and ions.
3. Why do metals conduct electricity?
4. What is an alloy?
5. Why are pure metals too soft for some uses?
6. Why are alloys stronger than pure metals?

1. Ionic bonding happens between..
2. What do metals give to non-metals?
3. What type of attraction is between the positive and negative ions?
4. What structure do ionic compounds form?
5. What are the melting points of ionic compounds like?
6. Why can solid ionic compounds **not** conduct electricity?
7. When can ionic compounds conduct electricity?

1. What is covalent bonding?
2. What type of atoms does covalent bonding happen between?
3. Draw dot and cross diagrams for the following:

Water (H_2O)

Methane (CH_4)

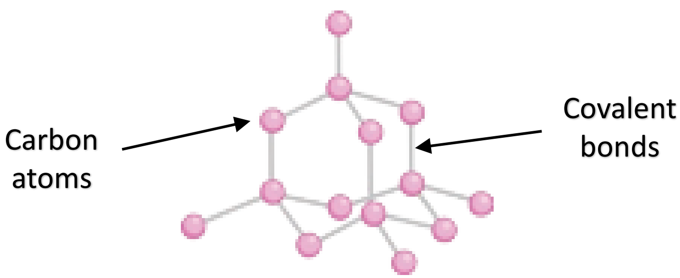
Oxygen (O_2)

5. Do simple covalent molecules have a high/low melting point?

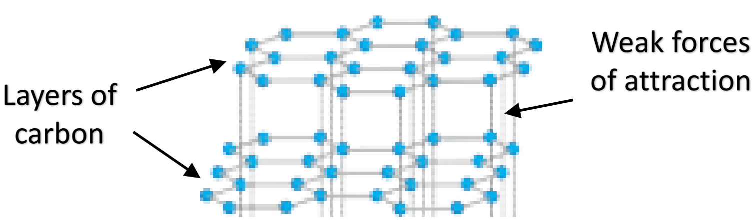
6. Why is this?

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.

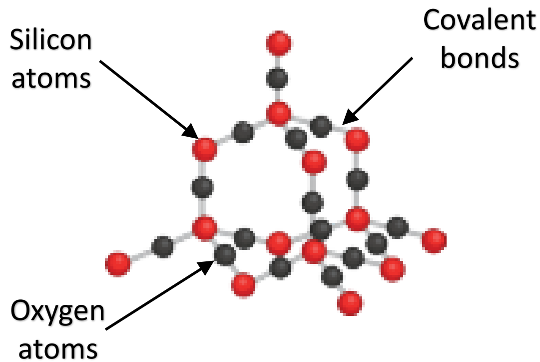


Giant Covalent Structure – Graphite
- Layers of **carbon** arranged in **hexagons.**
- Each carbon bonded to **three** other carbons.
- Leaves **one delocalised electron** → moves to carry electrical charge **throughout structure.**



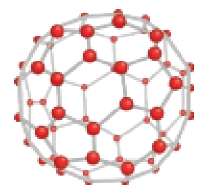
- Layers held together by **weak forces**
- Layers can **slide** over each other easily
- Makes graphite **soft/slippery** → good lubricant.
- Has **high melting point** as has many strong covalent bonds.

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.



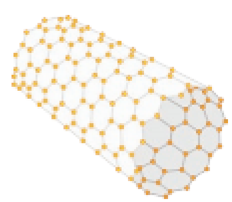
Fullerenes and Nanotubes
- Molecules of carbon shaped into hollow tubes or balls.
- Used to **deliver drugs into body**

Buckminsterfullerene
Formula = C_{60} →



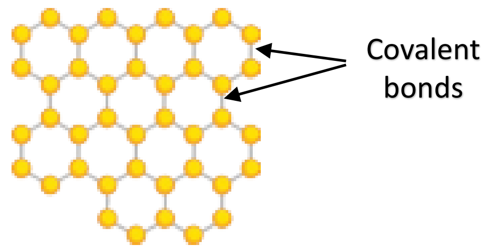
- **Carbon nanotubes** = long narrow tubes
- Can conduct electricity

- Can strengthen materials without adding weight.



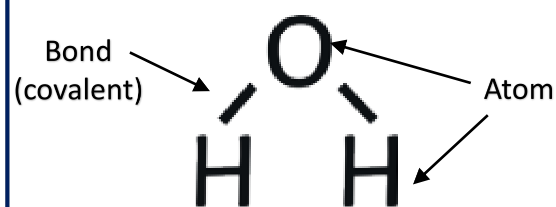
- Used in electronics and nanotechnology.

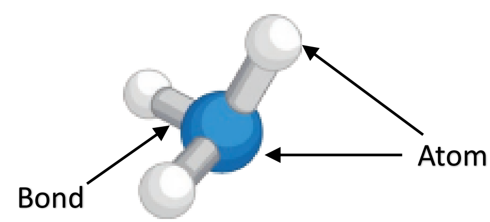
Graphene
- Graphene = one layer of graphite.
- Very strong → lots of strong covalent bonds.



- 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.





3D model

C2 – Bonding, structure, and the properties of matter

1. How many bonds do each carbon atom have in diamond?
2. What type of bonds are in diamond?
3. Why is diamond hard?
4. Why does diamond not conduct electricity?

1. What structure does silicon dioxide have?
2. Why does this structure have a high melting and boiling point?

1. What is graphene?
2. State a property of graphene.
3. How many bonds does each carbon have?
4. What does this allow graphene to do?

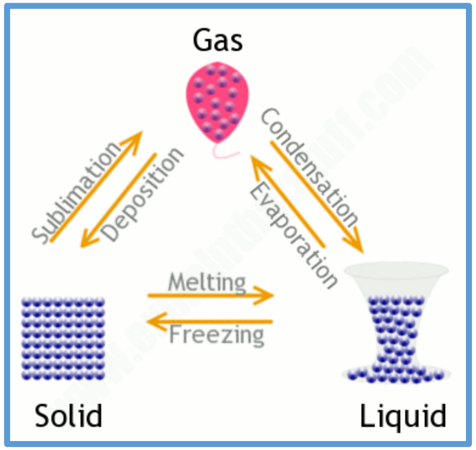
1. What element is graphite made from?
2. How many bonds does each carbon have?
3. Why can graphite conduct electricity?
4. What holds together the layers of graphite?
5. Why is graphite soft/slippery?
6. Does graphite have a high/low melting point?
7. Why?

1. What can fullerenes be used for?
2. What is the formula of buckminsterfullerene?
3. State two uses of carbon nanotubes.

1. What are three ways that H_2O could be drawn?

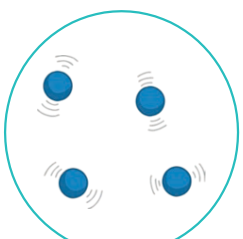
C2 – Bonding, structure, and the properties of matter

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



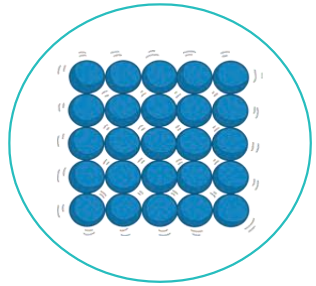
- When heated, particles **gain energy**.
- **Attractive forces** between particles begin breaking when melting or boiling points are reached
- **Amount of energy** needed to change state depends on how strong forces are.

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



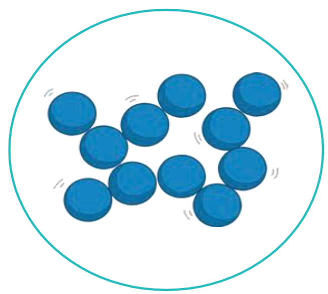
- Gases **are able to flow** – fill containers
- **Can be compressed as there is space between particles**

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.

C2 – Bonding, structure, and the properties of matter

1. What are the three states of matter?
2. What happens to particles when they are heated?
3. What happens to attractive forces when particles are heated?
4. What does the amount of energy needed to change state depend on?

1. How are gas particles arranged?
2. How do gas particles move?
3. Do particles in a gas have more or less kinetic energy than those in solids and liquids?
4. Can gases be compressed? Why?

1. How are solid particles arranged?
2. Do solid particles move?
3. Do particles in a solid have a high or low amount of kinetic energy?
4. Can solid particles flow?
5. Can solids be compressed?

1. How are liquid particles arranged?
2. Do particles in a liquid move?
3. Do the particles in a liquid have more or less kinetic energy than solids?
4. Can liquid particles flow?
5. Can liquids be compressed?

1. Where are state symbols used?
2. Write the symbols for solid, liquid, gas and aqueous.
3. What does aqueous mean?

1. If the temperature is lower than melting point, the substance is..
2. If the temperature is between melting and boiling point, the substance is..
3. When would a substance be gas?

1. State two limitations of the particle model.

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.

$2Mg$ + O_2 → $2MgO$
(2 x 24) + (2 x 16) → 2(24 + 16)
48 + 32 → 2 x 40
80 → 80

Reacting masses

Use conservation of mass to predict masses:

magnesium + oxygen → magnesium oxide
8g ? 10g

Both sides need to be equal:
10g – 8g = 4g 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 <u>oxygen</u> in magnesium oxide (MgO).

A_r of magnesium = 24 A_r of oxygen = **16**

M_r of MgO = 24 + 16 = **40**

$$\% \text{ mass} = \frac{\mathbf{A_r}}{\mathbf{M_r}} = \frac{16}{\mathbf{40}} = 0.4 \times 100 = 40\%$$

X 100 to make a %

40% of the mass of MgO is oxygen

Mass Changes

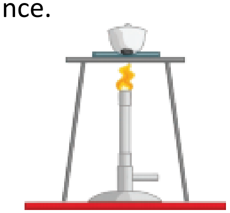
- Mass is always conserved in a reaction.
- Sometimes it may seem like the mass has increased/decreased.

- If a **reactant** is a gas – mass may **increase**.

magnesium + oxygen → magnesium oxide

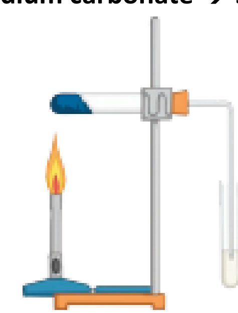
Oxygen is in the air before it combines with magnesium – you cannot find the mass of oxygen on the balance.

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**.

sodium carbonate → sodium oxide + carbon dioxide



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

16 O 8	12 C 6

CO_2

12 + (16 x 2) = 44

The Mole (HT only)

- Avogadro constant – 6.02 x 10^{23}

- One mole contains 6.02 x 10^{23} atoms or molecules
- The mass, in g, of one mole is the Ar (if an element) or Mr 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 x 10^{23} atoms of iron

Ammonia (NH_3) has an Mr of 17, so 1 mole of ammonia has a mass of 17g. and contains 6.02 x 10^{23} molecules of ammonia

C3 – Quantitative Chemistry

1. What is meant by conservation of mass?
2. Mass of reactants = ?
3. The M_r of the left side of an equation must be the same as..

1. How do you calculate the percentage mass of an element in a compound?
2. What do you do to convert a decimal into a percentage?

1. Should mass change in a reaction?
2. If a reactant is a gas, what will happen to the mass?
3. Why will it appear this has happened?
4. If a product is a gas, what will happen to the mas?
5. Why will it appear this has happened?

1. What does M_r stand for?
2. What is the relative formula mass?
3. Where can you find the relative atomic mass (A_r) of an element?

1. How many atoms are in one mole?
2. How do we know what the mass of one mole of an element is?
3. How do we know the mass of one mole of a compound?

C3 – Quantitative Chemistry

Concentrations of Solutions

- Concentration =mass of dissolved substance in specific volume (eg dm^3)
- More substance dissolved = more concentrated solution

Concentration = mass ÷ volume
(g/dm^3) (g) (dm^3)

Can be rearranged to find mass dissolved:

mass = concentration x volume
(g) (g/dm^3) (dm^3)

$1000cm^3 = 1dm^3$
$cm^3 \rightarrow dm^3$ = 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 $25cm^3$ of a $96g/dm^3$ solution

$96g/dm^3$ means 96g in every 1000cm

25c m^3

How do we get from 1000 to 25? (÷40)

Do the same to the other side (÷40)

2.4g

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)	$Na + O_2 \rightarrow Na_2O$
write down the mass or % <u>given in the question</u>	$18.4 + 6.4 \rightarrow 24.8$
Write the mass of one mole of each element or compound	23 32 62 (e.g 18.4 ÷ 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	$4Na + O_2 \rightarrow 2Na_2O$

Calculating reacting masses (HT)

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

Step	**Calculation**
Write the balanced equation	$2Ca + O_2 \rightarrow 2CaO$
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.

C3 – Quantitative Chemistry

1. What does concentration mean?
2. How can you make a solution more concentrated?
3. State the equation to calculate concentration in g/dm^3.
4. What is the unit for volume?
5. How many cm^3 are in a dm^3?

Calculating mass in a given volume

1. What does $36.5 g/dm^3$ mean?

2. Calculate the mass of dissolved solid in 25 cm^3 of a $36.5 g/dm^3$ solution

36.5 → ___ g

1000 → 25c m^3

Do the same to the other side (÷40)

How do we get from 1000 to 25? (÷40)

Moles and Equations (HT only)

12g of magnesium (Mg) reacted with 8g of oxygen (O_2) to produce 20g magnesium oxide (MgO). Use the masses to write a balanced equation

Step	Example
Write the equation for the reaction (unbalanced)	
write down the mass or % given in the question	
Write the mass of one mole of each element or compound	
Divide the mass given in question by the mass of one mole	
Turn the answers into whole number simple ratio	
Put the numbers into the equation	

1. What is a limiting reactant?

2. Complete the calculation: Calculate the mass of calcium needed to make 224g of calcium oxide

Step	**Calculation**
Write the balanced equation	$2Ca + O_2 \rightarrow 2CaO$
Write the masses of each substance	
Write down the given mass in the question.	
Work out the 'scale' factor (ie what did you have to do to the original number to get to the desired mass	
Do the same to the other side	

C4 – Chemical Changes

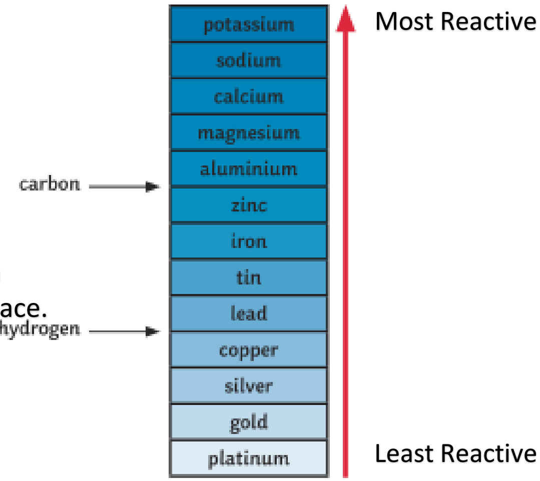
The Reactivity Series

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

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

Potassium is more reactive than magnesium

Potassium **displaces** magnesium from the compound and takes it's place.



Reactions of acids with metals

- Metal + acid → salt + hydrogen

E.g. iron + sulfuric acid → iron sulfate + hydrogen

salt

To name salt :
1st name Metal
2nd name Acid used

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

$2Mg + O_2 \rightarrow 2MgO$

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

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

Metal hydroxides are alkaline

Reactions of acids with alkalis

neutralisation

- Acid + alkali → salt + water

Hydrochloric acid + sodium hydroxide → sodium chloride + water

salt

$HCl + NaOH \rightarrow NaCl + H_2O$

Reactions of acids with carbonates

- Acid + carbonate → salt + water + carbon dioxide

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

salt

$H_2SO_4 + CaCO_3 \rightarrow CaSO_4 + H_2O + CO_2$

C4 – Chemical Changes

1. What is meant by displacement?
2. Name a very reactive metal
3. Name two metals which are less reactive than hydrogen.

---

1. Define extraction.
2. What is an ore?
3. How do you extract a metal less reactive than carbon?
4. What is meant by reduction?
5. What is meant by a 'native metal'?
6. Give an example of a metal found in native form.

---

1. State the general equation for the reaction of metal with oxygen.
2. Write a word equation for the reaction of iron with oxygen.

---

1. State the general equation for the reaction of metal with water.
2. Are hydroxides acid/alkaline?

---

1. State the general equation for the reaction of metal with acid.
2. State the salts produced from hydrochloric acid, sulfuric acid and nitric acid.

---

1. State the general equation for the reaction of acid with an alkali.

---

1. State the general equation for the reaction of acid with carbonates.

C4 – Chemical Changes

Redox Reactions (HT only)

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

- Metal + acid = redox reaction

Example

$H_2SO_4 + Ca \rightarrow CaSO_4 + H_2$

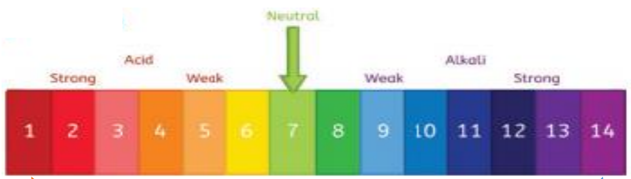
Ionic equation: $2H^+ + Ca \rightarrow Ca^{2+} + H_2$

Half equation 1: $Ca \rightarrow Ca^{2+} + 2e^-$ Lost 2 electrons (oxidation)

Half equation 2: $2H^+ + 2e^- \rightarrow H_2$ Gained 2 electrons (reduction)

pH Scale

- Shows how acidic or alkaline solution is.
- pH 1-6 = acid
- pH 7 = neutral
- pH 8-14 = alkali



Lots of H^+ | Less H^+ | Less OH^- | Lots of OH^-

In aqueous solutions:

Acids – produce H^+ ions

Alkalis – produce OH^- ions

In neutralisation reactions:

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

Strong/Weak Acids (HT only)

Strong acid = completely dissociates in a solution

e.g. $HCl \rightarrow H^+ + Cl^-$

Examples = nitric acid and sulfuric acid

Weak acid = partially dissociates in solution.

e.g. $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

$\rightleftharpoons$ = reversible reaction

Hasn't fully turned into ions – only partially

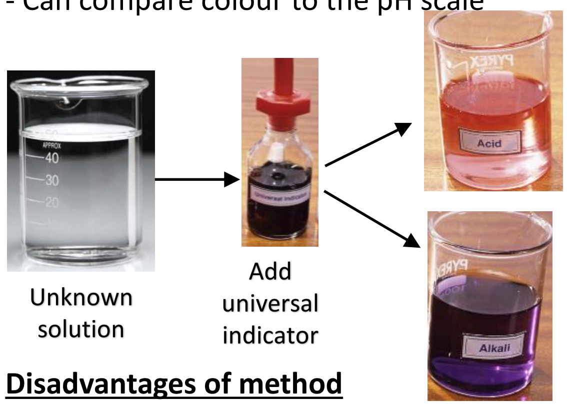
Concentration = how much is dissolved in every cm^3

Strong/weak = how well it ionises

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

Measuring pH of a solution

- Can use **universal indicator**
- Gives the solution a colour
- Can compare colour to the pH scale



Changes colour to give pH value

Disadvantages of method

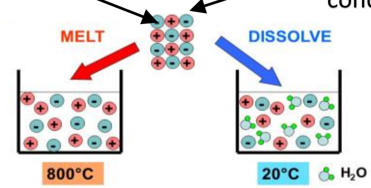
- Colour is **subjective** – different people may see different colours
- Doesn't give an exact pH number (could use **pH probe** to make more **accurate**).

Electrolysis

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

Ionic compound

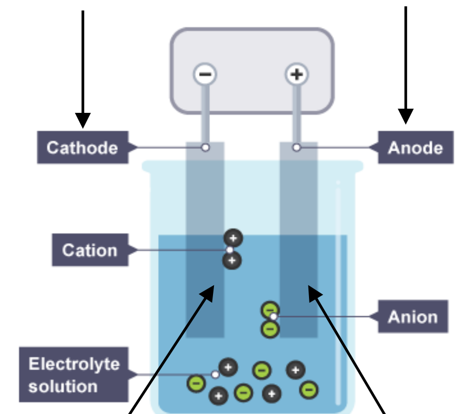
Solid – ions can't move to conduct electricity



- 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)



Positive ions move to the negative electrode (opposites attract)

Negative ions move to the positive electrode (opposites attract)

C4 – Chemical Changes

1. What is a redox reaction?
2. In terms of electrons, what does oxidation mean?
3. In terms of electrons, what does reduction mean?

1. What is the pH range for an acid?
2. What is the pH range for an alkali?
3. If a substance has a pH of 7, what type of substance is it?
4. What ions do acids produce in solution?
5. What ions do alkalis produce in a solution?
6. State the ionic equation for neutralisation reactions.

1. Define a strong acid.
2. Give an example of a strong acid.
3. Define a weak acid.
4. What happens to H^+ concentration as the pH value decreases by 1?

1. Describe a simple method to test the pH of an unknown solution.
2. State 2 disadvantages of using universal indicator.
3. How can pH be measured more accurately?

1. What is meant by the term electrolysis?
2. What is electrolysis used for?
3. What must the compound be for electrolysis to take place?
4. Why can solid ionic compounds not conduct electricity?
5. What does inert mean?
6. Name the positive electrode.
7. Name the negative electrode.
8. Why do positive ions move to the negative electrode?

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.**

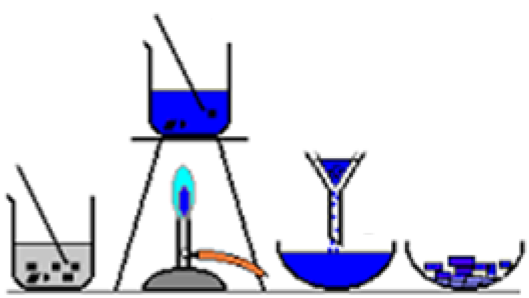
Equipment

- 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^3 **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 – Required Practical – Preparation of soluble salts

1. Write a method to prepare a pure, **dry** sample of copper sulfate crystals (6 marks).

Q2) Why do you heat the acid before adding the oxide?

Q3) Why is the oxide added in excess?

Q4) Why is the solution filtered?

Q5) Why is the solution left overnight in a warm, dry place?

Q6) Name 2 safety precautions you should take during this practical.

C4 – Chemical Changes

Electrolysis of Molten Ionic Compounds

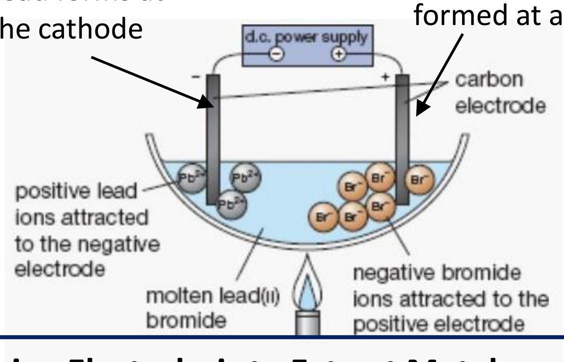
Molten = melted **so ions** can move.

- 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)**

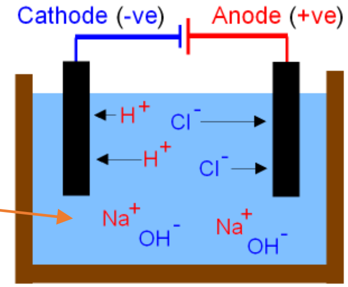
Example: Aluminium Oxide

- **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 MORE REACTIVE than hydrogen **K, Na, Ca, Mg, Zn, Fe** Then **HYDROGEN** is produced
If – ions are NOT Group 7 Eg sulphate SO_4^{2-} nitrate NO_3^- carbonate CO_3^{2-} OXYGEN is produced.	If + ions (metals) are LESS REACTIVE than hydrogen **Cu, Ag, Au** Then the METAL 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

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

Gained 2 electrons (reduction)

molecules of hydrogen gas produced

Anode – $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

molecules of oxygen produced

Lost electrons (oxidation)

Cathode – $Cu^{2+} + 2e^- \rightarrow Cu$

Gained electrons (reduction)

Copper atoms are formed at the cathode

Anode – $2Cl^- \rightarrow Cl_2 + 2e^-$

chlorine molecules are formed

Lost electrons (oxidation)

C4 – Chemical Changes

1. Why is an ionic compound melted before electrolysis takes place?
2. Metals are produced at the..
3. Non-metals are produced at the..

1. When is electrolysis used to extract a metal?
2. Why is electrolysis expensive?
3. Why is cryolite added to aluminium oxide before electrolysis?
4. Why does the positive anode need constantly replacing when electrolysing aluminium oxide?

1. Why is the compound dissolved in water before electrolysing?
2. What two ions are also present in aqueous solutions (along with the compound)?
3. Which two substances can be produced at the anode?
4. Which two substances can be produced at the cathode?
5. When would a metal be produced at the cathode?
6. When would oxygen be produced at the anode?

1. In terms of electrons, what happens at the positive electrode?
2. In terms of electrons, what happens at the negative electrode?
3. Write the half equation for the production of hydrogen.
4. Write the half equation for the production of oxygen from hydroxide ions.
5. Write the half equation for the production of copper from copper ions.
6. Write the half equation for the production of chlorine from chloride ions.

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.

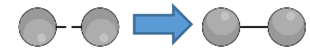
Endothermic Reactions

- 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

Energy change of reactions (HT)

During a reaction:

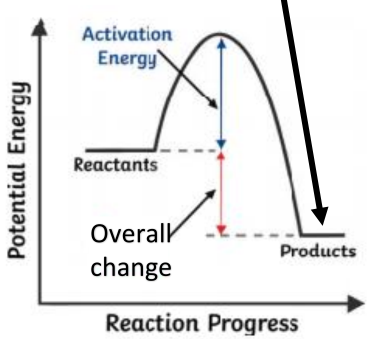
- Energy is **absorbed** in order to **break** bonds in the reactants 
- Energy is **released** when bonds are **made** in the products. 

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

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.

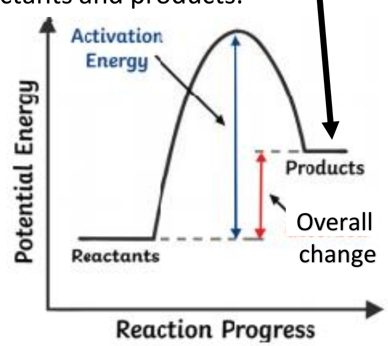
You may need to draw and label this in the exam!



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!



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

$$H_2 + F_2 \rightarrow 2HF$$

H–H F–F H–F

bonds broken bonds formed

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

Bonds broken =	Bonds formed
436 + 158	2 x 568
593	1136

Overall energy change = 593 – 1136

= -543 kJ/mol Exothermic

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

C5 – Energy Changes

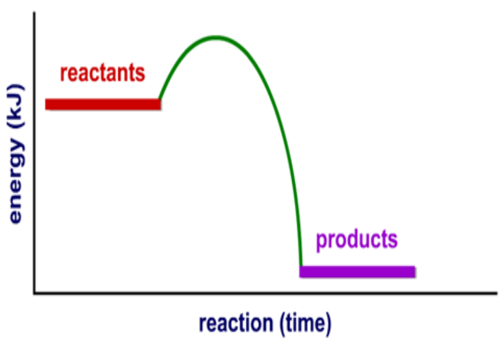
1. Which way is energy transferred in an exothermic reaction?
2. What happens to the temperature of the reaction mixture in an exothermic reaction?
3. State two examples of exothermic reactions.

1. Which way is energy transferred in an endothermic reaction?
2. What generally happens to the temperature of the reaction mixture of an endothermic reaction?
3. State two examples of endothermic reactions.

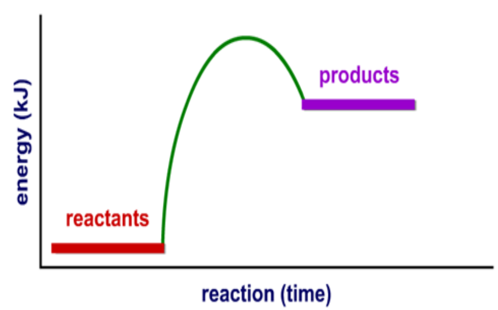
Higher Tier only

1. In terms of energy, what happens for bonds to be broken?
2. In terms of energy, what happens when bonds are formed?

1. Define activation energy.
2. On the graph below, draw and label the :
   - overall energy change
   - activation energy



1. What does an energy level diagram show?
2. On the graph below, draw and label the :
   - overall energy change
   - activation energy



Higher Tier only

1. Define overall energy change.
2. How do you calculate energy change?
3. Why, in terms of bond breaking and making, is a reaction exothermic?
4. Why, in terms of bond making and breaking, is a reaction endothermic?

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^3 beaker
- Measuring cylinder
- Liquid reactants

Method (example for hydrochloric acid and sodium hydroxide)

1. Using measuring cylinder to measure 30cm^3 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^3 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^3 more sodium hydroxide up to a maximum of 40cm^3.
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.**

C5 – Energy Changes **Required Practical – Temperature Changes**

1. Write a method to investigate how the volume of sodium hydroxide affects the change in temperature when reacting with hydrochloric acid (6 marks)

2. For the investigation above, name the :
Independent variable :
Dependent variable :
2 control variables :

3. Why do you use a polystyrene cup and lid instead of a beaker?

4. Why should you calculate the temperature change, instead of just using the final temperature?

5. Why is it important to stir the mixture?

6. Why do we do repeat readings?