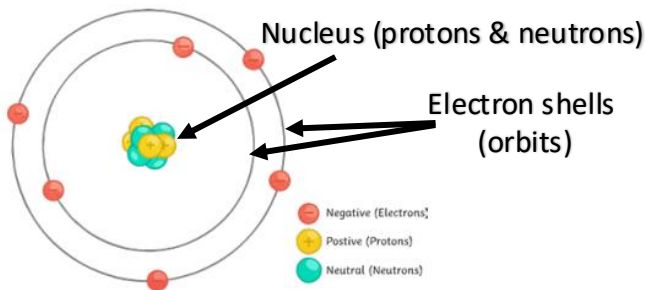


# C1 – Atomic Structure and The Periodic Table

## Atoms

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



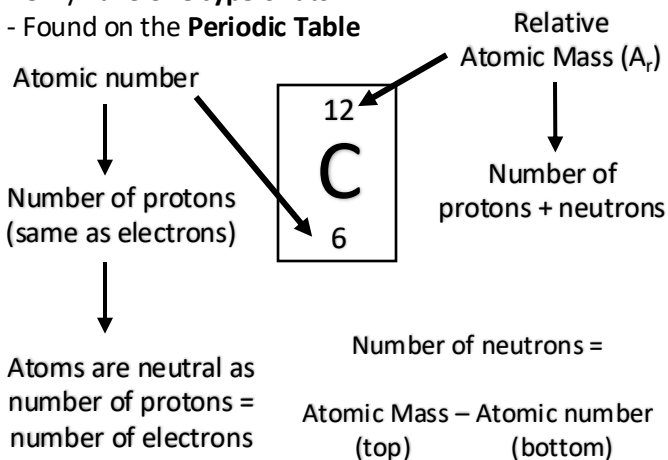
| Subatomic particle | Relative Mass | Charge   |
|--------------------|---------------|----------|
| Proton             | 1             | Positive |
| Neutron            | 1             | Neutral  |
| Electron           | Very small    | Negative |

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**



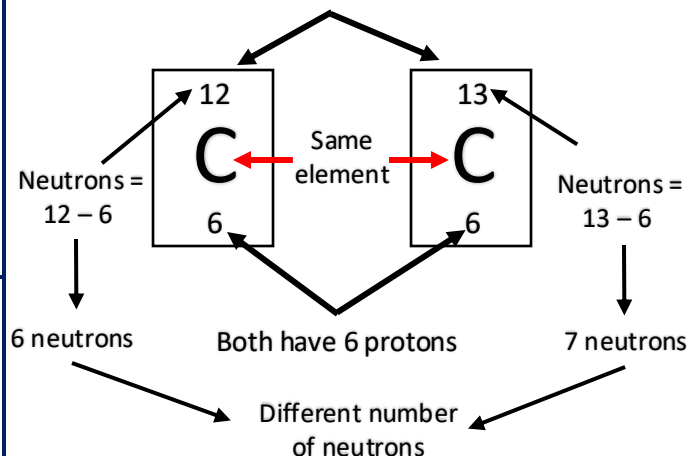
## Compounds

- 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..



## Mixtures and Separation

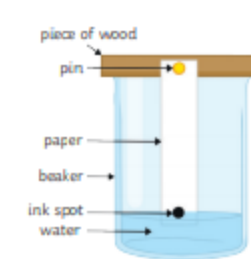
**Mixtures** – two or more elements or compounds **not** chemically joined.

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)



## 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.  $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

## 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?

1. Where are elements found?

2. What does the relative atomic mass of an element show?

3. What does the atomic number show?

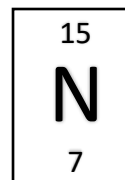
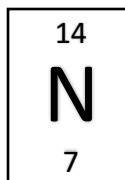
4. How do you calculate the amount of neutrons?

1. Define the word compound.

2. Give three examples of compounds.

1. What is an isotope?

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



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

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

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.

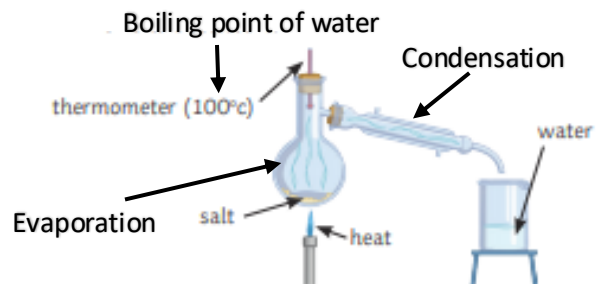
5. What is evaporation used to separate?

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.



- 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).

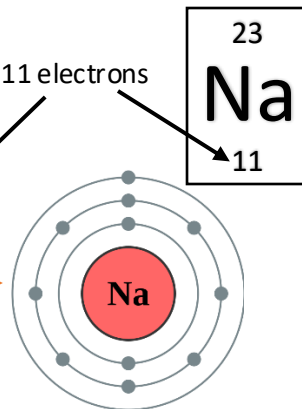
## Electronic Structure

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

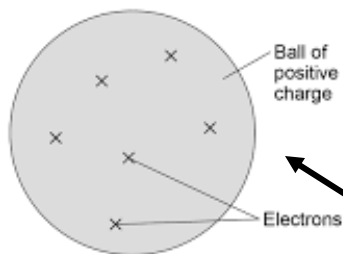
First shell = 2 electrons  
Second shell = 8 electrons  
Third shell = 8 electrons.

1<sup>st</sup> shell = 2  
2<sup>nd</sup> shell = 8  
3<sup>rd</sup> shell = 1

Total = 11 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?

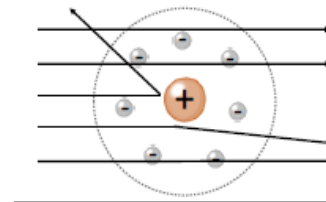
### Rutherford's scattering experiment

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

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

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?

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.

1. How many electrons do the halogens have in the outer shell?

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?

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?

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

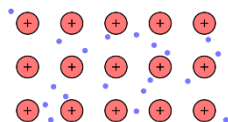
| Group | Ions | Example                                              |
|-------|------|------------------------------------------------------|
| 1     | +1   | $\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$     |
| 2     | +2   | $\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$ |
| 6     | -2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

Lost electrons

Gained electrons

## Metallic Bonding

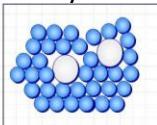
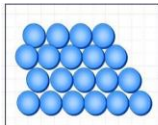
- Happens in **metals only**.
- Positive metal ions surrounded by **sea of delocalised electrons (can move)**.
- Ions tightly packed in rows.
- Strong **electrostatic forces of attraction** between positive ions and negative electrons.



## Alloys

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

Pure Metal

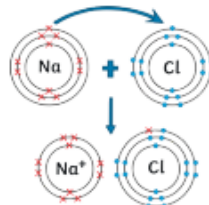


Alloy

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

## Ionic Bonding

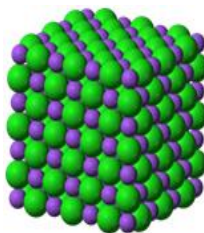
- Between a metal and non-metal.
- Metals give electrons to non-metals so both have a full outer shell.
- **Electrostatic force of attraction** between positive and negative ions.



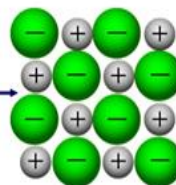
E.g. Sodium loses one electron to become  $\text{Na}^+$ . Chlorine gains one electron to become  $\text{Cl}^-$ . The two ions attract to form sodium chloride.

## 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

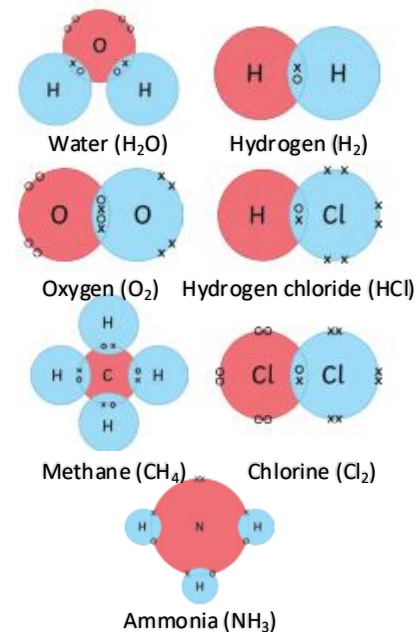
- **High melting point** – lots of energy needed to overcome electrostatic forces.
- **High boiling point**
- **Cannot conduct electricity as solid** – ions cannot move
- **Conducts electricity when molten or dissolved** – ions are free to move.

## 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:



## 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<sub>2</sub>O)

Methane (CH<sub>4</sub>)

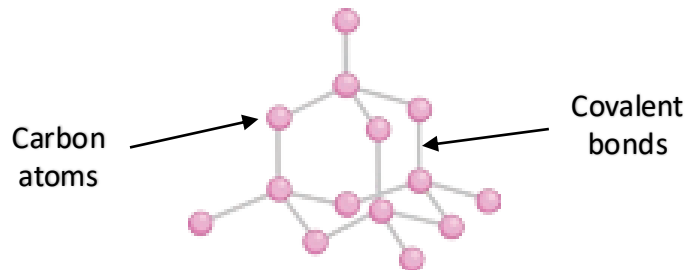
Oxygen (O<sub>2</sub>)

5. Do simple covalent molecules have a high/low melting point?
6. Why is this?

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

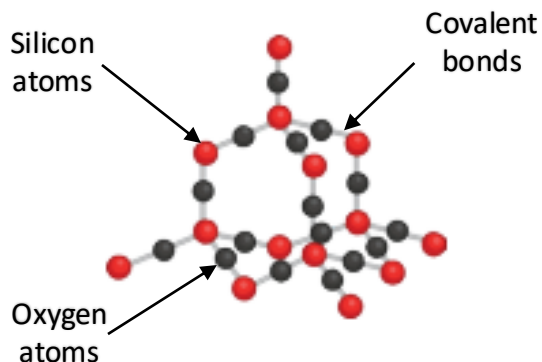
### Giant Covalent Structure – Diamond

- 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.



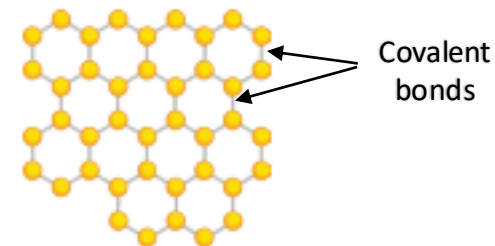
### Silicon Dioxide

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



### Graphene

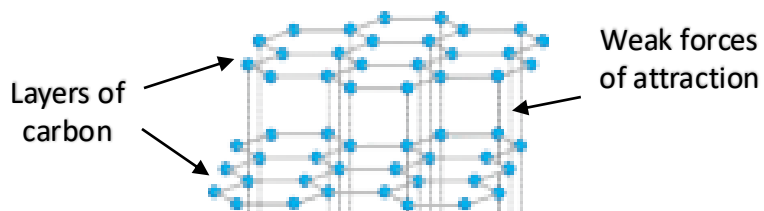
- Graphene = one layer of graphite.
- Very strong → lots of strong covalent bonds.



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

### 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**.

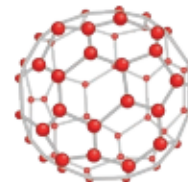


- Layers held together by **weak forces**
- Layers can **slide** over each other easily
- Makes graphite **soft/slippery** → good lubricant.
- Has **high melting point** as has many strong covalent bonds.

### Fullerenes and Nanotubes

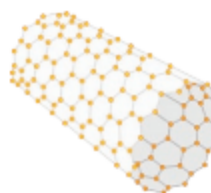
- Molecules of carbon shaped into hollow tubes or balls.
- Used to **deliver drugs into body**

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



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

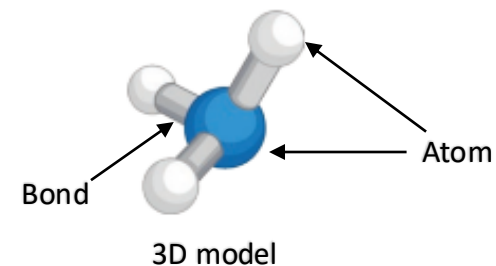
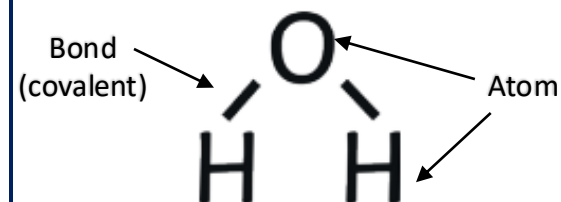
- Can strengthen materials without adding weight.



- Used in electronics and nanotechnology.

### Molecular models

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



## 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?

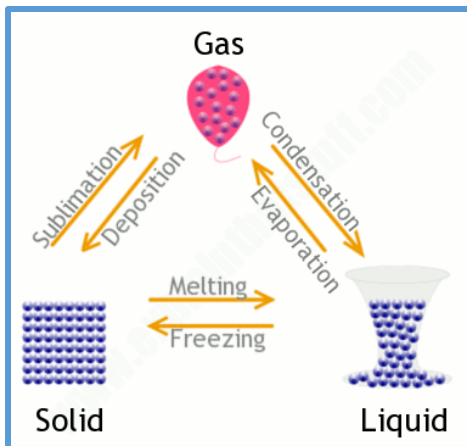
1. What can fullerenes be used for?
2. What is the formula of buckminsterfullerene?
3. State two uses of carbon nanotubes.

1. What are three ways that H<sub>2</sub>O could be drawn?

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

### States of Matter

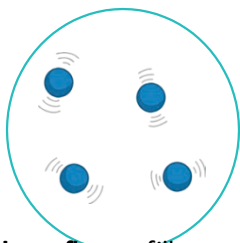
- Three states of matter: **solid, liquid & gas.**
- To change state, **energy must be transferred.**



- 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

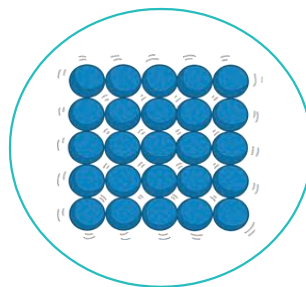
- Randomly arranged.
- Particles **move quickly** – all directions.
- Highest **amount of kinetic energy.**



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

### Solid

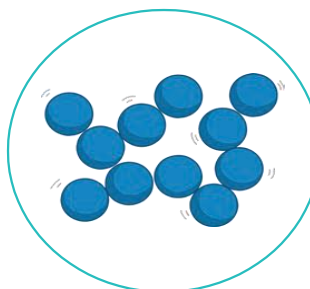
- **Regular** pattern (rows and columns)
- Particles **vibrate** in a **fixed position.**
- Particles have **low amount of kinetic energy.**



- 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?

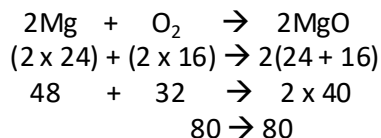
1. State two limitations of the particle model.

# C3 – Quantitative Chemistry

## Conservation of Mass

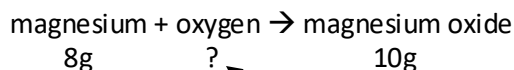
- Atoms cannot be created or destroyed during reactions.
- **Mass of reactants = mass of products.**

To show mass is conserved in a reaction:  
M<sub>r</sub> 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:  
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 oxygen in magnesium oxide (MgO).

A<sub>r</sub> of magnesium = 24      A<sub>r</sub> of oxygen = 16

M<sub>r</sub> of MgO = 24 + 16 = 40

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

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**.

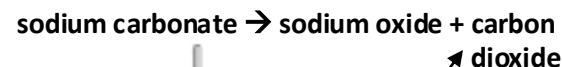


Oxygen is in the air before it combines with magnesium – you cannot find the mass of oxygen on the balance.

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**.

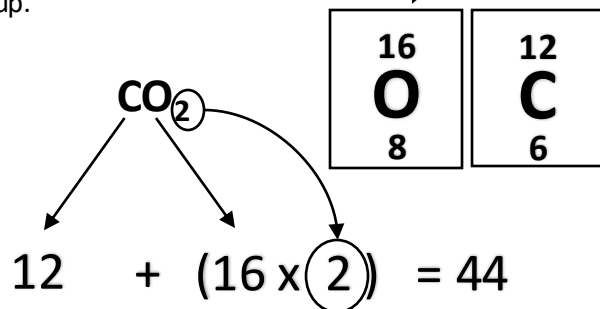


It will look like the mass has decreased as some of the atoms have been given off as gas and have escaped – so cannot be re-weighed.

## Atomic mass (A<sub>r</sub>) and Relative Formula Mass (M<sub>r</sub>)

- Atomic mass (A<sub>r</sub>) is the mass number – ie the mass of one atom
- Relative formula mass (M<sub>r</sub>) = all the **relative atomic masses (A<sub>r</sub>)** 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<sub>r</sub> (if an element) or M<sub>r</sub> if a compound or molecular element

Iron has a A<sub>r</sub> of 56, so 1 mole of iron is 56 g and contains  $6.02 \times 10^{23}$  atoms of iron

Ammonia (NH<sub>3</sub>) has an M<sub>r</sub> 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

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?

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?

4. If a product is a gas, what will happen to the mas?
5. Why will it appear this has happened?

# 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>)

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

$$\text{cm}^3 \rightarrow \text{dm}^3 =$$

**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

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)          | Na + O <sub>2</sub> → Na <sub>2</sub> O                     |
| write down the mass or % given in the question            | 18.4 + 6.4 → 24.8                                           |
| Write the mass of one mole of each element or compound    | 23    32    62<br>(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<br>(cancel down)<br>4        1        2 |
| Put the numbers into the equation                         | 4Na + O <sub>2</sub> → 2Na <sub>2</sub> 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                                                                               | 2Ca + O <sub>2</sub> → 2CaO |
| Write the masses of each substance                                                                        | 80 + 32 → 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  $\text{g/dm}^3$ .
4. What is the unit for volume?
5. How many  $\text{cm}^3$  are in a  $\text{dm}^3$ ?

### Calculating mass in a given volume

1. What does  $36.5\text{g/dm}^3$  mean?
2. Calculate the mass of dissolved solid in  $25\text{ cm}^3$  of a  $36.5\text{g/dm}^3$  solution

$$\begin{array}{ccc} 36.5 & & 1000 \\ & & \downarrow \\ & & 25\text{ cm}^3 \end{array}$$

Do the same to the other side ( $\div 40$ )

How do we get from 1000 to 25? ( $\div 40$ )

g

### Moles and Equations (HT only)

12g of magnesium (Mg) reacted with 8g of oxygen ( $\text{O}_2$ ) 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 % <u>given in the question</u>     |         |
| 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

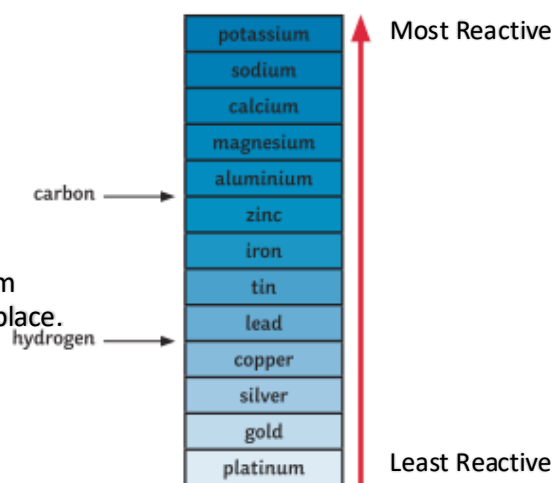
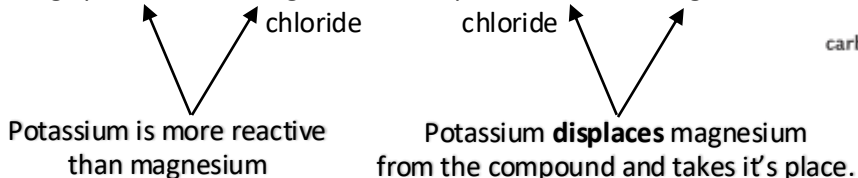
| Step                                                                                                       | Calculation                                       |
|------------------------------------------------------------------------------------------------------------|---------------------------------------------------|
| Write the balanced equation                                                                                | $2\text{Ca} + \text{O}_2 \rightarrow 2\text{CaO}$ |
| 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

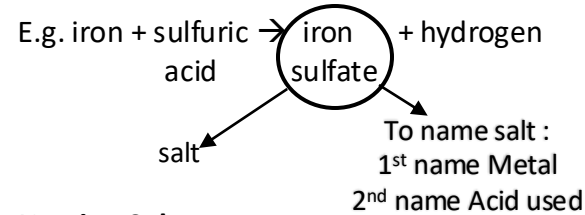
- A more reactive metal will replace a less reactive metal in a compound (**displacement**)

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



## 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

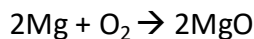
**More reactive than carbon** – electrolysis is used.

- 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

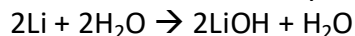
- Oxidation = gaining oxygen  
- Reduction = losing oxygen

## 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

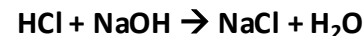
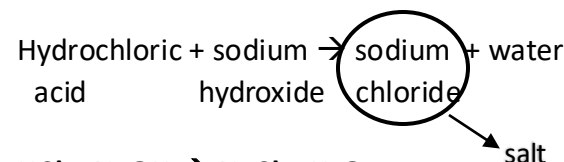
e.g. lithium + water → lithium hydroxide + hydrogen



Metal hydroxides are alkaline

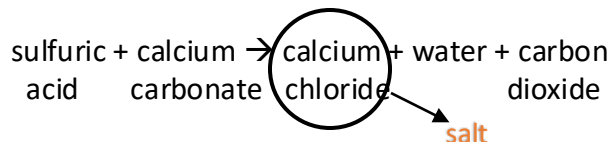
## Reactions of acids with alkalis

- Acid + alkali → salt + water (neutralisation)



## Reactions of acids with carbonates

- Acid + carbonate → salt + water + carbon dioxide



## C4 – Chemical Changes

1. What is meant by displacement?

2. Name a very reactive metal

3. Name two metals which are less reactive than hydrogen.

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. Define extraction.

2. What is an ore?

3. How do you extract a metal less reactive than carbon?

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 acid with an alkali.

4. What is meant by reduction?

1. State the general equation for the reaction of metal with water.

5. What is meant by a 'native metal'?

2. Are hydroxides acid/alkaline?

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

6. Give an example of a metal found in native form.

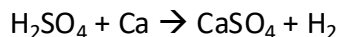
# C4 – Chemical Changes

## Redox Reactions (HT only)

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

- Metal + acid = redox reaction

### Example



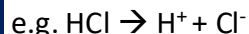
Ionic equation:  $2\text{H}^+ + \text{Ca} \rightarrow \text{Ca}^{2+} + \text{H}_2$  Lost 2 electrons (oxidation)

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)

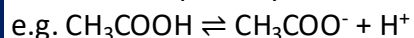
## Strong/Weak Acids (HT only)

**Strong acid** = completely dissociates in a solution



Examples = nitric acid and sulfuric acid

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



$\rightleftharpoons$  = reversible reaction

Hasn't fully turned into ions – only partially

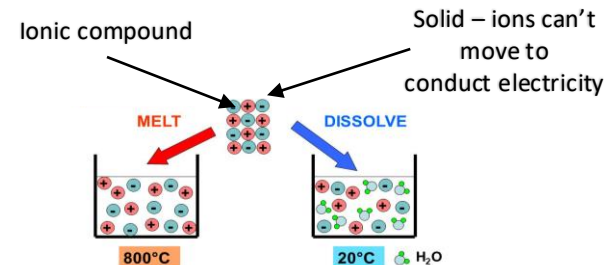
**Concentration** = how much is dissolved in every  $\text{cm}^3$

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

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

## Electrolysis

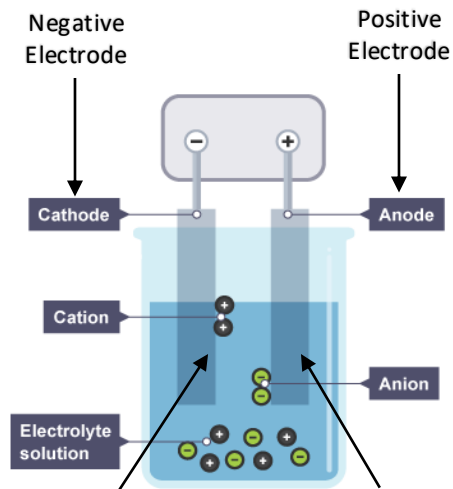
- **Splitting** up a **compound** using electricity.
- Used to extract metals from compounds, purify metals (eg copper)



- 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)

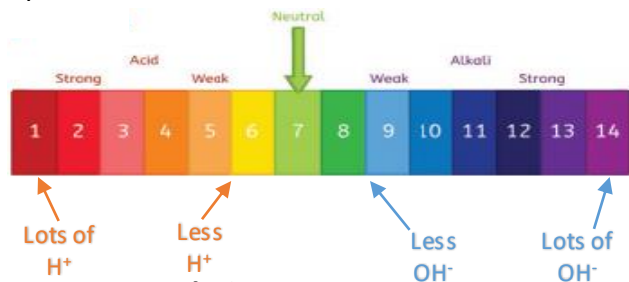


Positive ions move to the negative electrode (opposites attract)  
Negative ions move to the positive electrode (opposites attract)

## pH Scale

- Shows how acidic or alkaline solution is.

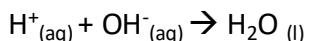
- pH 1-6 = acid
- pH 7 = neutral
- pH 8-14 = alkali



**In aqueous solutions:**

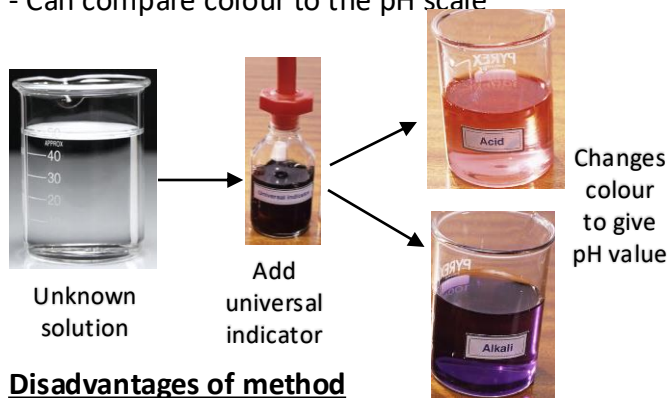
Acids – produce  $\text{H}^+$  ions  
Alkalis – produce  $\text{OH}^-$  ions

**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

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. 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. 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?

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. 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?

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**.

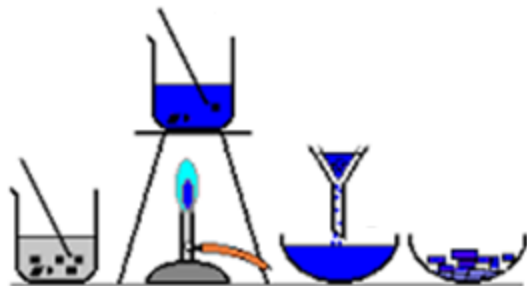
### Equipment

- 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<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 – 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.

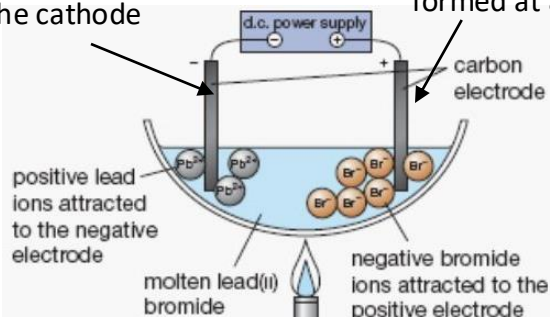
# C4 – Chemical Changes

## Electrolysis of Molten Ionic Compounds

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

### Example: Lead Bromide - PbBr<sub>2</sub>

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)**

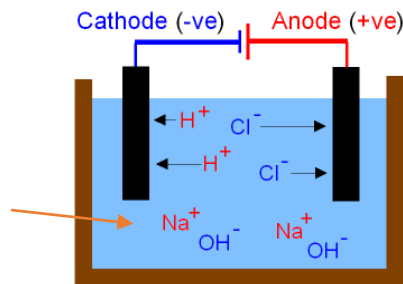
### Example: Aluminium Oxide

- **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<sub>2</sub> – it will degrade.

## Electrolysis of Aqueous Solutions

- Compound is dissolved in water so ions can move.

When aqueous – H<sup>+</sup> and OH<sup>-</sup> (from H<sub>2</sub>O) 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                                                                                                                                                                                 | - CATHODE                                                                                                              |
|-----------------------------------------------------------------------------------------------------------------------------------------------------------------------------------------|------------------------------------------------------------------------------------------------------------------------|
| Attracts – ions ('Anions')                                                                                                                                                              | Attracts + ions ('Cations')                                                                                            |
| If – ions are group 7 i.e.<br>chloride Cl <sup>-</sup><br>bromide Br <sup>-</sup><br>iodide I <sup>-</sup><br>Then the groups 7 element is produced as a gas                            | If + ions (metals) are MORE REACTIVE than hydrogen<br><b>K, Na, Ca, Mg, Zn, Fe</b><br>Then <b>HYDROGEN</b> is produced |
| If – ions are NOT Group 7<br>Eg sulphate SO <sub>4</sub> <sup>2-</sup><br>nitrate NO <sub>3</sub> <sup>-</sup><br>carbonate CO <sub>3</sub> <sup>2-</sup><br><b>OXYGEN</b> is produced. | If + ions (metals) are LESS REACTIVE than hydrogen<br><b>Cu, Ag, Au</b><br>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:

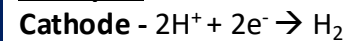
**Cathode** – positive ions **gain** electrons (**reduction**)

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

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

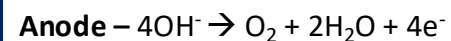
- Reactions at electrodes can be represented by half equations.

### Examples



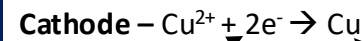
Gained 2 electrons (reduction)

molecules of hydrogen gas produced



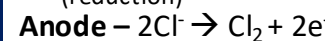
molecules of oxygen produced

Lost electrons (oxidation)



Gained electrons (reduction)

Copper atoms are formed at the cathode



chlorine molecules are formed

Lost electrons (oxidation)

## 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.

## C4 – Chemical Changes – Required Practical – Electrolysis of Aqueous Solutions

### Aim

To investigate the electrolysis of an aqueous solution using inert (unreactive) **electrodes**.

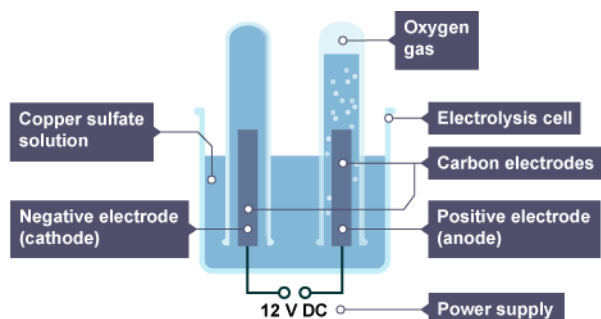
### Equipment

- Beaker
- Two test tubes (or measuring cylinders)
- Graphite electrodes
- Two splints
- Aqueous solution
- DC powerpack

Change method depending on the question.

### Method (example copper sulfate solution.)

1. Pour some copper sulfate solution into a beaker.
2. Place two graphite rods into the copper sulfate solution. Attach one electrode to the negative terminal of a dc supply, and the other electrode to the positive terminal.
3. Completely fill two small test tubes with copper sulfate solution and position a test tube over each electrode as shown in the diagram. **(use measuring cylinders if measuring volume of gas produced)**
4. Turn on the power supply and observe what happens at each electrode.
5. Test any gas produced with a glowing splint and a burning splint.
6. Record observations and the results of your tests.



### Common questions

**Q1)** How do you test for hydrogen gas?

**A1)** Lit splint will make a squeaky pop.

**Q2)** How do you test for oxygen gas?

**A2)** Glowing splint – will relight.

**Q3)** Explain why copper is produced at the cathode.

**A3)** Copper ions are **positive**, so are attracted to the negative electrode (opposites attract). Copper is less reactive than hydrogen so is discharged. The copper ions **gain electrons** and are **reduced** to form **copper atoms**.

**Q4)** Why do hydrogen ions move to the cathode?

**A4)** Hydrogen ions are **positive** so move to the negative electrode as **opposites attract**.

**Q5)** Why are measuring cylinders better to collect the gas?

**A5)** Because they are more accurate when measuring the volume of gas produced.

## C4 – Chemical Changes – Required Practical – Electrolysis of Aqueous Solutions

Q1. Draw a labelled diagram to show the equipment needed to electrolyse copper chloride.

Q2. Write a method for the electrolysis of aqueous copper chloride solution.

Q2) How do you test for hydrogen gas?

Q3) How do you test for oxygen gas?

Q4) Explain why copper is produced at the cathode.

Q5) Why do hydrogen ions move to the cathode?

Q6) Why are measuring cylinders better to collect the gas?

# 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. 

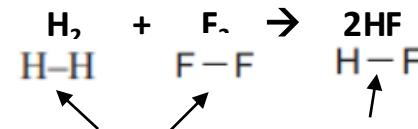
**Bond energy** = the amount of energy that is released when a bond is made or that is needed to break a bond

## 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



bonds broken                      bonds formed

| Bond | Bond Energy /<br>kJ mol <sup>-1</sup> |
|------|---------------------------------------|
| F—F  | 158                                   |
| H—H  | 436                                   |
| H—F  | 568                                   |

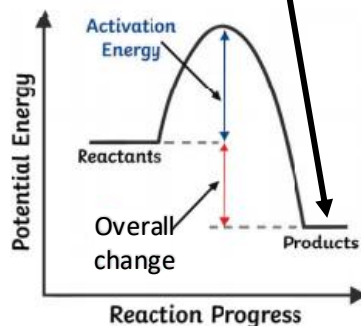
|                                    |                                 |
|------------------------------------|---------------------------------|
| Bonds broken =<br>436 + 158<br>593 | Bonds formed<br>2 x 568<br>1136 |
|------------------------------------|---------------------------------|

**Overall energy change = 593 – 1136**  
**= -543 kJ/mol Exothermic**

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

## Reaction Profiles – Exothermic

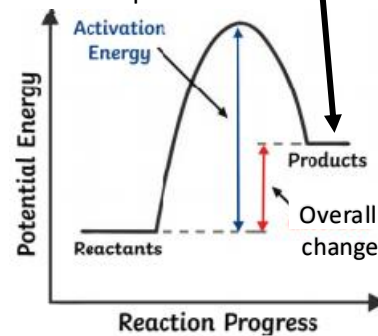
- Energy level diagrams show **difference in energy** between reactants and products.
- Exothermic = Energy of products is **lower than** reactants (energy is released)
- **Activation Energy** = minimum amount of energy needed to start the reaction.
- **Energy change** = the difference in energy between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

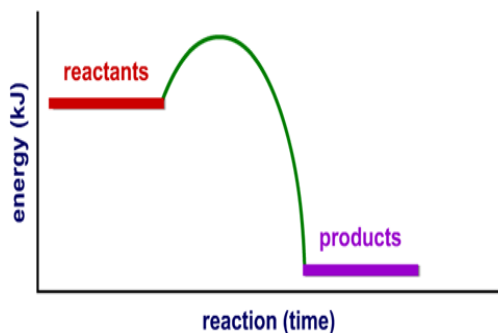
1. Which way is energy transferred in an exothermic reaction?
2. What happens to the temperature of the reaction mixture in an exothermic reaction?
3. State two examples of exothermic reactions.

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.

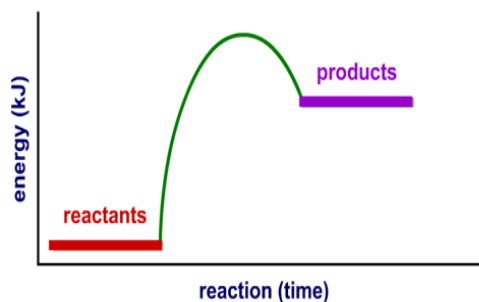
### Higher Tier only

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<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**.

## 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?