Chemistry Paper 1 (H) Knowledge Recall Booklet

Whitley Academy

Paper Chemistry 1H 8464/C/1H

For this paper, the following list shows the major focus of the content of the exam:

- 5.2.2 How bonding and structure are related to the properties of substances
- 5.3.2 Use of amount of substance in relation to masses of pure substances
- 5.4.1 Reactivity of metals
- 5.4.2 Reactions of acids
- 5.4.3 Electrolysis
- 5.5.1 Exothermic and endothermic reactions

Required practical activities that **will be assessed**:

- Required practical activity 8: preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate, using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution.
- Required practical activity 9: investigate what happens when aqueous solutions are electrolysed using inert electrodes. This should be an investigation involving developing a hypothesis.
- Required practical activity 10: investigate the variables that affect temperature changes in reacting solutions such as, eg, acid plus metals, acid plus carbonates, neutralisations, displacement of metals.

Your teacher will tell you specific topic areas to focus on …

Topic	RAG			Revision technique					Date completed	Teacher Signed
	R	A	G	Flashcards	Mindmap	Notes	Video Watched (With notes)	Frog resource		

Chemistry Paper 1 (F) Knowledge Recall Booklet

Paper Chemistry 1F 8464/C/1F

For this paper, the following list shows the major focus of the content of the exam:

- 5.1.2 The periodic table
- 5.2.2 How bonding and structure are related to the properties of substances
- 5.2.3 Structure and bonding of carbon
- 5.4.1 Reactivity of metals
- 5.4.2 Reactions of acids
- 5.4.3 Electrolysis

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Required Practical – Making Soluble Salts

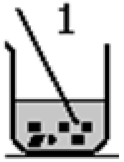
Recall it …

Use the information in the following page(s) to answer these questions …

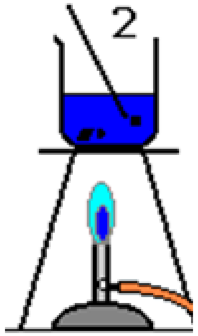
1. Outline all the steps / method for this investigation in sufficient detail.
2. What are the reactants to make a soluble salt?
3. What safety precautions are required?
4. Describe the method for electrolysis?

Required Practical – Preparation of soluble salt

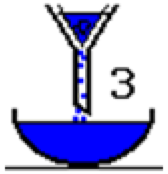
Soluble salts can be made from **acids** by reacting them with **solid insoluble substances**, such as **metals, metal oxides, hydroxides or carbonates**. The solid is added to the acid until no more reacts and the excess solid is **filtered** off to produce a solution of the salt. Salt solutions can be **crystallised** to produce **solid** salts. **You will complete this as a required practical.**



1. Measure the required volume of acid with a measuring cylinder and add the weighed solid (insoluble metal, oxide, hydroxide or carbonate) in small portions with stirring.



2. Safety goggles required - the mixture may be heated to speed up the reaction. When no more of the solid dissolves it means ALL the acid is neutralised and there should be a little excess solid. You should see a residue of the solid (oxide, hydroxide, carbonate) left at the bottom of the beaker.



3. Filter the solution to remove the excess solid metal/oxide/carbonate, into an evaporating dish. On filtration, only a solution of the salt is left.

4. Then hot concentrated solution is left to cool and crystallise.
   After **crystallisation**, you collect and dry the crystals with a filter paper.
   If the solution is heated, the solvent will evaporate faster. Heating a solution until all the solvent has evaporated is known as **heating to dryness**.

Required Practical – Electrolysis

Recall it …

Use the information in the following page(s) to answer these questions …

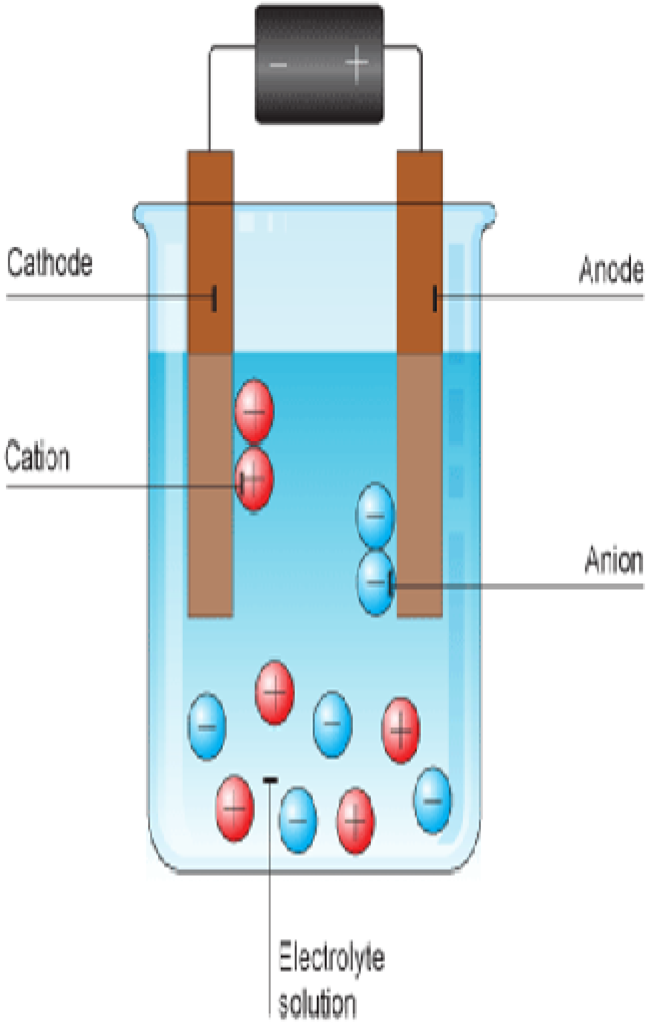
1. What happens in electrolysis?
2. What is the anode? What happens at the anode?
3. What is the cathode? What happens at the cathode?
4. Describe the method for electrolysis?

Required Practical - Electrolysis

3. Electrolysis of Aqueous Solution (aq)

CONTENT LINK

- In electrolysis, the ions move towards the oppositely charged electrodes.
- When you electrolyse aqueous ionic solutions - you get three products..
- When electrolysis happens in aqueous solution, the less reactive element, either hydrogen or the metal, is usually produced at the Cathode. At the Anode, you get either;
  - Oxygen gas given off, from discharged hydroxide ions produced from water.
  - Halogen produced if the electrolyte is a solution of a halide



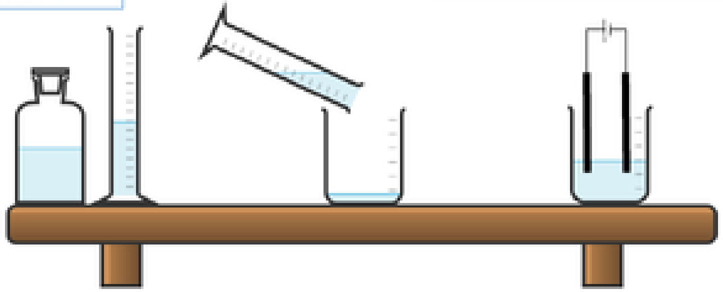
Positive Anode Negative is Cathode

Electrolysis

1. Measure 50 cm³ of copper (II) chloride solution

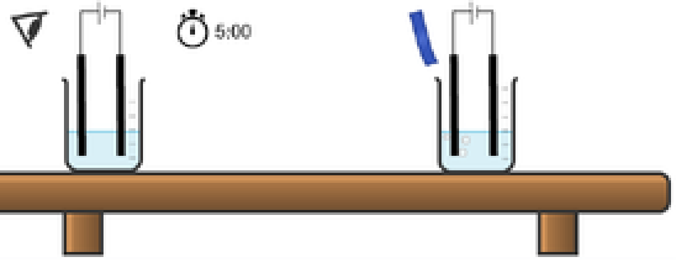
2. Pour into a 100 cm³ beaker

3. Insert the graphite electrodes into the solution. Connect one to the positive outlet of the power pack and the other to the negative.



4. Turn the power pack to 4V and switch on for no longer than 5 minutes. Observe what is happening at each electrode.

5. If you can see bubbles forming at the anode (positive electrode) hold a piece of damp, blue litmus paper next to the bubbles. If the paper turns white, the bubbles are chlorine.



6. Repeat steps 1 to 5 for solutions of copper (II) sulfate, sodium chloride and sodium sulfate.

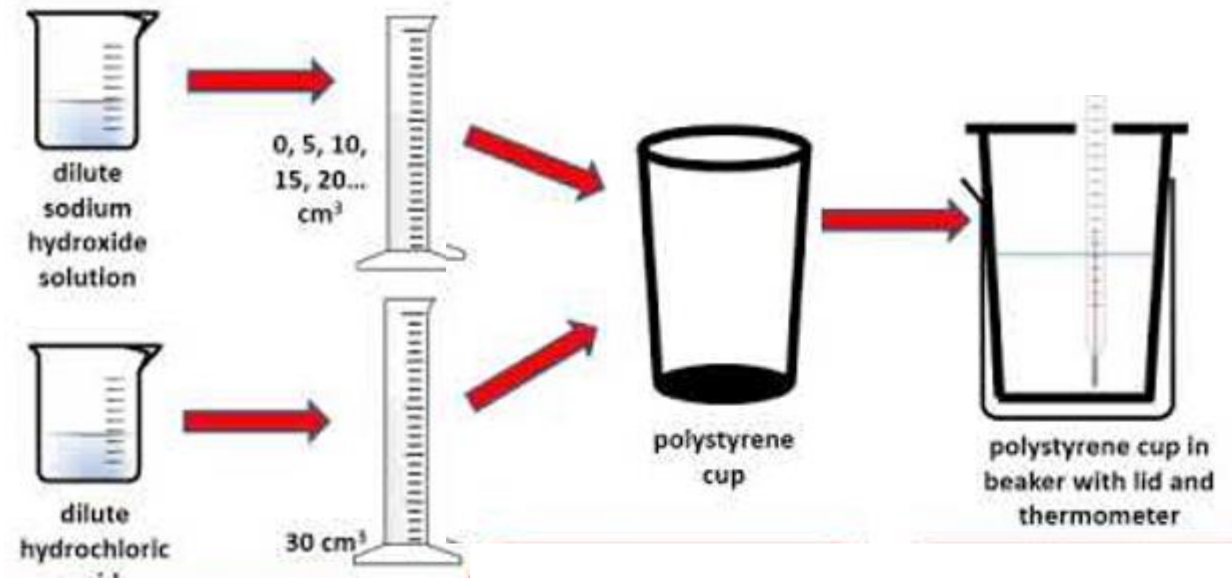
Required Practical – Temperature Change

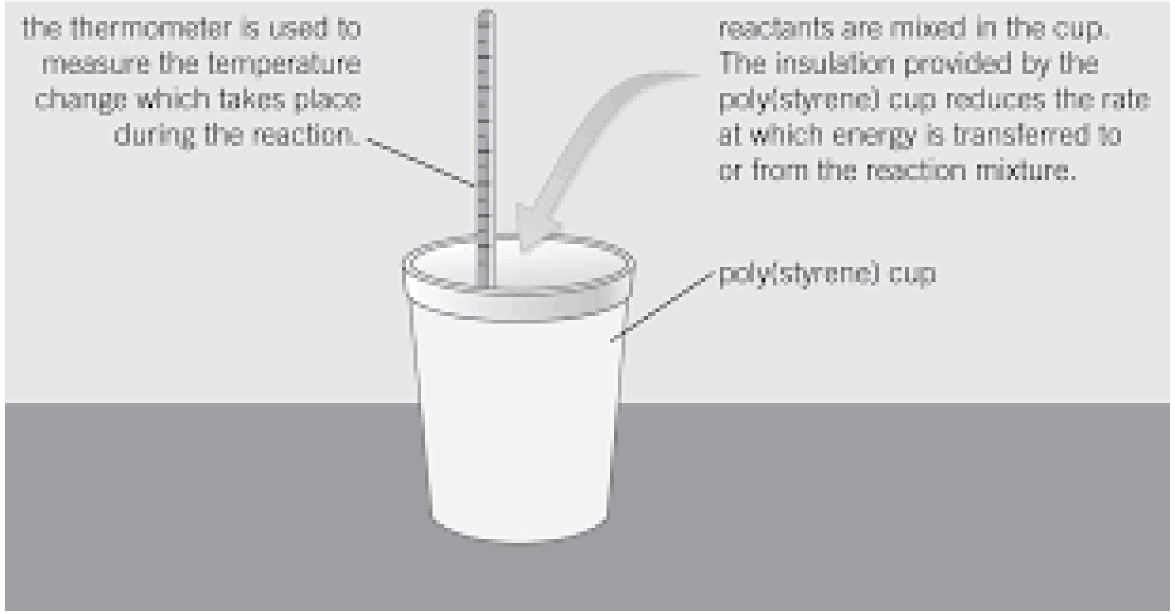
Recall it …

Use the information in the following page(s) to answer these questions …

1. What are the reactants for this experiment?
2. What is the independent, dependent and control variables for this experiment?
3. Describe the method for the experiment?
4. Why is a polystyrene cup used?

Required Practical – Temperature Change





Method:

1. Measure $30cm^3$ dilute hydrochloric acid and put it into the polystyrene cup.
2. Stand the cup inside the beaker. This will make it more stable
3. Use the thermometer to measure the temperature of the acid. Record your result in the table below.
4. Measure $5cm^3$ sodium hydroxide solution.
5. Pour the sodium hydroxide into the polystyrene cup. Fit the lid and gently stir the solution with the thermometer through the hole.
6. Look carefully at the temperature rise on the thermometer.
7. When the reading on the thermometer stops changing, record the highest temperature reached in the table.
8. Repeat steps 4-7 to add further $5cm^3$ amounts of sodium hydroxide to the cup each time, recording your temperature reading in the results table.
9. Repeat until a maximum of $40cm^3$ of sodium hydroxide has been added.
10. Wash out all the equipment and repeat the experiment for your second trial.

Changes of state and heating curves

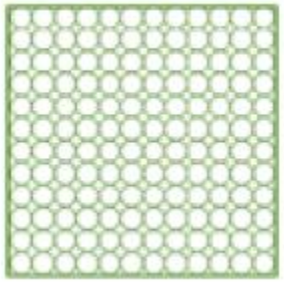
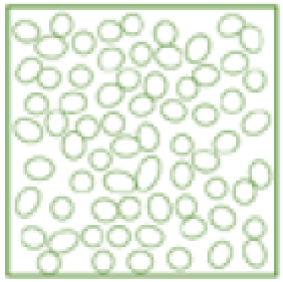
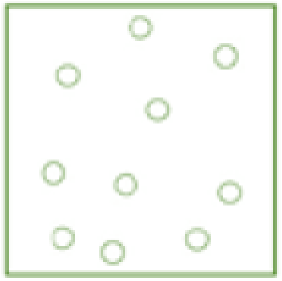
Recall it …

Use the information in the following page(s) to answer these questions …

1) Describe how particles behave in a solid?
2) Describe how particles behave in a liquid?
3) Describe how particles behave in a gas?
4) Describe 4 changes of state?
5) What happens at the melting point and the boiling point?
6) What does the energy needed to change state depend upon?
7) Why may some substances have high melting and boiling points?
8) What happens when the temperature does not change in a heating curve?

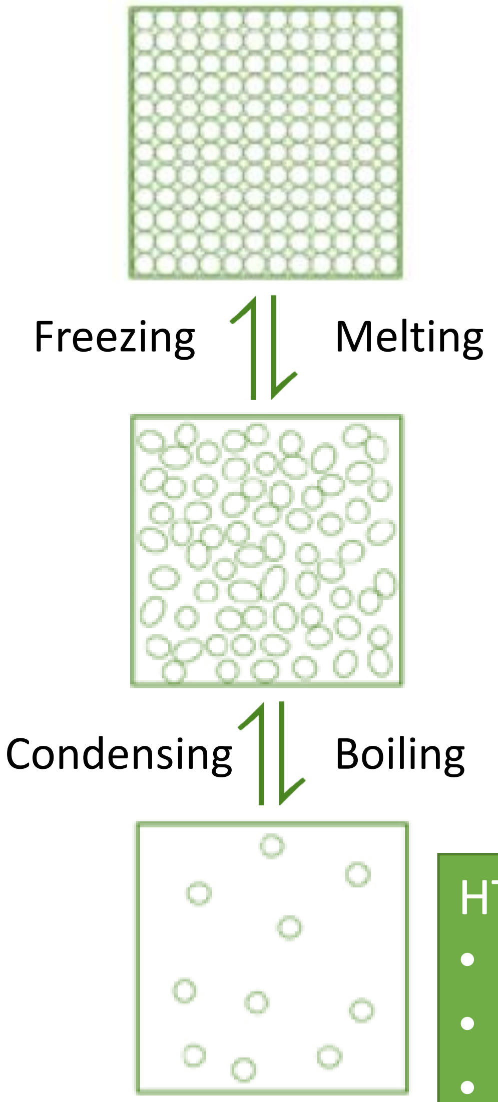
States of matter and state symbols

There are **three** states of matter – **solid, liquid** and **gas**. To explain the properties of the states, the **particle theory** is used. It is based on the fact that all matter is made up of tiny particles and describes the **movement** and **distance** between particles.

Solid	Liquid	Gas
Close together, regular pattern, vibrate on the spot.	Close together, random arrangement, move around each other.	Far apart, random arrangement, move quickly.
		

In chemical equations, the three states are shown as **(s), (l), (g)** and **(aq)** for aqueous solutions.

Changes of state



Melting and **freezing** take place at the **melting point**. **Boiling** and **condensing** take place at the **boiling point**.

The **amount of energy** required to change the state depends on the **strength of the forces** between the particles of the substance.

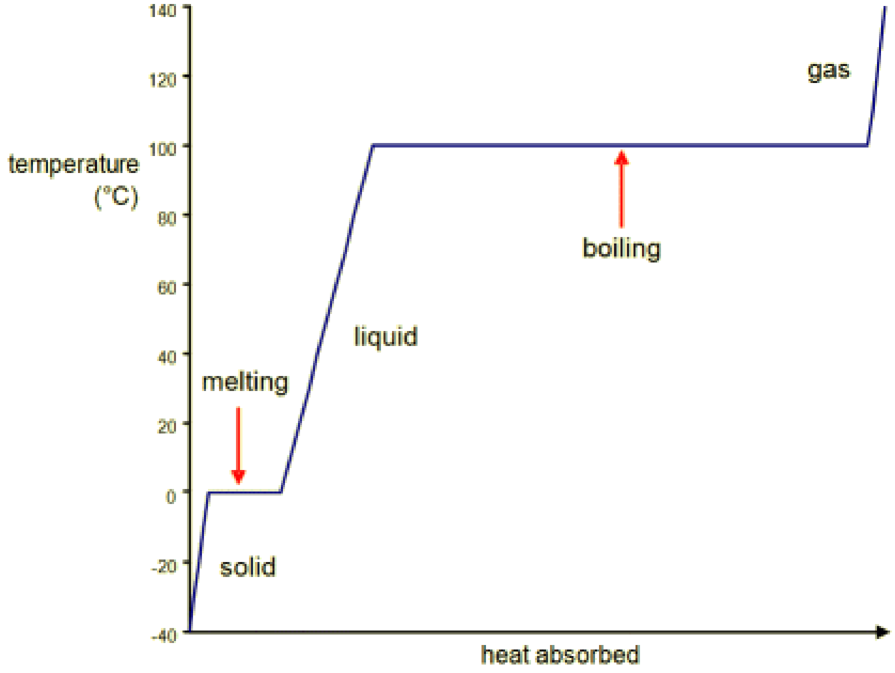
The **stronger the forces** between the particles the **higher the melting and boiling point** of the substance.

The type of bonding and the structure of the substance depend on the particles involved.

HT ONLY - There are limitations of the particle model of matter:

- There are no forces
- All particles are shown as spheres
- The spheres are solid

The graph shows a **heating curve** of a solid, which shows the temperature of a substance plotted against the amount of energy it has absorbed:



A substance must **absorb** heat energy so that it can **melt** or **boil**. The **temperature** of the substance does **not change** during **melting**, **boiling** or **freezing**, even though energy is still being transferred.

Recall it …

Structures

Use the information in the following page(s) to answer these questions …

1) Describe the structure of sodium chloride (NaCl)
2) Describe and explain the properties of sodium chloride (NaCl)
3) Describe the structure of small molecules?
4) Describe and explain the properties of small molecules?
5) How the size of small molecules affect their melting and boiling points? Explain why?
6) What are polymers? Describe and explain the properties of polymers?
7) Describe and explain the properties of metals?
8) Describe and explain the properties of alloys?
9) Describe and explain the properties of giant covalent structures?
10) Describe and explain the properties of diamond?
11) Describe and explain the properties of graphite?
12) Describe and explain the properties of graphene?
13) Describe and explain the properties of fullerenes?

Properties of ionic compounds

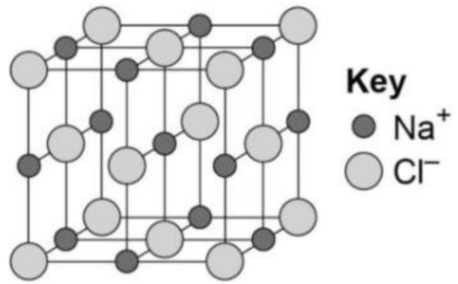
Structure

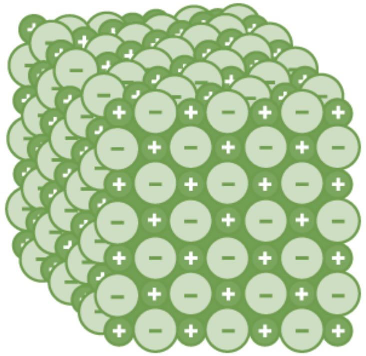
Ionic compounds have **regular structures** called **giant ionic lattices**.

There is **strong electrostatic forces** of attraction in all directions between **oppositely changed ions**.

Properties

- **High melting and boiling points** – large amounts of energy is needed to break the many **strong bonds** and overcome the electrostatic attraction.
- **Conduct electricity** when molten or dissolved in water – **ions are free to move** and can carry charge.





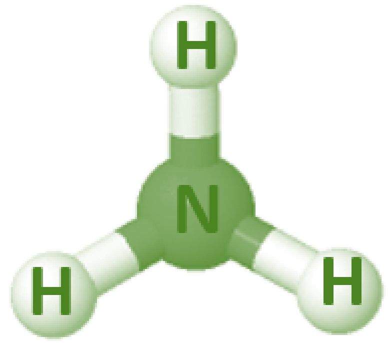
Properties of small molecules

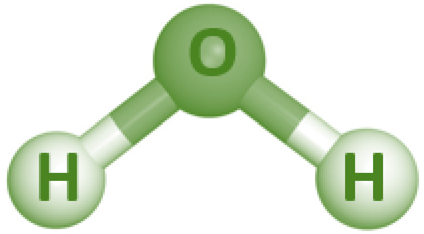
Structure

They have **weak forces between the molecules**. These weak forces are overcome when they change state **not** the strong covalent bonds.

Properties

- **Low melting and boiling points** – small amounts of energy is needed to break the **intermolecular forces**. Most are gases or liquids.
- **Do not conduct electricity** – Particles do not have an overall electric charge.





Intermolecular forces increase with the size of the molecules. So larger molecules have higher melting and boiling points.

Polymers

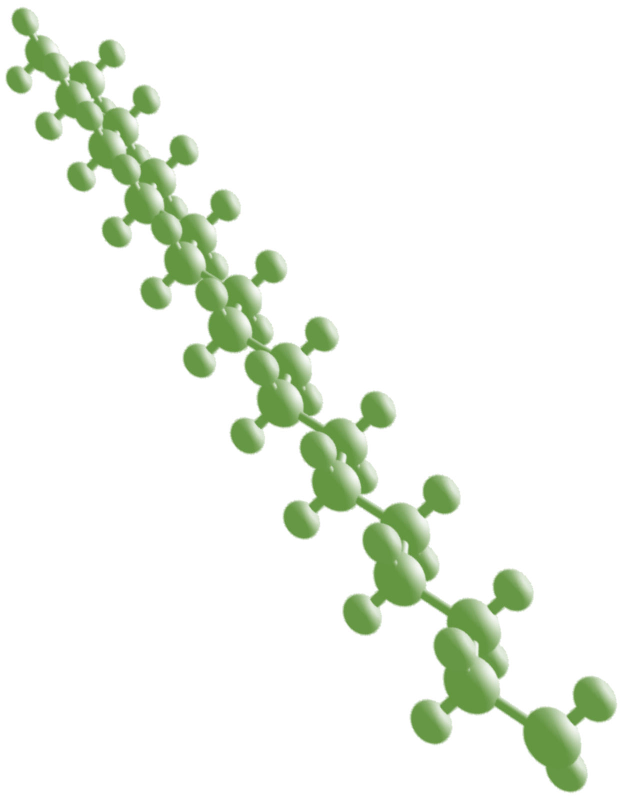
Some covalently bonded substances have very **large** molecules, such as **polymers**.

Structure

Polymers are made up from many small reactive molecules that bond to each other to form **long chains**. The atoms in the polymer molecules are linked to other atoms by **strong covalent bonds**. The **intermolecular forces** between polymer molecules **are relatively strong**.

Properties

- **Solid** at room temperature – **Strong intermolecular forces**.

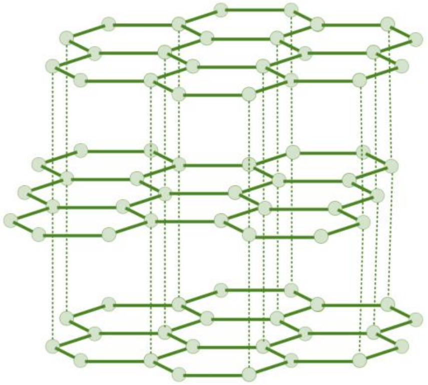


Giant covalent structures

Structure
All atoms within the structure are linked by **strong covalent bonds**. These bonds **must be broken** for a solid to melt or boil.

Properties

- **Very high melting and boiling points** – very large amounts of energy is needed to break the **covalent bonds**.
- **Do not conduct electricity** – Particles do not have an overall electric charge.

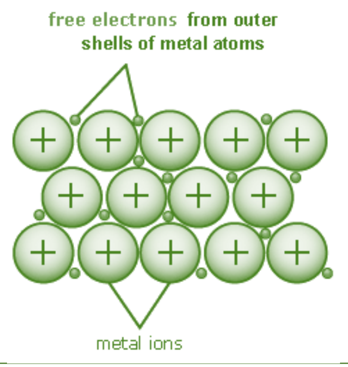


Properties of metals and alloys

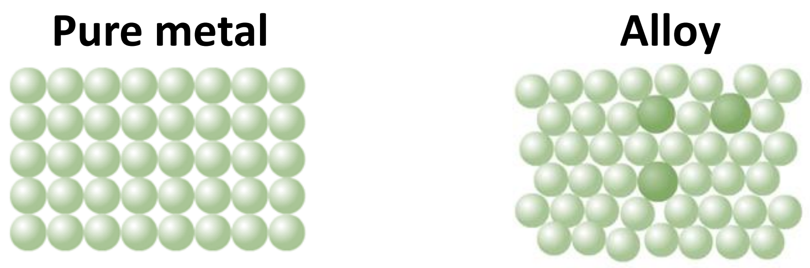
The giant structure of atoms with strong metallic bonding gives most metals a **high melting** and **boiling point**.

Metals are **malleable** (can be hammered into shape) and **ductile** (can be drawn out into a wire) because the **layers** of atoms (or ions) in a giant metallic structure can **slide** over each other

Delocalised electrons in metals enable **electricity** and **heat** to pass through the metal easily.



A **metal mixed** with other **elements** is called an **alloy.** Alloys are **harder** than pure metals. Alloys are made from **two or more** different metals.



The **different sized** atoms of the metals **distort** the **layers** in the structure, making it more **difficult** for them to **slide** over each other, and so make the **alloys harder** than pure metals.
For example, **gold** is naturally **soft** but adding **copper** to make jewellery **stronger** and last longer.

Recall it … Molar Calculations

Use the information in the following page(s) to answer these questions …

1) Describe what is a mole? What are the units?
2) What is the avogadro constant?
3) How do you calculate a mole and give a worked example?
4) How do you re-arrange the formula for a mole? Give a worked example?
5) Describe the steps involved in calculating the mass of the reactants or products using a balanced symbol equation?
6) Describe what is concentration?
7) Describe how to calculate concentration? Give a word equation?
8) Describe how to re-arrange the formula for concentration? Give a word equation?
9) Describe how the concentration of an aqueous solution can be increased?

Use of amount of substance - PART 1

Chemical amounts are measured in **moles**. The symbol for the unit mole is **mol**.

The mass of **one mole** of a substance **in grams** is numerically **equal** to its **relative formula mass**. **One mole of a substance contains the same number of the stated particles, atoms, molecules or ions as one mole of any other substance**.

The **number** of atoms, molecules or ions in a mole of a given substance is the **Avogadro constant**. The value of the Avogadro constant is **6.02×10^{23} per mole**.

Number of moles = $\frac{\text{mass (g)}}{A_r}$ or $\frac{\text{mass (g)}}{M_r}$

Mass (g) = number of moles x A_r
or number of moles x M_r

How many moles of sulfuric acid molecules are there in 4.7g of sulfuric acid (H_2SO_4)? Give your answer to 1 significant figure.

$$\frac{4.7}{98} = 0.05 \text{ mol}$$

What is the mass of 7.2×10^{-3} moles of aluminium sulfate ($Al_2(SO_4)_3$)? Give your answer to 1 decimal place.

$7.2 \times 10^{-3} \times 342 = 2.5$g

Use of amount of substance - PART 1

HIGHER TIER

The **masses** of **reactants** and **products** can be calculated from **balanced symbol equations**.

Chemical equations can be interpreted in terms of **moles**. Example:

$$H_2 + Cl_2 \rightarrow 2HCl$$

This equation shows that **one** mole of **hydrogen** reacts with **one** mole of **chlorine** to form **two** moles of **hydrochloric acid**.

The balanced equation is useful because it can be used to calculate what mass of hydrogen and chlorine react together and how much hydrogen chloride is made.

A_r: H (1)	so mass of 1 mole of H_2	= 2 x 1 = 2g
A_r: Cl (35.5)	so mass of 1 mole of Cl_2	= 35.5 x2 = 71g
M_r : HCl (1 + 35.5)	so mass of 1 mole of HCl	= 36.5g

The balanced equation tells us that one mole of hydrogen reacts with one mole of chlorine to give two moles of hydrogen chloride molecules, so turning this to masses:

1 mole of hydrogen	= 1 x 2	= 2g
1 mole of chlorine	= 1 x 71	= 71g
2 moles of hydrochloric acid	= 2 x 36.5	= 73g

Use of amount of substance - PART 1
HIGHER TIER

Sodium hydroxide reacts with chlorine to make bleach:

$$2NaOH + Cl_2 \rightarrow NaOCl + NaCl + H_2O$$

If you have a solution containing 100.0g of sodium hydroxide, what mass of chlorine gas do you need to convert it to bleach?

M_r : NaOH (23 + 16 + 1) so mass of 1 mole of NaOH = 40g
M_r : Cl_2 (35.5 x 2) so mass of 1 mole of Cl_2 = 71g

So 100.0g of sodium hydroxide is **100/40 = 2.5 moles**

The balanced symbol equation tells us that for every two moles of sodium hydroxide, you need one mole of chlorine to react with it.
So you need **2.5/2 = 1.25 moles** of chlorine
One mole of chlorine is 71g, so you will need **1.25 x 71g = 88.75g** of chlorine to react with 100.0g of sodium hydroxide.

Use of amount of substance - PART 1
HIGHER TIER

The **balancing numbers** in a **symbol** equation can be calculated from the **masses** of **reactants** and **products** by **converting** the **masses in grams** to **amounts in moles** and converting the number of moles to **simple whole number ratios**.

8.5g of sodium nitrate ($NaNO_3$) is heated until its mass is constant.
6.9g of sodium nitrite ($NaNO_2$) and 1.6g of oxygen gas (O_2) is produced.

$$NaNO_3 \rightarrow NaNO_2 + O_2$$

M_r: $NaNO_3$ = 23 + 14 + (16x3) = 85
M_r: $NaNO_2$ = 23 + 14 + (16x2) = 69
M_r: O_2 = 16x2 = 32

Number of moles = $\frac{\text{mass (g)}}{M_r}$

Then to convert masses to moles use:

Moles of $NaNO_3$ = 8.5/85 = 0.1 mol
Moles of $NaNO_2$ = 6.9/69 = 0.1 mol
Moles of O_2 = 1.6/32 = 0.05 mol

$NaNO_3$	:	$NaNO_2$	:	O_2
0.1	:	0.1	:	0.05

Dividing the ratio by the smallest number gives 2:2:1 $2NaNO_3 \rightarrow 2NaNO_2 + O_2$

Use of amount of substance - PART 1
HIGHER TIER

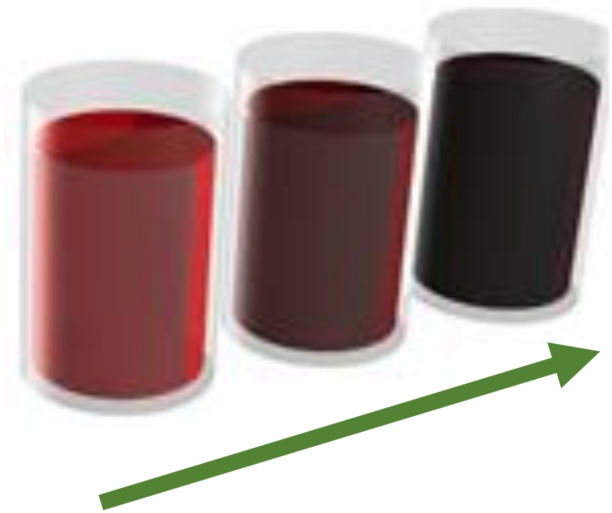
In a chemical reaction involving **two** reactants, it is common to use an **excess** of one of the reactants to ensure that all the reactant is **used up**. The reactant that is completely used up is called the **limiting reactant** because it **limits** the **amount** of **products**.

4.8g of magnesium ribbon reacts with 7.3g of HCl.
Which is the limiting reactant?

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

A_r: Mg (24) and A_r: Cl (35.5)
4.8g of Mg = 4.8/24 moles = 0.2 mol
7.3g of HCl = 7.3/36.5 moles = 0.2 mol
From the balanced equation:
1 mole of Mg reacts with 2 moles of HCl,
therefore **0.2 mol of Mg will need 0.4 mol of HCl** to react completely, there is only 0.2 mol of HCl, so the HCl is the limiting reactant.

Calculations - PART 2



Chemists quote the amount of substance (solute) dissolved in a certain volume of the solution. The units used to express the concentration can be grams per decimetre cubed **(g/dm^3)**. **A decimetre (1dm ^{3}) cubed is equal to 1000cm^3.**

The blackcurrant juice is getting more concentrated – the darker colour indicates more squash is in the same volume of its solution

If you know the mass of the solute dissolved in a certain volume of solution, you can work out the concentration using:

$$\textbf{Concentration (g/dm}^3\textbf{)} = \frac{\textbf{amount of solute (g)}}{\textbf{Volume of solution (dm}^3\textbf{)}}$$

Remember if you are using cm^3 to multiply the volume by 1000 to covert to dm^3

Example 1:

50g of sodium hydroxide is dissolved in water to make up 200cm^3 .

What is the concentration in dm^3?

50g/200cm^3= 0.25g/cm^3
0.25g/cm^3 x 1000 = 250g/dm^3

Calculations - PART 2

Example 2:

A solution of sodium chloride has a concentration of $200 g/dm^3$.

What is the mass of sodium chloride in $700 cm^3$ of solution?

Convert $700 cm^3$ into dm^3

$700/1000 = 0.7\ dm^3$

Then rearrange the equation

amount of solute = concentration x volume of solution

(g) (g/dm^3) (dm^3)

$200 g/dm^3 \times 0.7\ dm^3 = 140 g$

HIGHER:

You can increase the concentration of an aqueous solution by:

- Adding more solute and dissolving it in the same volume of its solution.
- Evaporating off some of the water from the solution so you have the same mass of solute in a smaller volume of solution.

Recall it … Reactivity of Metals

Use the information in the following page(s) to answer these questions …

1) Describe what is an ore? How do we remove metals from their oxides?
2) What is the reactivity series?
3) Describe observations of metal reactivity with water?
4) Describe observations of metal reactivity with acid?
5) Which metals are extracted by electrolysis?
6) Which metals are extracted by reduction with carbon?
7) Which metals are found native?
8) Describe the reaction between metals and acid, giving an example equation?
9) Describe how does the name of the acid change the name of the salt produced?
10) Give the equations for the reactions between

a) Acid and Alkali

b) Acid and Base

c) Acid and Metal carbonate

11) Describe how to work out the chemical formula of salts?
12) What is an indicator? Describe the colour changes and pH values of acids, alkalis and neutral substances?
13) What types of ions are produced by acids and alkalis? What happens to these ions during neutralisation reactions?

Metal Oxides

A **metal compound** within a **rock** is an **ore.** Ores are **mined** and then **purified**.

Whether it is worth extracting a particular metal depends on:

- **How easy it is to extract it from its ore**
- **How much metal the ore contains**
- **The changing demands for a particular metal**

Most **metals** in ores are **chemically bonded** to **other elements** in compounds. Many of these metals have been **oxidised** (have oxygen added) by oxygen in the air to form their oxides.

Iron + oxygen → iron (III) oxide

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

To extract metals from their oxides, the metal oxides must be **reduced** (have oxygen removed).

Metals can be arranged in order of reactivity in a **reactivity series**.

Order of reactivity	Reaction with water	Reaction with acid
Potassium	Fizz, giving off hydrogen gas and leaving an alkaline solution of metal hydroxide	Reacts violently and explodes
Sodium		
Lithium		
Calcium		Fizz, giving off hydrogen gas and forming a salt
Magnesium	Very slow reaction	
Aluminium		
Zinc		
Iron		
Tin	No reaction with water at room temperature	React slowly with warm acid
Lead		
Copper	No reaction	No reaction
Silver		
Gold		

Reactivity series

Metals can be arranged in order of reactivity in a **reactivity series**.

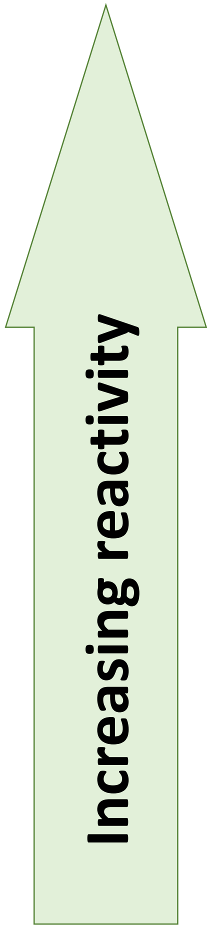
When metals react with other substances the metal atoms form **positive ions**.

The reactivity of a metal is linked to its tendency to form positive ions.

The **non-metals hydrogen** and **carbon** are often included in the series as they can be used to extract less reactive metals.

Metal + acid → salt + hydrogen

Potassium
Sodium
Lithium
Calcium
Magnesium
CARBON
Zinc
Iron
Lead
HYDROGEN
Copper
Silver
Gold



Extracting metals

The **reactivity** of a metal determines the **method** of **extraction.**

Metals **above** carbon must be extracted from their ores by using **electrolysis.**

Metals **below** carbon can be extracted from their ores by **reduction** using **carbon.**
REDUCTION involves the loss of oxygen.

metal oxide + carbon → metal + carbon dioxide

Gold and **silver** do not need to be extracted. They occur **native** (naturally).

Potassium
Sodium
Calcium
Magnesium
Aluminium
CARBON
Zinc
Iron
Lead
HYDROGEN
Copper
Silver
Gold
Platinum

Increasing reactivity

Oxidation and reduction in terms of electrons (HT)

A **more reactive metal can displace a less reactive metal from its compound** in **displacement reactions**.

Iron + copper(II) sulfate → iron sulfate + copper

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

Higher:

OILRIG

Oxidation **I**s **L**oss of electrons
Reduction **I**s **G**ain of electrons

When reactions involve **oxidation and reduction**, they are known as **redox reactions**

Higher:

An **ionic equation** shows only the atoms and ions that change in a reaction:

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

Half equations show what happens to each reactant:

$Fe \rightarrow Fe^{2+} + 2e^-$ — *The iron atoms are **oxidised** (lose 2 electrons) to form **ions**.*

$Cu^{2+} + 2e^- \rightarrow Cu$ — *The 2 **electrons** from the iron are **gained** (**reduction**) by copper ions as they become **atoms**.*

Reactions of acids - PART 1

Acids react with some **metals** to produce **salts** and **hydrogen**.

Metal + acid → salt + hydrogen

Reactions between metals and acids only occur if the metal is **more reactive** than the **hydrogen** in the acid. If the metal is too reactive, the reaction with acid is **violent**.

The **salt** that is made depends on the **metal** and **acid** used.

Salts made when metals react *nitric acid* are called *nitrates*.

Zinc + Nitric acid → Zinc Nitrate + Hydrogen

Salts made when metals react with *sulfuric acids* are called *sulfates*.

Iron + Sulfuric Acid → Iron Sulfate + Hydrogen

Salts made when metals react with *hydrochloric* acid are called *chlorides*.

Magnesium + Hydrochloric acid → Magnesium Chloride + Hydrogen

Reactions of acids - PART 1 - HIGHER

In the reaction between magnesium and hydrochloric acid, the hydrogen ions are **displaced** from the solution by magnesium as the magnesium is more reactive than hydrogen.

The following **ionic equation** occurs:

$$Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

The chloride ions are not included as they do not change in the reaction. These are known as **spectator ions.**

The reaction can be further represented by **half equations,** showing that the reaction between a metal and acid is a **redox reaction.**

$$Mg \rightarrow Mg^{2+} + 2e^{-}$$

The magnesium atoms lose two electrons, they have been **oxidised**.

$$2H^{+} + 2e^{-} \rightarrow H_2$$

The hydrogen ions have gained electrons, they have been **reduced**.

Reactions of acids - PART 1

Acids are neutralised by **alkalis** (**eg: soluble metal hydroxides**) and **bases** (**eg: insoluble metal hydroxides and metal oxides**) to produce **salts** and **water** and by **metal carbonates** to produce **salts**, **water** and **carbon dioxide**.

The salt **name** depends on the **acid** used and the **positive ions** in the **alkali, base or carbonate**.

Making Soluble Salts from acids and alkalis

Salts can be made by reacting an acid with an alkali.

Acid + Alkali → Salt + Water

Making Soluble Salts from acids and bases

Salts can be made by reacting an acid with a insoluble base.

Acid + Bases → Salt + Water

Making Soluble Salts from acids and metal carbonates

Salts can be made by reacting an acid with a metal carbonate.

Acid + Metal carbonate → Salt + Water + Carbon dioxide

Salts are made of positive metal ions (or ammonia ions - NH_4^+) and a negative ion from the acid. Like all ionic compounds, salts have **no overall charge**, so once you know the charges on the ions, you can work out the **formula**.
Example: **magnesium sulfate** is **$MgSO_4$**

ion	formula	ion	formula
Group 1	Li^+ Na^+ K^+	Transition metals	Cu^{2+} Fe^{3+}
Group 2	Mg^{2+} Ca^{2+}	Group 7	F^- Cl^- Br^-
Aluminium	Al^{3+}	Nitrate	NO_3^-
Ammonium	NH_4^+	Sulphate	SO_4^{2-}

Reactions of acids - PART 2

Indicators are substances which change **colour** when you add them to acids and alkali.

Litmus goes red in acid and blue in alkali.

Universal indicator, made from many dyes is used to tell you **pH**. The scale runs from 0 (most acidic) to 14 (most alkaline). Aqueous solutions of **acids** have a pH value **less than 7**, and for **alkalis greater than 7** and anything in the middle is **neutral** (pH 7). You can use a pH meter to record the change of a pH over time.

Acids produce **hydrogen ions (H^+)** in aqueous solutions and **alkalis** produce **hydroxide ions (OH^-).** In **neutralisation** reactions between an acid and alkali, hydrogen ions react with hydroxide ions to produce **water**.

Neutralisation symbol equation:

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

pH	Examples of solutions
0	Battery acid, strong hydrofluoric acid
1	Hydrochloric acid secreted by stomach lining
2	Lemon juice, gastric acid, vinegar
3	Grapefruit juice, orange juice, soda
4	Tomato juice, acid rain
5	Soft drinking water, black coffee
6	Urine, saliva
7	"Pure" water
8	Sea water
9	Baking soda
10	Great Salt Lake, milk of magnesia
11	Ammonia solution
12	Soapy water
13	Bleach, oven cleaner
14	Liquid drain cleaner

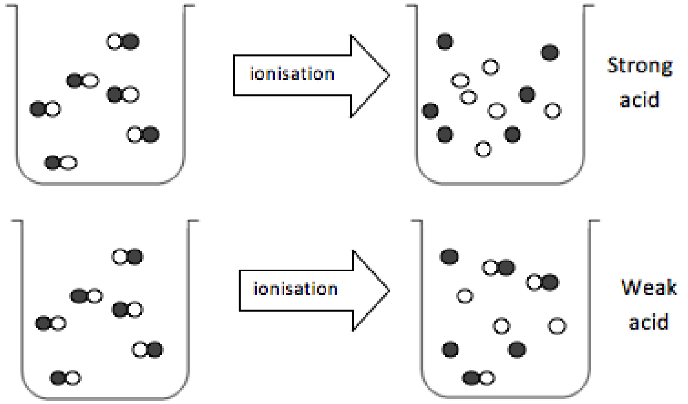
Reactions of acids - PART 2 - HIGHER

Acids must **dissolve** in water to show their acidic properties.
A **concentrated acid** has a relatively **large amount of solute** dissolved in the solvent.
A **dilute acid** has a relatively **smaller amount of solute** dissolved in the solvent

The molecules split up to form **hydrogen** ions.

A **strong** acid is **completely ionised** in aqueous solution. E.g. Hydrochloric, nitric and sulfuric acid.

A **weak** acid is only **partially ionised** in aqueous solution. E.g. Ethanoic, citric and carbonic.



A **weak acid** (aq) has a **lower pH** than a **strong acid** (aq) of the same concentration.

This is because a **weak acid** has a **lower concentration of hydrogen ions.**

As the pH decrease by one unit, the hydrogen ion concentration of the solution increase by a factor of 10.

Concentration of hydrogen ions in mol/dm^3	pH
0.10	1.0
0.010	2.0
0.0010	3.0
0.00010	4.0

Electrolysis

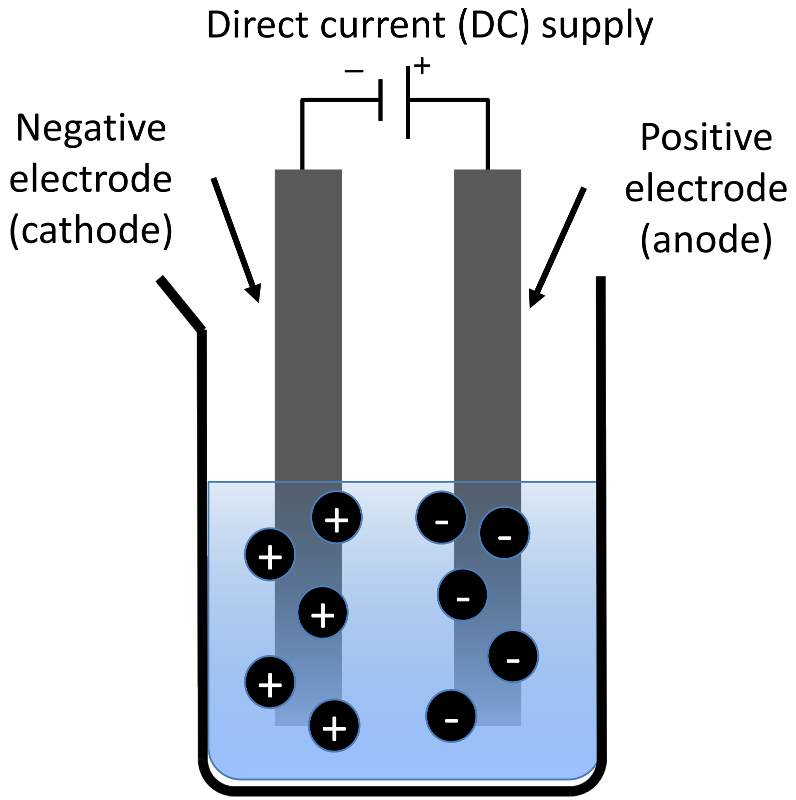
Recall it …

Use the information in the following page(s) to answer these questions …

1) Describe what is an electrolyte?
2) What is the anode and cathode?
3) Describe what happens at the anode?
4) Describe what happens at the cathode?
5) What happens during electrolysis?
6) Describe what happens at the anode and cathode during the electrolysis of lead bromide?
7) Describe what happens at the anode and cathode during the electrolysis of sodium chloride?
8) What are the uses of the products from the electrolysis of sodium chloride?
9) Why is the electrolysis of aluminium oxide expensive?
10) What is mixed with aluminium to reduce the temperature to heat aluminium oxide?
11) Describe what is produced at the anode and cathode during the electrolysis of aluminium oxide?

Electrolysis - PART 1

When an **ionic compound** is **melted** or **dissolved in water**, the **ions** are **free** to **move** about the liquid or solution. These liquids and solutions are able to **conduct electricity** and are called **electrolytes**. Passing an **electric current** though electrolytes causes the ions to move to the electrodes.



Positive ions go to **negative** electrode (cathode) and are **reduced** (gain of electrons).

Negative ions go to the **positive** electrode (anode) and are **oxidised** (loss of electrons).

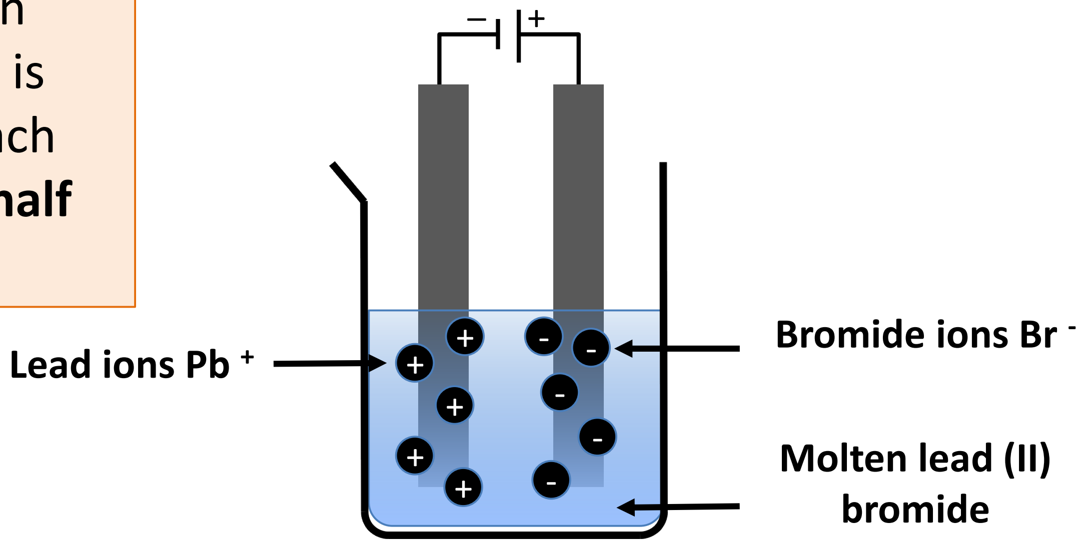
Ions are **discharged** at the electrodes producing **elements**. This is called **electrolysis**.

Electrolysis - PART 1

When an **ionic compound** is electrolysed in a **molten** state using inert electrodes, the **metal** is produced at the **cathode** and the **non-metal** is produced at the **anode**.

Higher: You can represent what is happening at each electrode using **half equations**.

lead bromide → lead + bromine



The positively charged lead ions Pb + (cations) are attracted to cathode and the negatively charged bromide ions Br - are attracted to the anode.

Higher:
At the cathode
$Pb^{2+} + 2e^{-} \rightarrow Pb$

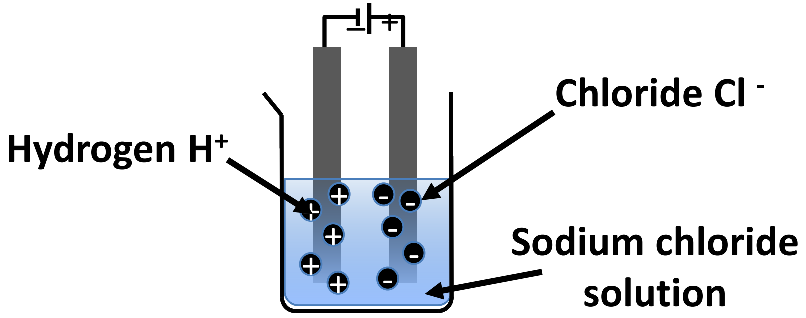
Higher:
At the anode
$2Br^{-} \rightarrow Br_2 + 2e^{-}$ or $2Br^{-} - 2e^{-} \rightarrow Br_2$

Electrolysis - PART 1

The **ions** discharged when an **aqueous solution** is electrolysed using inert electrodes depend on the relative **reactivity** of the elements involved.

At the **negative** electrode:

Metal will be produced on the electrode if it is **less** reactive than **hydrogen**.
Hydrogen will be produced if the metal is **more** reactive than hydrogen.



sodium chloride → hydrogen + chlorine + sodium hydroxide

Uses of the products:
Chlorine: Bleach and PVC
Hydrogen: Margarine
Sodium hydroxide: Bleach and soap

At the **positive** electrode:

Oxygen is formed at positive electrode.

Higher: At the anode
$4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
or $4OH^- - 4e^- \rightarrow O_2 + 2H_2O$

If you have a **halide** ion (Cl^-, I^-, Br^-) then you will get **chlorine, bromine or iodine** formed at that electrode.
This happens because in the aqueous solution, **water molecules** break down producing **hydrogen** ions and **hydroxide** ions that are discharged.
$H_2O(l) \rightarrow H^+(aq) + OH^-(aq)$

Higher:
At the cathode
$2H^+ + 2e^- \rightarrow H_2$

Higher:
At the anode
$2Cl^- \rightarrow Cl_2 + 2e^-$ or $2Cl^- - 2e^- \rightarrow Cl_2$

Electrolysis - PART 2

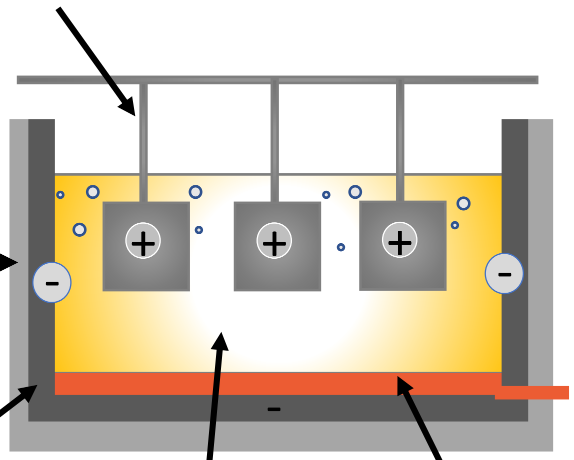
Metals can be extracted from **molten compounds** using electrolysis.

It is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon.

Large amounts of **energy** are used in the extraction process to melt the compounds and to produce the electrical current.

Aluminum is manufactured by electrolysis of molten aluminum oxide.

Aluminium oxide → aluminium + oxygen



Aluminium oxide has a very **high melting point** so is mixed with molten **cryolite** to lower the temperature required to carry out the electrolysis.

Aluminium goes to the negative electrode and sinks to bottom.

Higher: $Al^{3+} + 3e^- \rightarrow Al$

Oxygen forms at positive electrodes.

Higher: $2O^{2-} \rightarrow O_2 + 4e^-$

The oxygen reacts with the carbon electrode making carbon dioxide causing damage. The electrode needs **replacing** due to this reaction.

$C + O_2 \rightarrow CO_2$

Endothermic and Exothermic reactions

Use the information in the following page(s) to answer these questions …

1) Describe what is an exothermic reaction?
2) Give three examples of exothermic reactions?
3) Describe what is an endothermic reaction?
4) Give two examples of endothermic reactions?
5) What is activation energy?
6) Draw a reaction profile and label the activation energy?
7) Draw the reaction profile for an exothermic reaction?
8) Draw the reaction profile for an endothermic reaction?
9) How is a double bond represented?
10) What happens to bonds in a chemical reaction?
11) What is bond energy?
12) Describe how to calculate the overall energy change using bond energy?

Exothermic and endothermic reactions part 1 – Exothermic reactions

Energy is conserved in chemical reactions. The amount of energy in the Universe at the end of a chemical reaction is the same as before the reaction takes place.

$$H_2\,(g) + Cl_2\,(g) \rightarrow 2HCl\,(g)$$

In the above reaction energy is released, it gets hotter.

An exothermic reaction is one that transfers energy to the surroundings so the temperature of the surroundings increases – "it gets hotter".

The two HCl molecules made will not hold as much energy as the H_2 and Cl_2 molecules at the start, so the spare energy is released as heat.

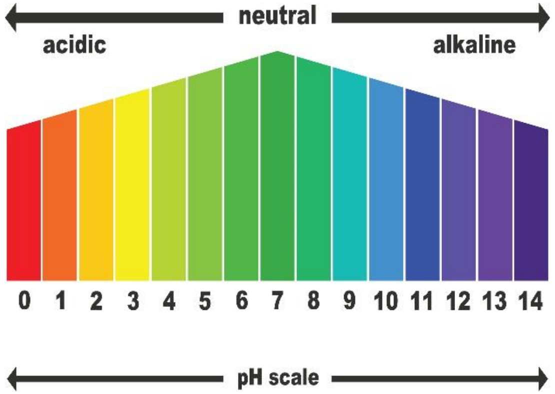
Exothermic and endothermic reactions part 1 – Exothermic reactions

There are a number of common exothermic reactions, they include:

Combustion

Oxidation

Neutralisation



Know all three of these examples of exothermic reactions

Everyday uses of exothermic reactions include –
Self-heating cans
Hand warmers

Know both of these uses for exothermic reactions

Exothermic and endothermic reactions part 1 – Endothermic reactions

We have already learnt that energy is conserved in chemical reactions.

$$2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O\ (l)$$

In the above reaction, energy is taken in- it gets colder.
An endothermic reaction is one that takes energy from the surroundings so the temperature of the surroundings decreases – "it gets colder".
The sodium ethanoate, carbon dioxide and water molecules made will hold more energy than the ethanoic acid and sodium carbonate molecules at the start, so the energy needed is taken in as heat.

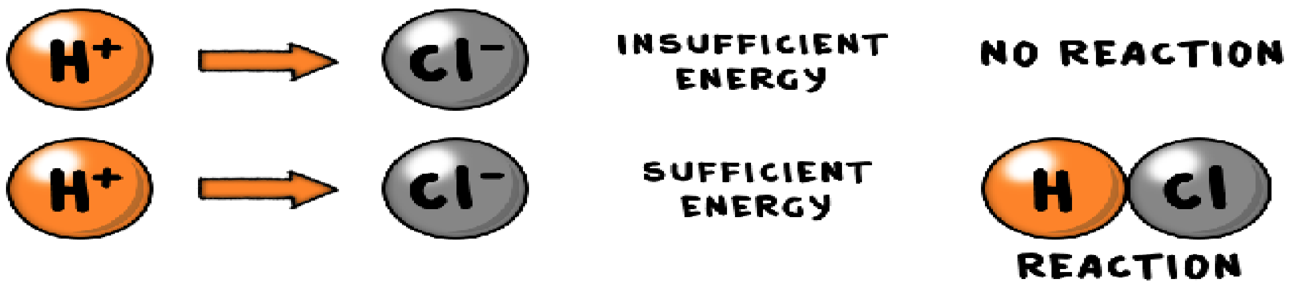
Other examples of endothermic reactions are

- Thermal decomposition
- Sports injury packs

Know all three of these examples of endothermic reactions

Energy Changes part 2 – Reaction profiles

Chemical reactions can only occur when reacting particles collide with each other with sufficient energy.



The minimum amount of energy that particles must have to react is called the activation energy

You have given a reaction its activation energy when you have used a lit spill to light a Bunsen burner. Without the activation energy from the lit spill the methane gas and oxygen in the air will not combust and release the heat energy.

When we look at this reaction we see the following.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

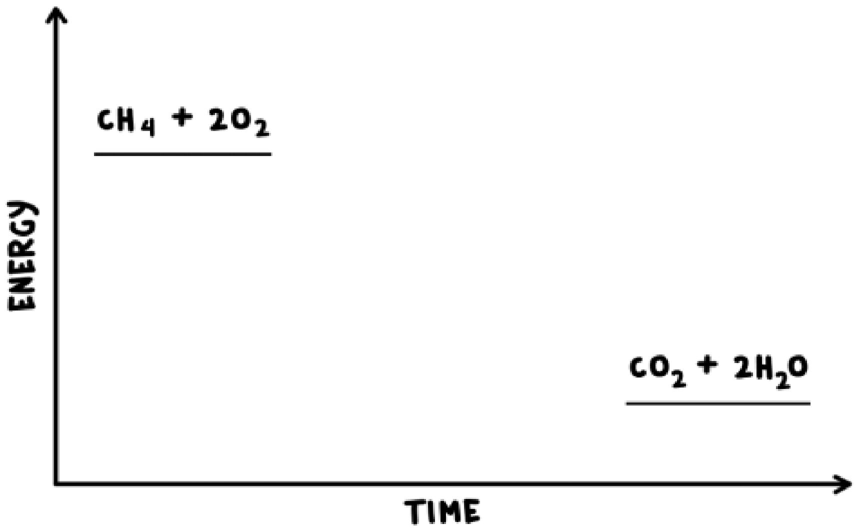
You will be expected to balance this equation.

Energy Changes part 2 – Reaction profiles

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

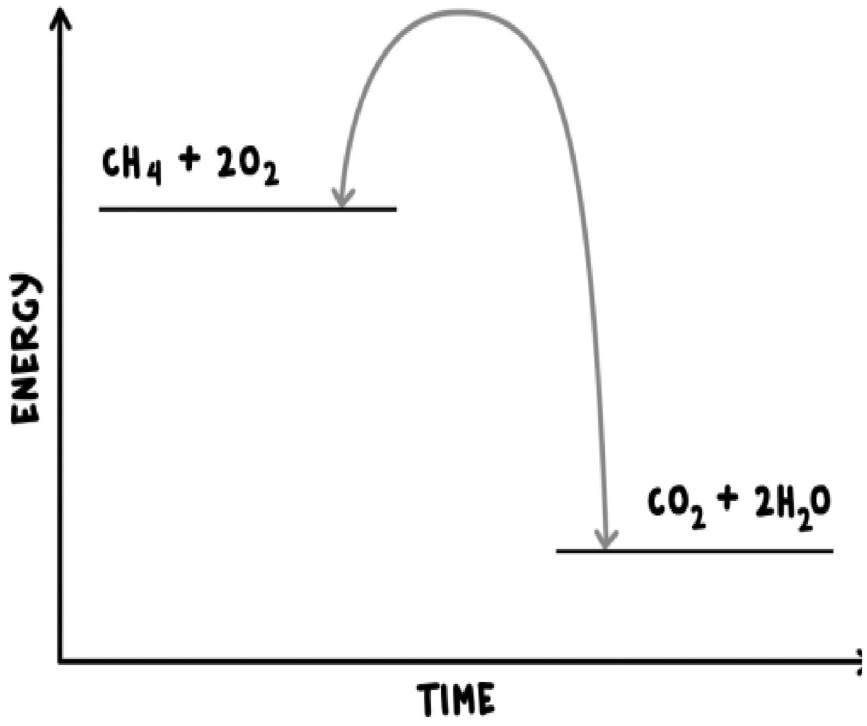
We know this reaction is exothermic, this means energy is released. So the CH_4 and $2O_2$, the reactants, must have more energy than the products, CO_2 and $2H_2O$

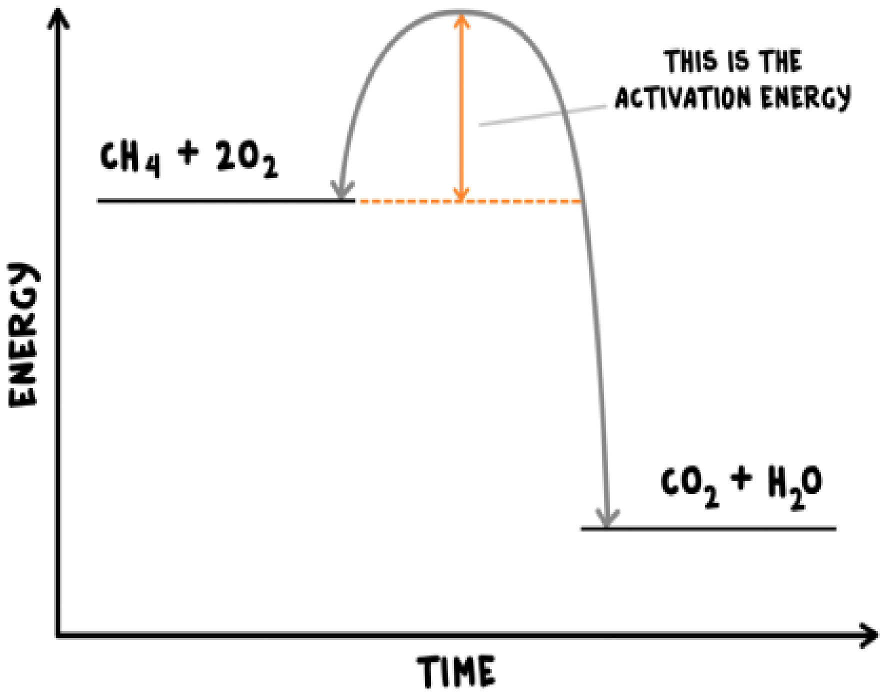
We can show this as a reaction profile. On it we need to include the formulae or names of the products and reactants. We also need to show the relative energies of the reactants and products



Energy Changes part 2 – Reaction profiles

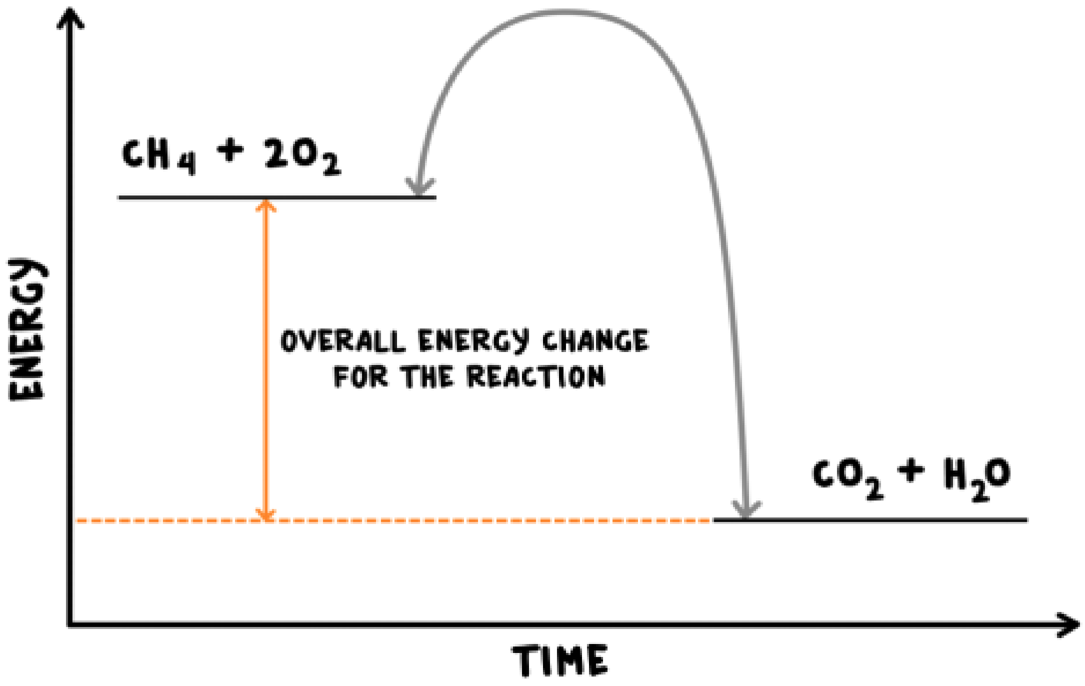
More information needs to be included in the reaction profile. This will show the activation energy of the reaction. It is shown by a curved line rising above the reactants energy.





Energy Changes part 2 – Reaction profiles

We can now see the overall change in energy within the reaction.



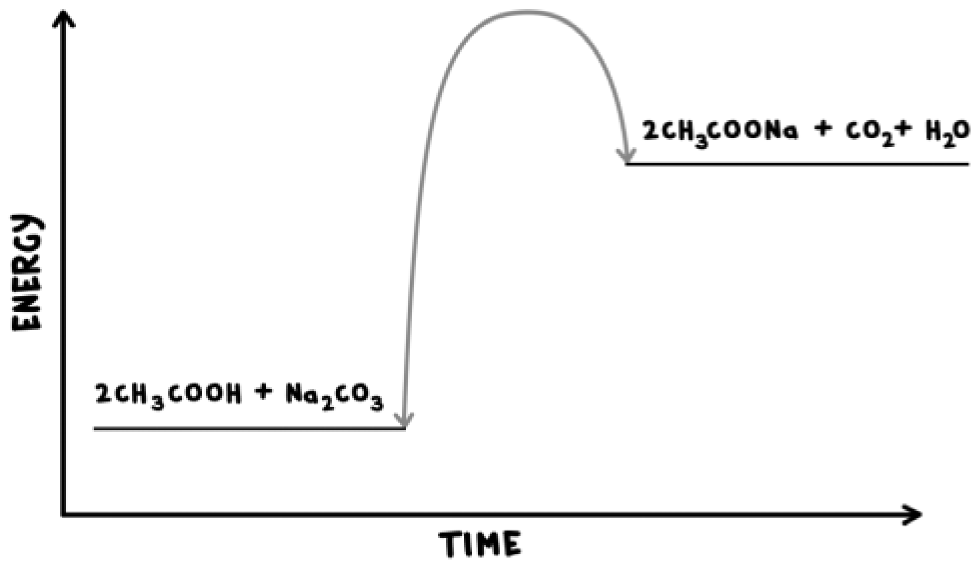
The products have less energy than the reactants. This will have been lost as heat as the reaction is exothermic.

Energy Changes part 2 – Reaction profiles

We saw earlier that the following reaction was endothermic:

$2CH_3COOH(aq) + Na_2CO_3(s) \rightarrow 2CH_3COONa(aq) + CO_2(g) + H_2O\ (l)$

What would the reaction profile look like for this reaction?



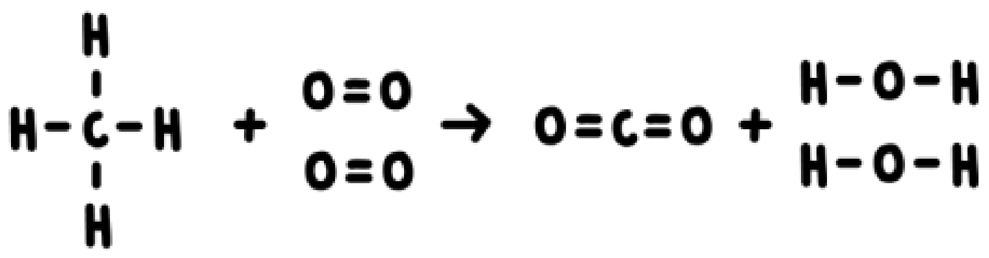
We can see that the products have more energy than the reactants. This will have been taken in as heat energy it feels colder.

The energy change of reactions (HT only)

For the reaction of methane with oxygen we can write out the balanced symbol equation:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

We can draw out the bonds between the atoms e.g.



Each line represents a bond, two lines represent a double bond e.g. in the oxygen molecule.

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The energy change of reactions (HT only)

During a chemical reaction:

- **Energy must be supplied to break bonds in the reactants**
- **Energy is released to form bonds in the products.**

The energy needed to break bonds and the energy released when bonds are formed can be calculated from bond energies.

Bond	Bond Energy kJ/mol
C-H	411
O=O	494
C=O	799
O-H	459

This means that 411kJ/mol of energy needs to be put in to break the carbon-hydrogen bond.
It also means that 459 kJ/mol is given out when the oxygen hydrogen bond is made in water.

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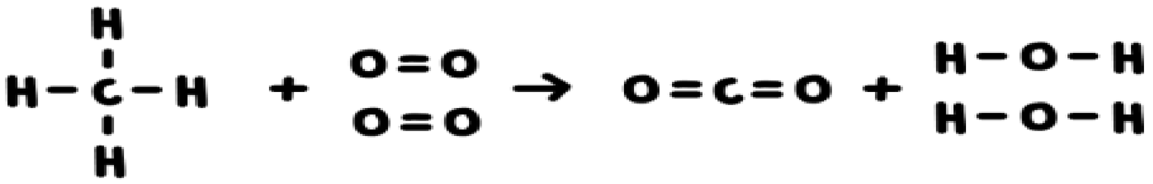
The energy change of reactions (HT only)

The **difference** between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the **overall energy change** of the reaction.

Worked example



For the reactants	For the products
There are four C-H bonds so 4 x 411kJ/mol = 1,644kJ/mol	There are two C=O bonds so 2 x 799kJ/mol = 1,598kJ/mol
There are two O=O bonds so 2 x 494kJ/mol = 988kJ/mol	There are four O-H bonds so 4 x 459kJ/mol = 1,836kJ/mol
The sum of these is the energy supplied to break the bonds in the reactants it is 1,644 + 988 = 2,632kJ/mol	The sum of these is the energy released when bonds in the products are formed it is 1,598 + 1,836 = 3,434kJ/mol

The energy change of reactions (HT only)

We already know that the difference between the sum of the energy needed to break bonds in the reactants and the sum of the energy released when bonds in the products are formed is the overall energy change of the reaction.

This means that

Overall energy change = energy needed to break the bonds – energy released as bonds are made

= 2,632kJ/mol - 3,434kJ/mol

Overall energy change = - 802kJ/mol

This is an exothermic reaction, so the sum of the difference between the calculations is negative. For an endothermic reaction it would be positive.

Students should be able to calculate the energy transferred in chemical reactions using bond energies supplied.

The energy change of reactions (HT only)

Know these two definitions- they are often asked for in the exam.

In an exothermic reaction, the energy released from forming new bonds is greater than the energy needed to break existing bonds

In an endothermic reaction, the energy needed to break existing bonds is greater then the energy released from forming new bonds