7. Chemical reactions

7.1.1 Physical & Chemical Changes

Physical & Chemical Change

Physical change

- Physical changes (such as **melting** or **evaporating**) do not produce any new chemical substances
- These changes are often **easy to reverse** and mixtures produced are usually relatively easy to separate

Chemical change

- In chemical reactions, **new chemical** products are formed that have very different **properties** to the reactants
- Most chemical reactions are impossible to **reverse**
- Energy changes also accompany chemical changes and energy can be given out (**exothermic**) or taken in (**endothermic**)
- The majority of chemical reactions are exothermic with only a small number being endothermic

7.1.2 Rate (Speed) of Reaction

Rates of Reaction Factors

Effect of concentration

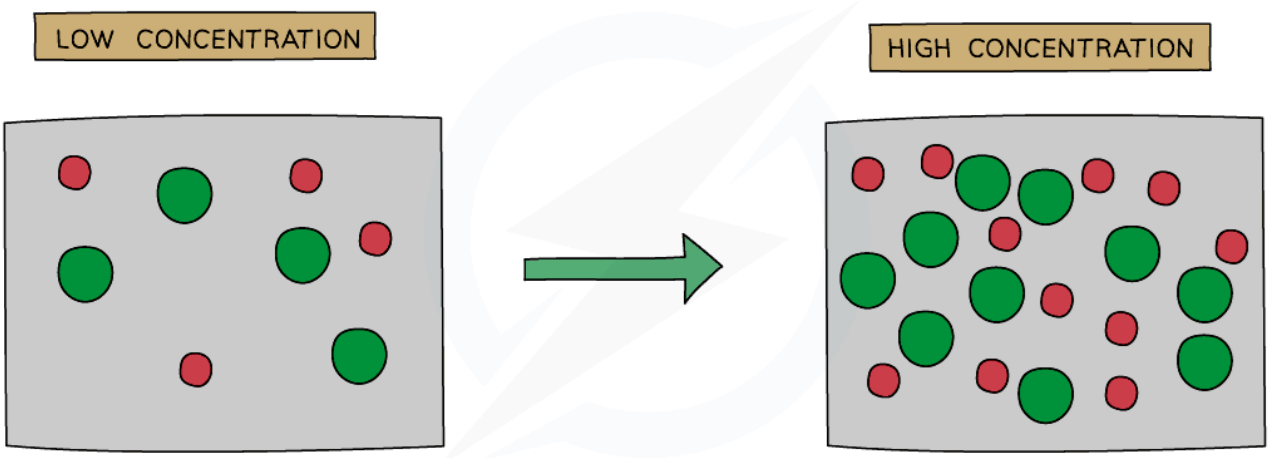


Diagram showing increase in concentration of solution

Explanation:

- **Increase in the concentration of a solution, the rate of reaction will increase**
- This is because there will be more reactant particles in a given volume, allowing more frequent and successful collisions per second, increasing the rate of reaction

Effect of surface area

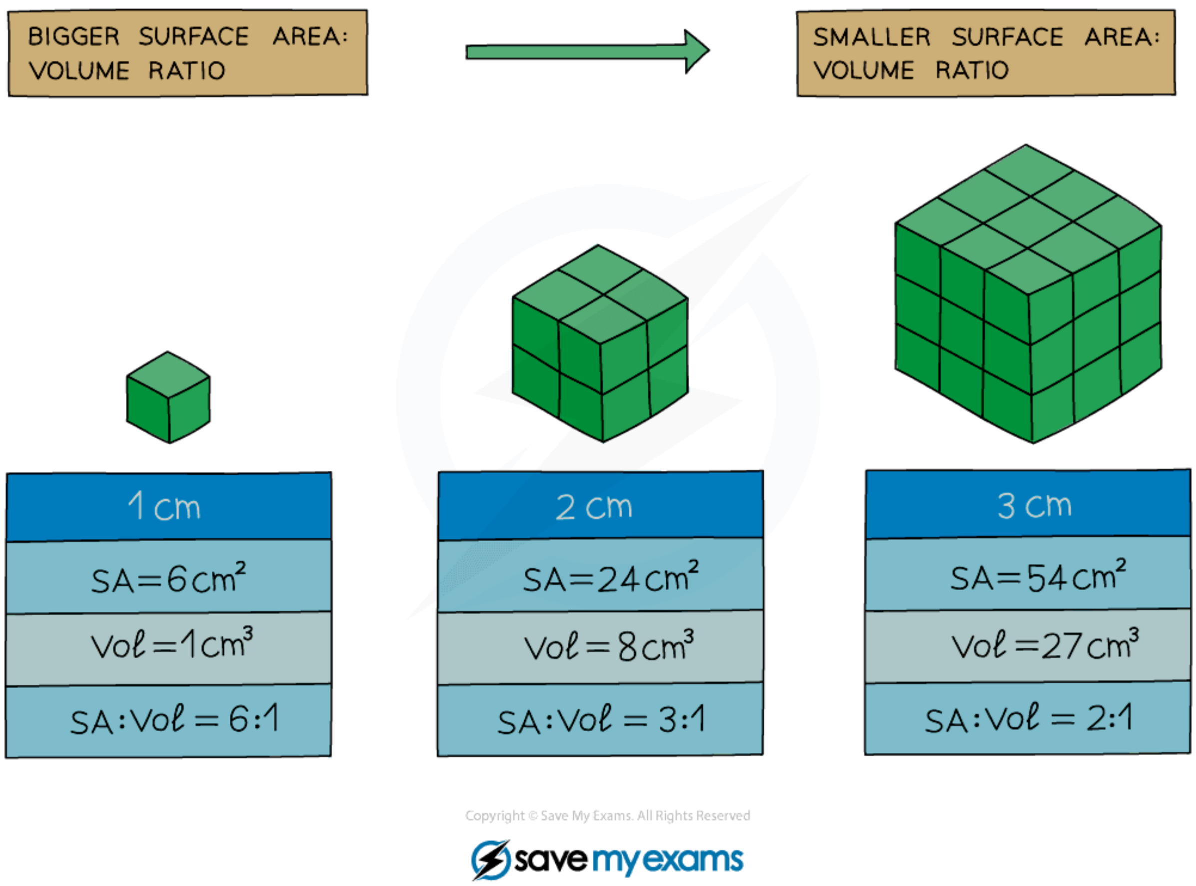


Diagram showing surface area to volume ratio of various sized cubes

Explanation:

- **Increase in the surface area of the solid, the rate of reaction will increase**
- This is because more surface area particles will be exposed to the other reactant so there will be more frequent and successful collisions per second, increasing the rate of reaction

Effect of temperature

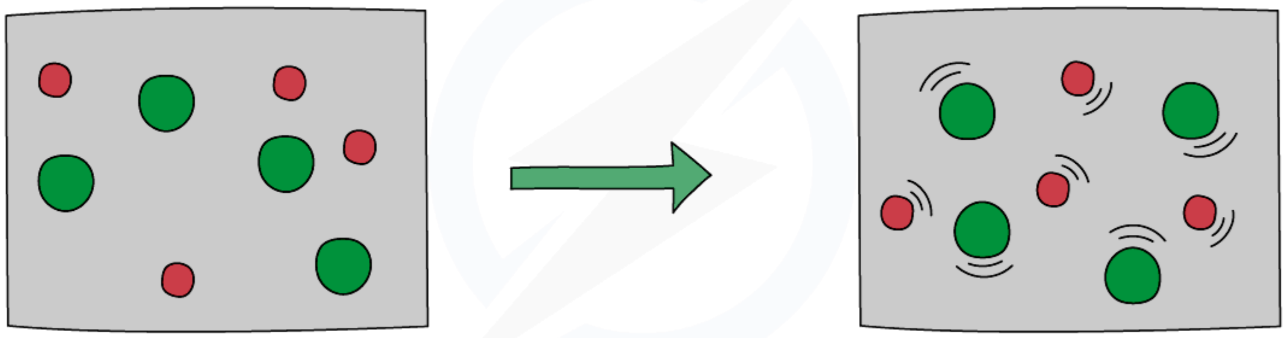
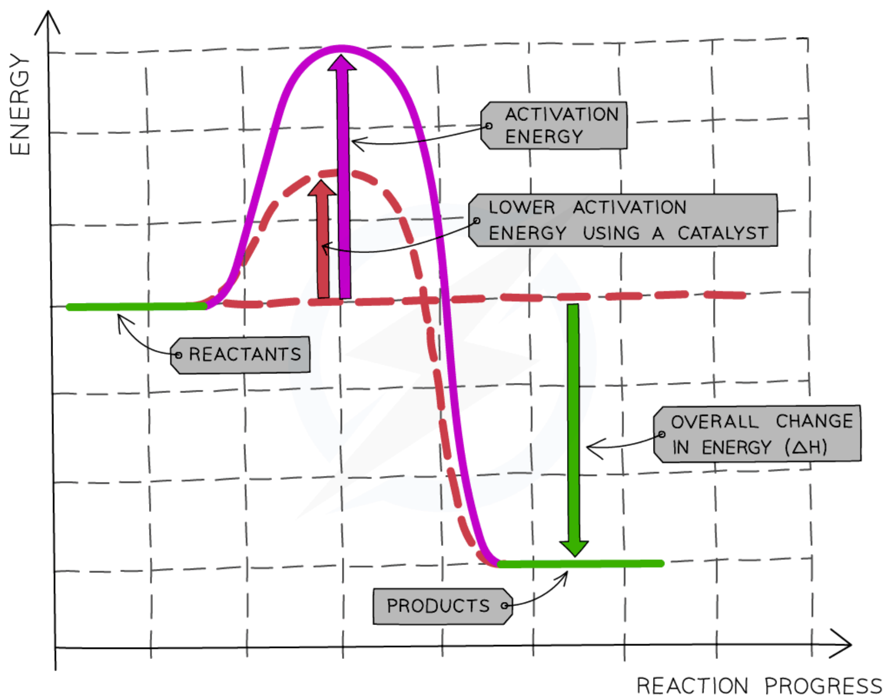


Diagram showing the effect of temperature on particles

- **Increase in the temperature, the rate of reaction will increase**
- This is because the particles will have more kinetic energy than the required activation energy, therefore there will be more frequent and successful collisions per second, increasing the rate of reaction

Effect of using a catalyst



Graph showing the effect of the use of a catalyst on the rate of reaction

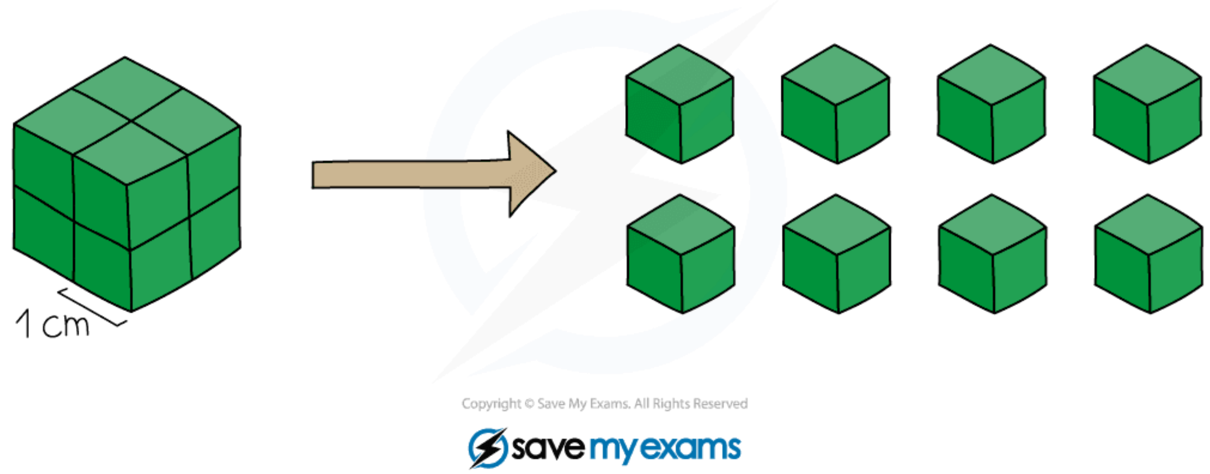
Explanation:

- Catalysts reduce the activation energy as they create alternative pathways requiring lower activation energy, allowing more successful and frequent collisions
- **This shows that when a catalyst is used, the rate of reaction will increase**

Explosive combustion

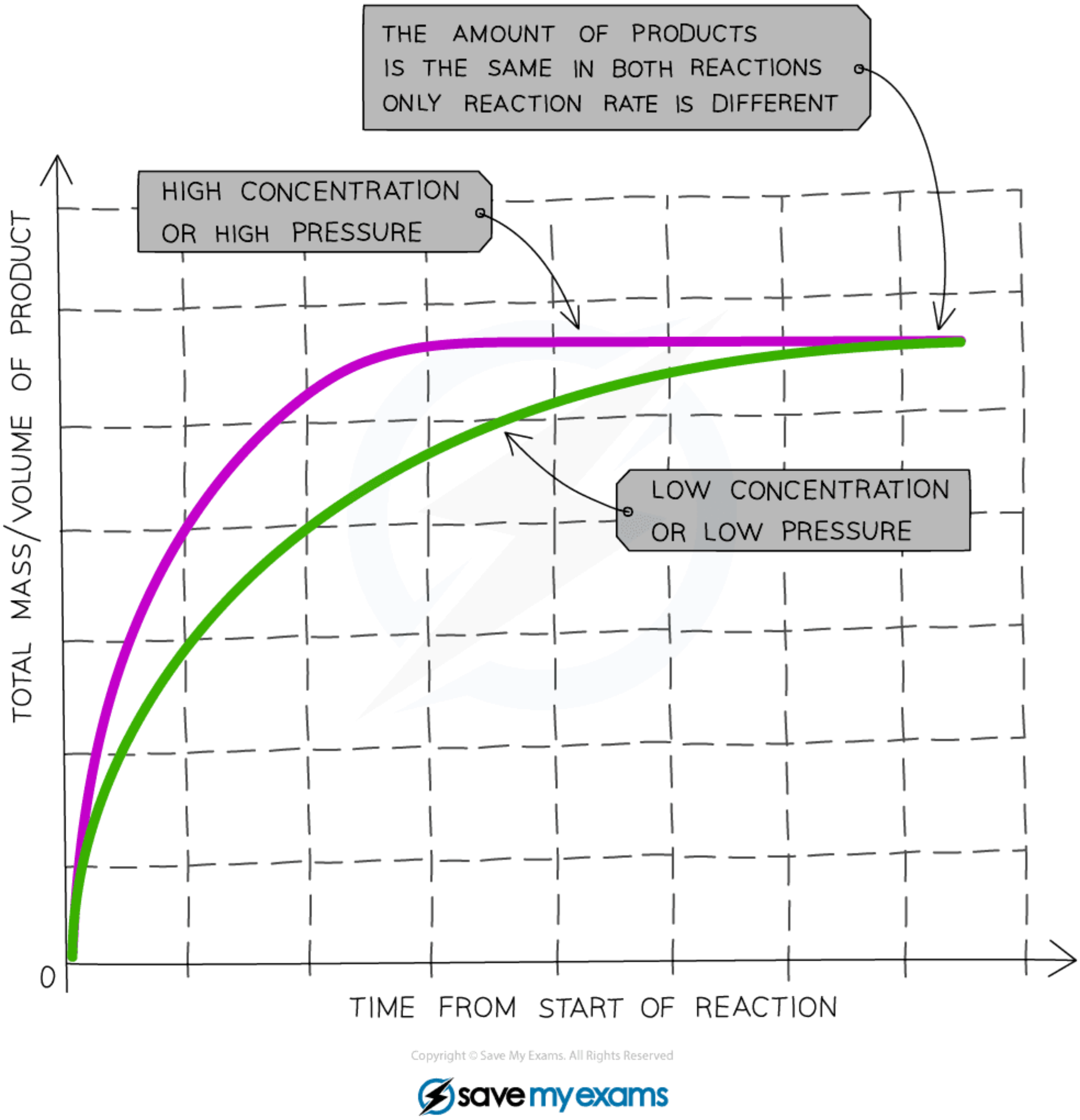
- Explosive combustion occurs when there are many **fine particles** in the air
- Many industrial processes such as metal working, coal mining or flour milling produce very fine and tiny particles
- These particles have a very large **surface area** and are **combustible** in air
- Even a small spark may cause them to ignite and since the surface area is so large, the rate of reaction can be incredibly **fast,** hence they are explosive
- Methane gas mixed with air in coal mines can also form an explosive mixture

Interpreting Data



A single 2 cm length cube has a surface area of 2 x 2 x 6 = 24cm^2. Cutting it into 8 x 1cm cubes means it now has a surface area of 1 x 1 x 6 x 8 = 48cm^2

Concentration

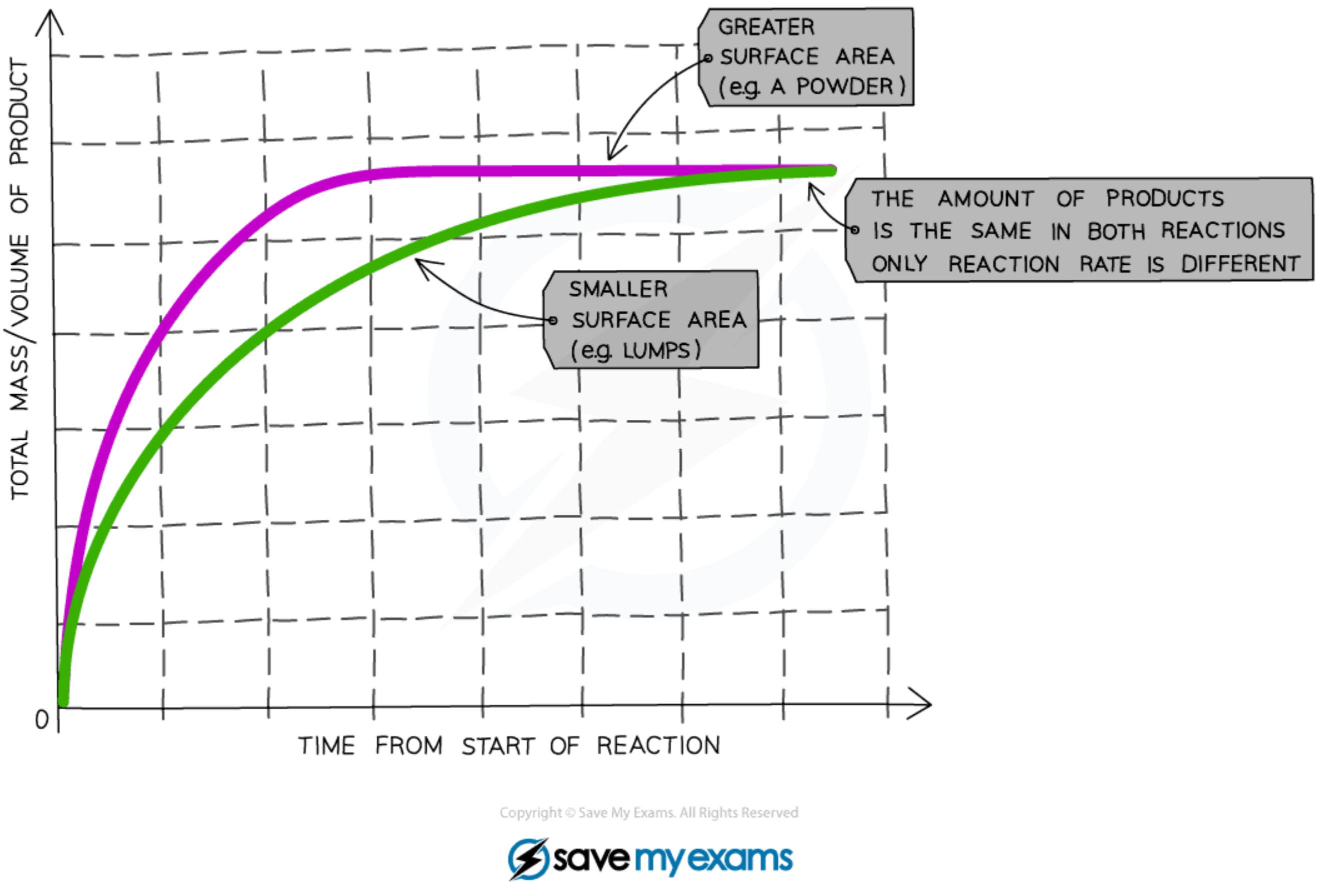


Graph showing the effect of the concentration of a solution on the rate of reaction

Explanation:

- Compared to a reaction with a reactant at a low concentration, the graph line for the same reaction but at a higher concentration has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with increased concentration of a solution, the rate of reaction will increase**

Particle size

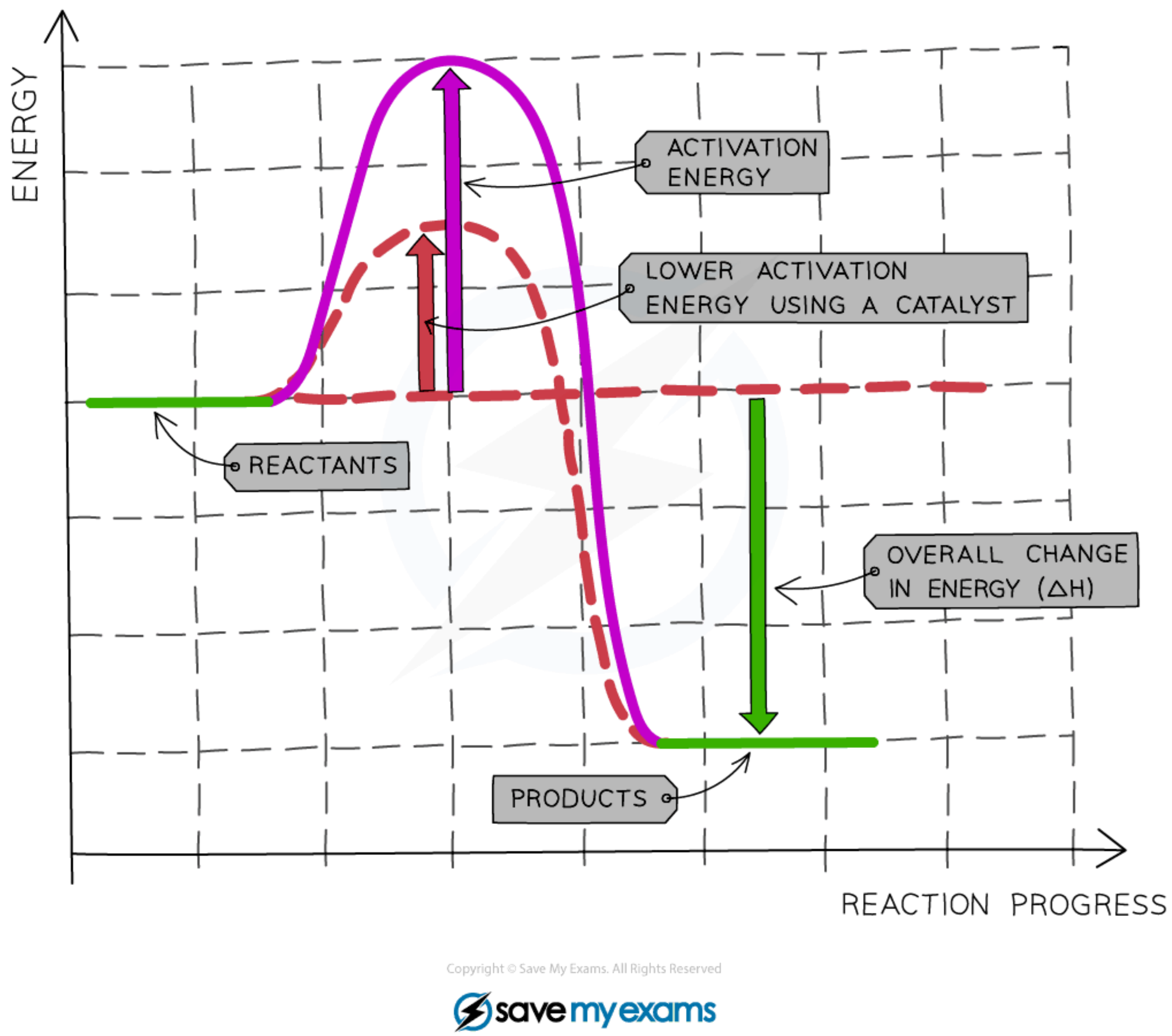


Graph showing the effect of the surface area of a solid on the rate of reaction

Explanation:

- Compared to a reaction with lumps of reactant, the graph line for the same reaction but with powdered reactant has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with increased surface area of the solid, the rate of reaction will increase**

Catalyst

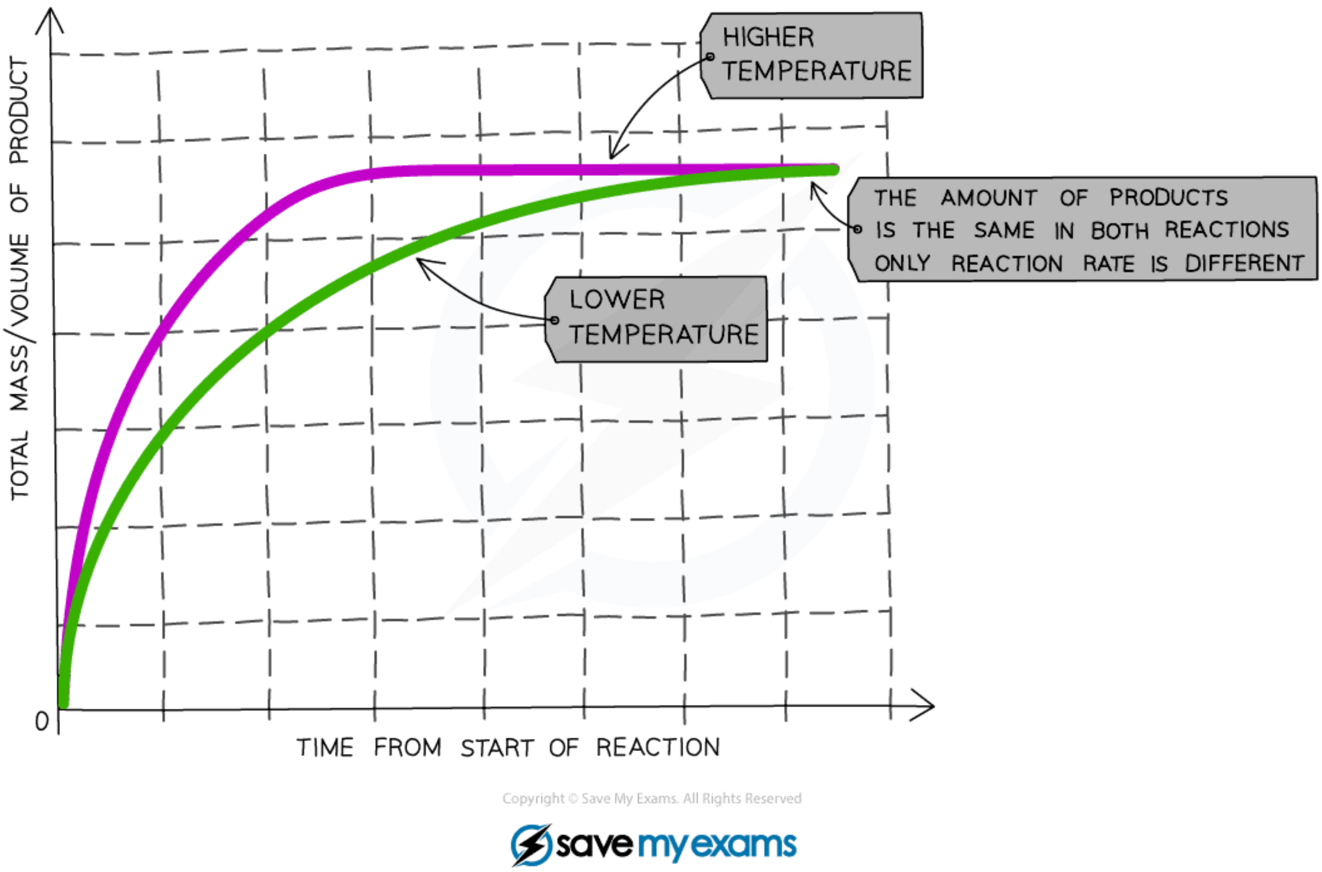


Graph showing the effect of the use of a catalyst on the rate of reaction

Explanation:

- The diagram shows that when a catalyst is used, the activation energy is reduced as it creates an alternative pathway requiring lower activation energy, allowing more successful and frequent collisions
- **This shows that when a catalyst is used, the rate of reaction will increase**

Temperature



Graph showing the effect of temperature on the rate of reaction

Explanation:

- Compared to a reaction at a low temperature, the graph line for the same reaction but at a higher temperature has a steeper gradient at the start and becomes horizontal sooner
- **This shows that with increased temperature, the rate of reaction will increase**

Extended Only

Investigating the Rate of a Reaction

Effect of surface area of a solid on the rate of reaction:

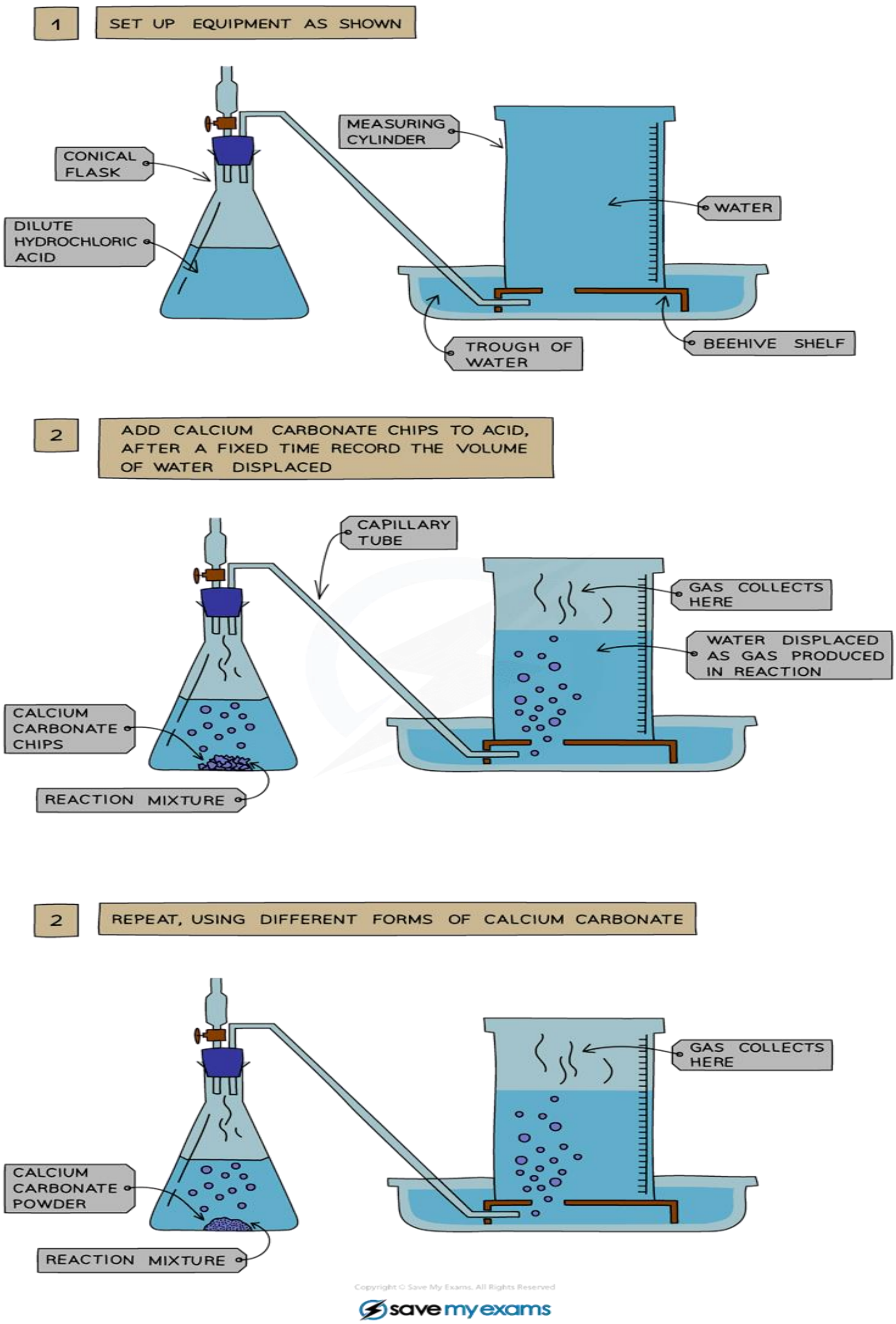


Diagram showing the process of downwards displacement to investigate the effect of the surface area of a solid on the rate of reaction

Method:

- Add dilute hydrochloric acid into a conical flask
- Use a capillary tube to connect this flask to a measuring cylinder upside down in a bucket of water (downwards displacement)
- Add calcium carbonate chips into the conical flask and close the bung
- Measure the volume of gas produced in a fixed time using the measuring cylinder
- Repeat with different sizes of calcium carbonate chips (solid, crushed and powdered)

Result:

- Smaller sizes of chips causes an increase in the surface area of the solid, so the rate of reaction will increase
- This is because more surface area of the particles will be exposed to the other reactant so there will be more frequent and successful collisions, increasing the rate of reaction

Effect of concentration of a solution on the rate of reaction:

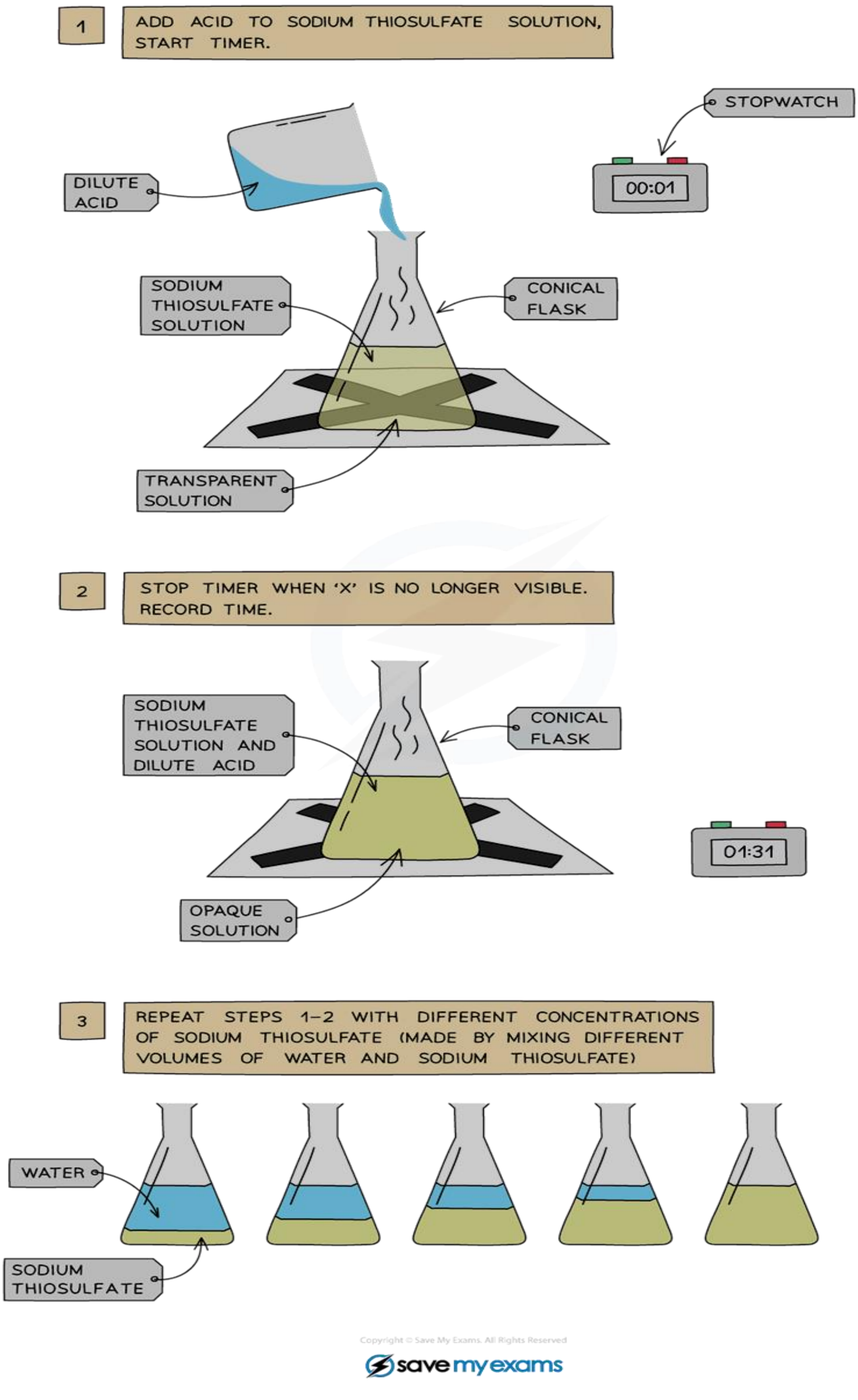


Diagram showing the apparatus needed to investigate the effect of concentration on the rate of reaction

Method:

- Measure 50 cm^3 of Sodium Thiosulfate solution into a flask
- Measure 5 cm^3 of dilute Hydrochloric acid into a measuring cylinder
- Draw a cross on a piece of paper and put it underneath the flask
- Add the acid into the flask and immediately start the stopwatch
- Look down at the cross from above and stop the stopwatch when the cross can no longer be seen
- Repeat using different concentrations of Sodium Thiosulfate solution (mix different volumes of sodium thiosulfate solution with water to dilute it)

Result:

- With an increase in the concentration of a solution, the rate of reaction will increase
- This is because there will be more reactant particles in a given volume, allowing more frequent and successful collisions, increasing the rate of reaction

Effect of temperature on the rate of reaction:

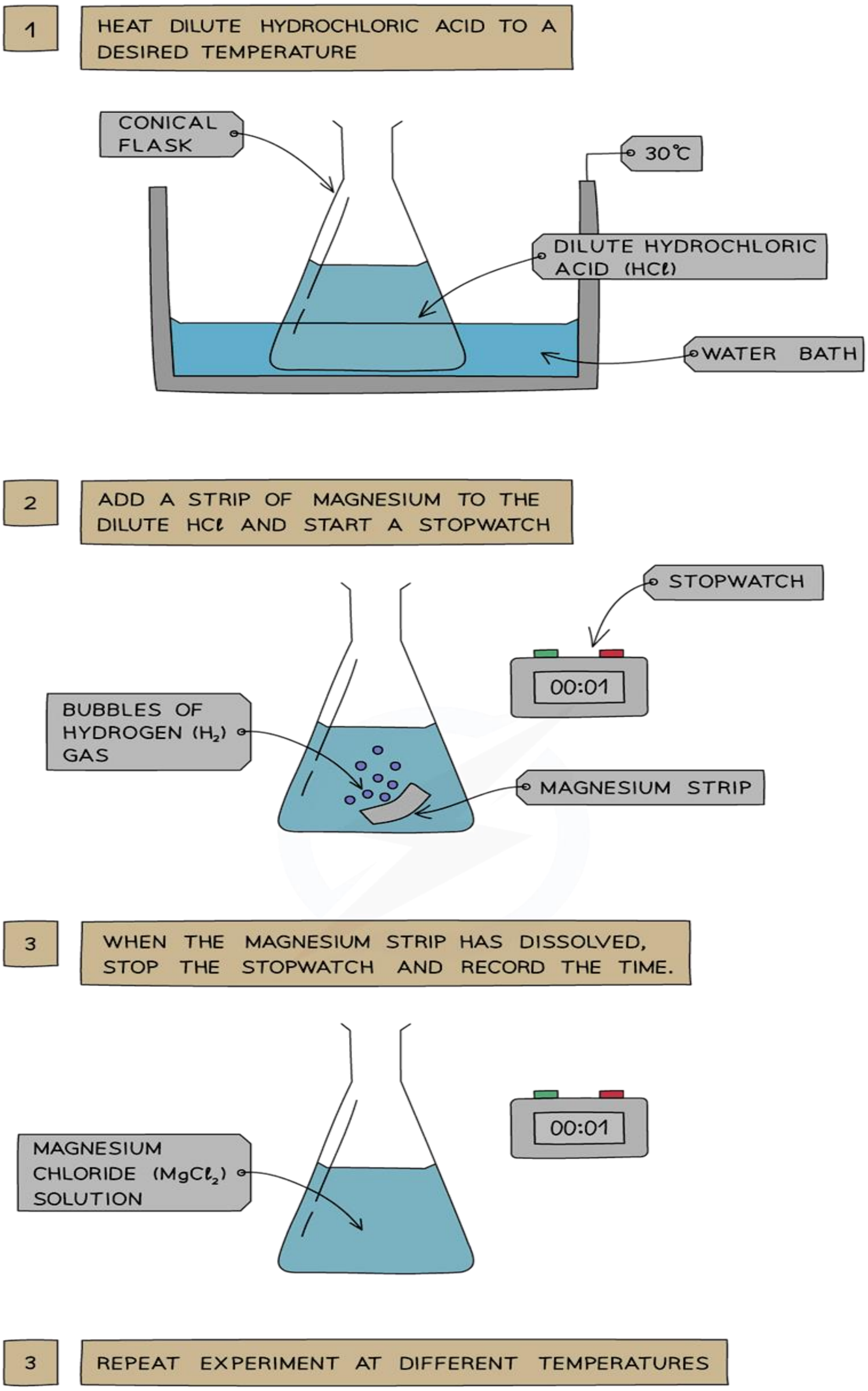


Diagram showing the apparatus needed to investigate the effect of temperature on the rate of reaction

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Method:

- Dilute Hydrochloric acid is heated to a set temperature using a water bath
- Add the dilute Hydrochloric acid into a conical flask
- Add a strip of Magnesium and start the stopwatch
- Stop the time when the Magnesium fully dissolves
- Repeat at different temperatures and compare results

Result:

- With an increase in the temperature, the rate of reaction will increase
- This is because the particles will have more kinetic energy than the required activation energy, therefore more frequent and successful collisions will occur, increasing the rate of reaction

Effect of a catalyst on the rate of reaction:

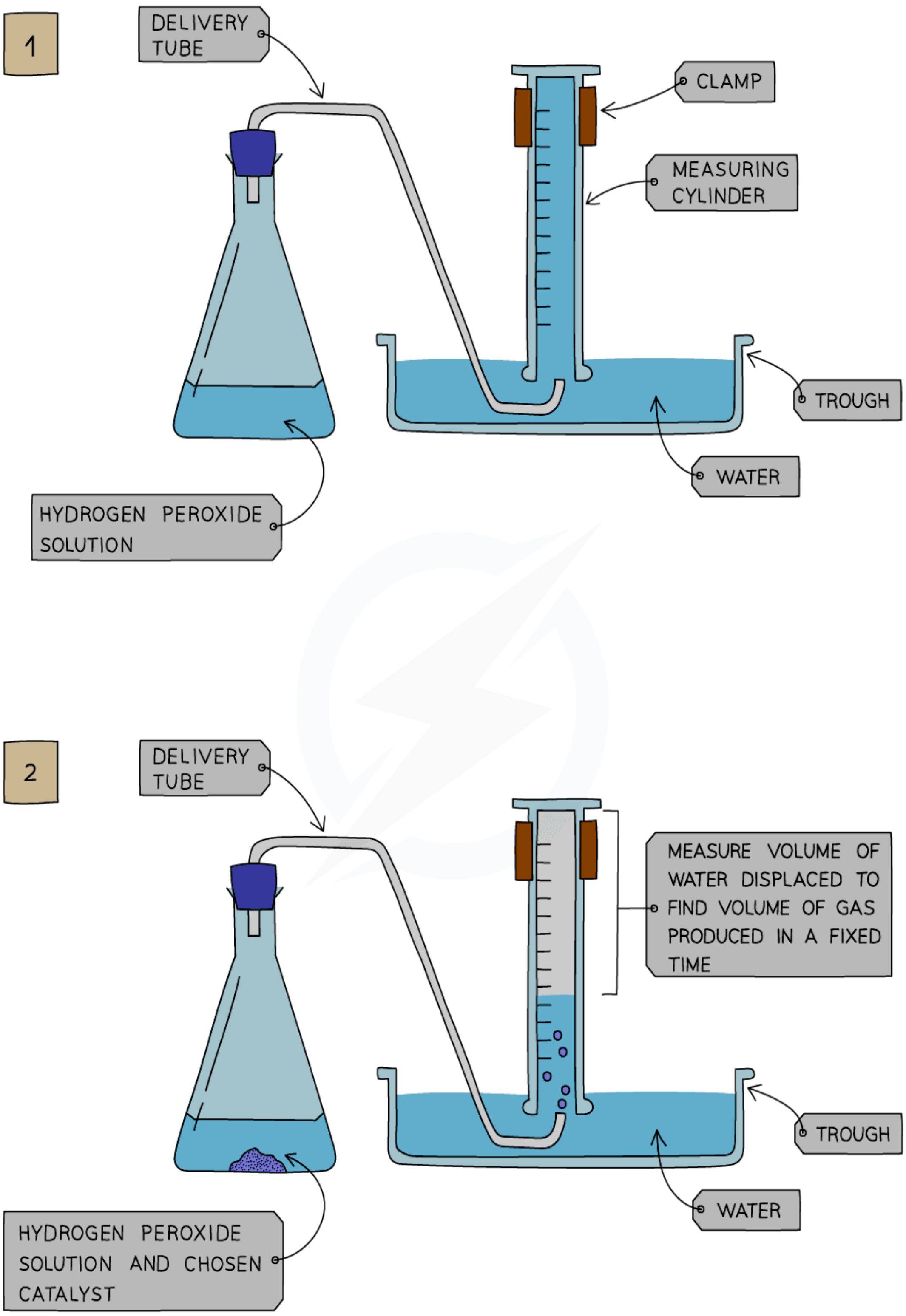

Diagram showing the apparatus needed to investigate the effect of a catalyst on the rate of reaction

Method:

- Add Hydrogen Peroxide into a conical flask
- Use a capillary tube to connect this flask to a measuring cylinder upside down in a bucket of water (downwards displacement)
- Add the catalyst Manganese(IV) Oxide into the conical flask and close the bung
- Measure the volume of gas produced in a fixed time using the measuring cylinder
- Repeat experiment without the catalyst of Manganese(IV) Oxide and compare results

Result:

- Using a catalyst will increase the rate of reaction
- The catalyst will provide an alternative pathway requiring lower activation energy so more colliding particles will have the necessary activation energy to react
- This will allow more frequent and successful collisions, increasing the rate of reaction

Extended Only

Temperature & Concentration

Temperature

- Particles need to have a **minimum amount of energy** to react when they collide
- This is called the **activation energy**
- At low temperatures only a small number of particles will have enough activation energy so the reaction will be **slow**
- At higher temperatures the particles have **more** kinetic energy so they move faster and with more energy
- The collisions are thus more energetic and there is a greater number of particles with sufficient energy to react, so the rate of reaction increases

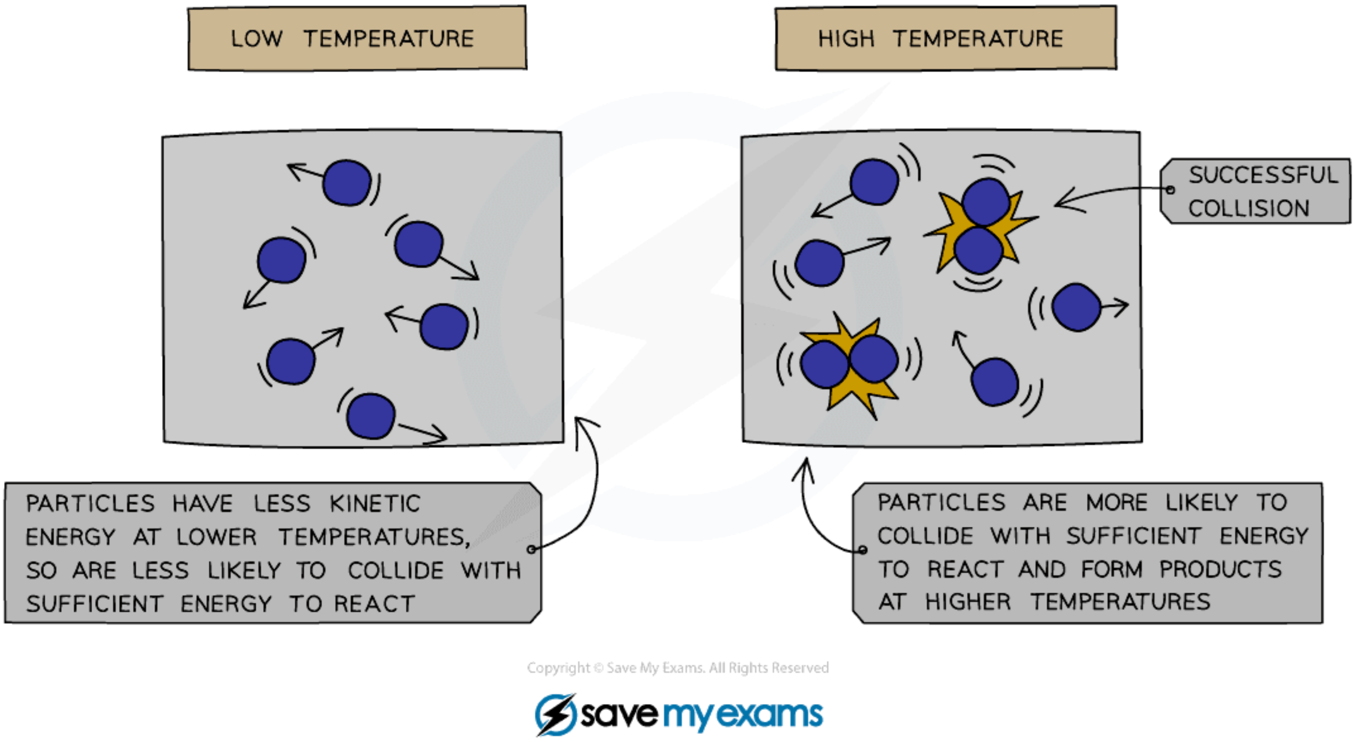

Diagram showing the increased kinetic energy that particles have at higher temperatures

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Concentration

- Increasing the concentration means there are more particles per cm^3, so there is less space between the particles
- Since there are more particles then it follows that there are more collisions, hence the rate of reaction increases

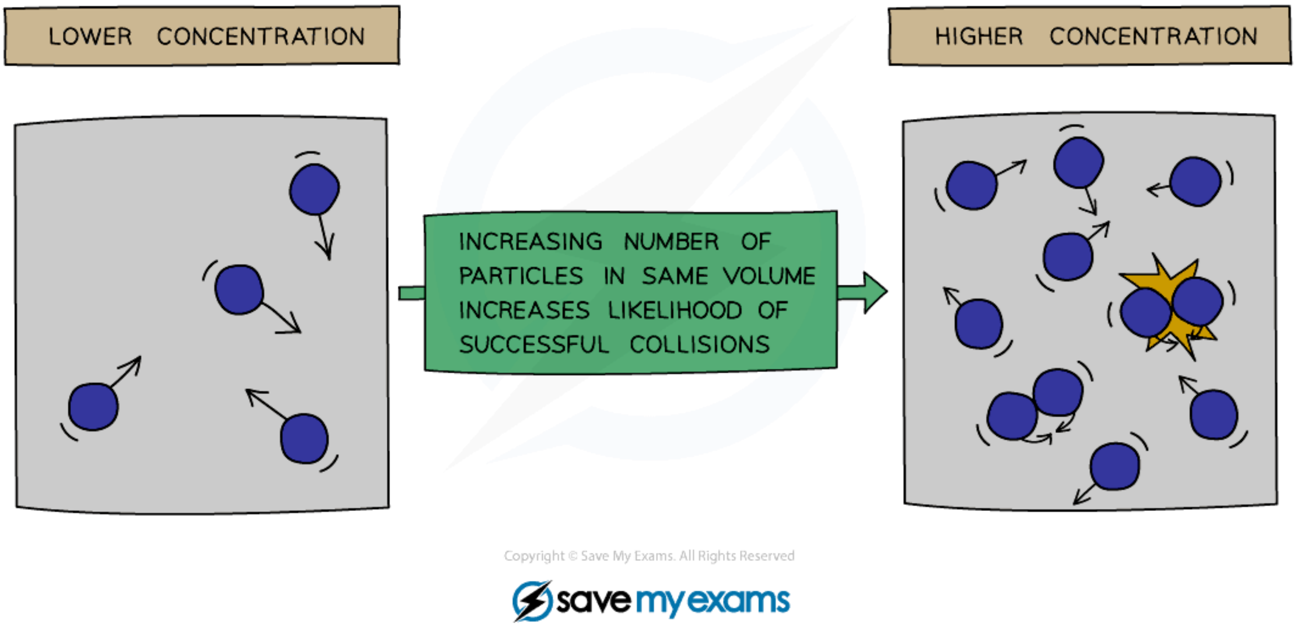


Diagram showing the decreases in space between particles at higher concentrations

Exam Tip

When answering questions on the effect of concentration on the rate of reaction, you should mention that there are more particles **per unit volume** (usually cm^3) and this causes an increase in the rate of collisions.

Extended Only

Photochemistry

Photochemical reactions

- These reactions occur only when light is present
- The greater the intensity of ultraviolet light then the greater the rate of reaction
- Eg the substitution of hydrogen atoms in methane by chlorine:

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

Silver salts in photography

- Black and white photography film surfaces contain crystals of silver bromide
- When exposed to light they decompose to silver:

$$2AgBr \rightarrow 2Ag + Br_2$$

- AgBr is colourless at low concentrations but the Ag appears grey-black
- Parts of the film appear black, grey or white depending on the exposure:
    - Stronger light = black or dark grey
    - Weaker light = light grey
    - Not exposed = white

Photosynthesis

- This is the process in which plants produce food for reproduction and growth
- The equation is:

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

- The process requires sunlight and chlorophyll
- Chlorophyll is the green pigment in plants which absorbs sunlight and acts as the catalyst for photosynthesis

7.2 Reversible Reactions

Reversible Reactions

Reversible reactions

- Some reactions go to **completion**, where the reactants are used up to form the product molecules and the reaction stops when all of the reactants are used up
- In reversible reactions, the product molecules can themselves react with each other or decompose and form the reactant molecules again
- It is said that the reaction can occur in **both directions**: the forward reaction (which forms the products) and the reverse direction (which forms the reactants)

Chemical equations for reversible reactions

- When writing chemical equations for reversible reactions, two arrows are used to indicate the forward and reverse reactions
- Each one is drawn with just half an arrowhead – the top one points to the right, and the bottom one points to the left

Example

- The reaction for the **Haber Process** which is the production of ammonia from hydrogen and nitrogen:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

Hydrated and anhydrous salts

- Hydrated salts are salts that contain **water of crystallisation** which affects their molecular shape and colour
- Water of crystallisation is the water that is stoichiometrically included in the structure of some salts during the crystallisation process
- A common example is copper(II) sulfate which crystallises forming the salt copper(II) sulfate pentahydrate, $CuSO_4\mathbf{.}5H_2O$
- Water of crystallisation is indicated with a dot written in between the salt molecule and the surrounding water molecules
- Anhydrous salts are those that have lost their water of crystallisation, usually by heating, in which the salt becomes **dehydrated**

Dehydration of Hydrated Copper (II) Sulfate:

Hydrated Copper (II) Sulfate ⇌ Anhydrous Copper (II) Sulfate + Water

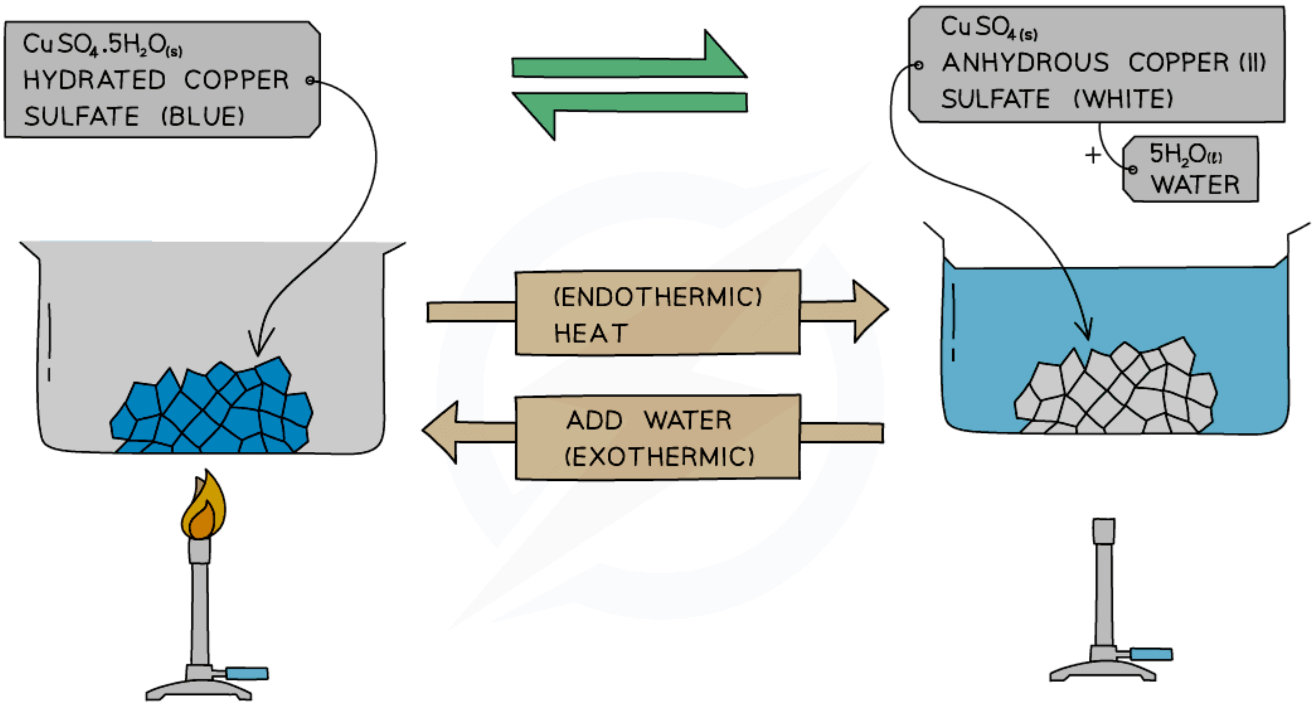


Diagram showing the dehydration of Hydrated Copper (II) Sulfate

Explanation:

- When anhydrous copper (II) sulfate crystals are added to water they turn **blue** and heat is given off (exothermic); this reaction is reversible
- When Copper (II) Sulfate crystals are heated in a test tube, the blue crystals turn into a **white** powder and a clear, colourless liquid (water) collects at the top of the test tube
- The form of Copper (II) Sulfate in the crystals is known as Hydrated Copper (II) Sulfate because it contains water of crystallisation
- When Hydrated Copper (II) Sulfate is heated, it loses its water of crystallisation and turns into anhydrous Copper (II) Sulfate:

$$CuSO_4.5H_2O\ (s) \rightleftharpoons CuSO_4\ (s) + 5H_2O\ (l)$$

Dehydration of Hydrated Cobalt (II) Chloride:

Hydrated Cobalt (II) Chloride ⇌ Anhydrous Cobalt (II) Chloride + Water

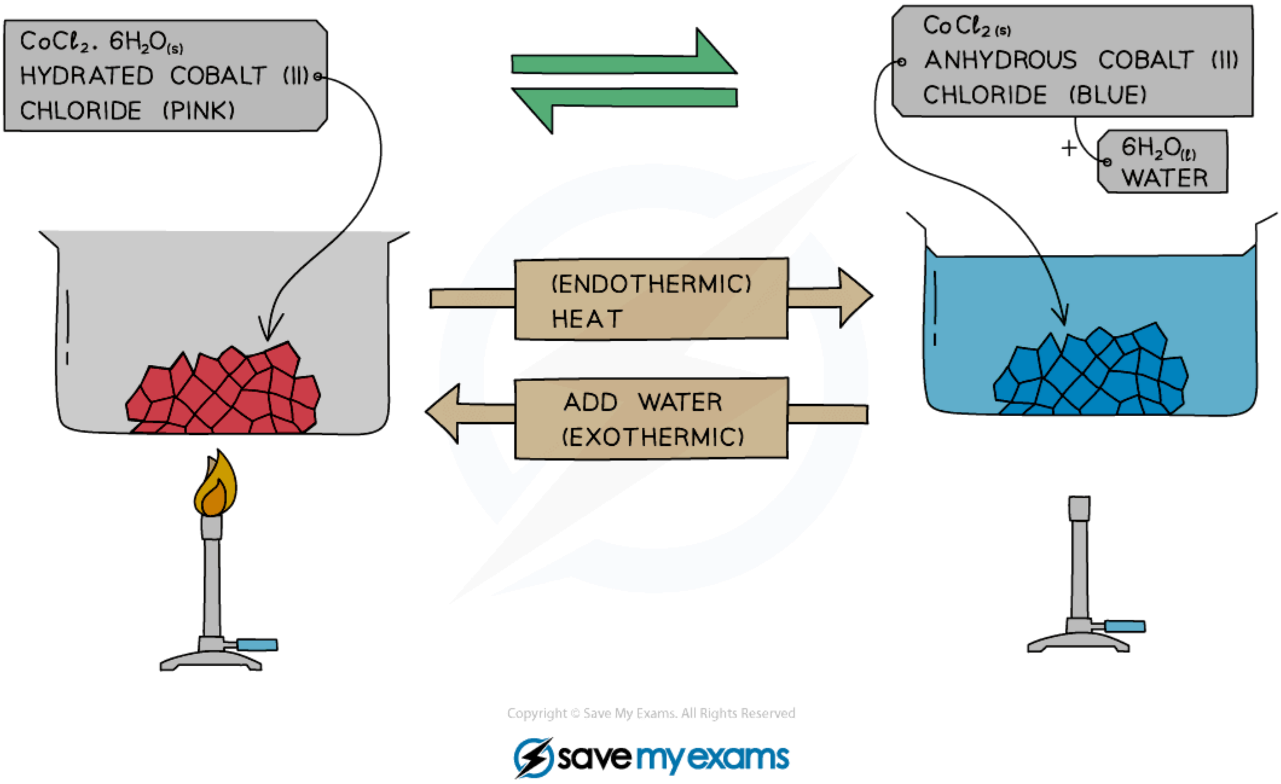


Diagram showing the dehydration of Hydrated Cobalt (II) Chloride

Hydration of Cobalt(II) Chloride

- When anhydrous **blue** cobalt(II) chloride crystals are added to water they turn **pink** and the reaction is **reversible**
- When the cobalt(II) chloride crystals are heated in a test tube, the **pink** crystals turn back to the **blue** colour again as the water of crystallisation is lost
- The form of cobalt(II) chloride in the crystals that are pink is known as hydrated cobalt (II) chloride because it contains water of crystallisation
- When hydrated cobalt(II) chloride is heated, it loses its water of crystallisation and turns into anhydrous cobalt(II) chloride:

$$CoCl_2.6H_2O\ (s) \rightleftharpoons CoCl_2\ (s) + 6H_2O\ (l)$$

Exam Tip

Both the hydration of $CoCl_2$ and $CuSO_4$ are chemical tests which are commonly used to detect the presence of water. You should remember the equations and colour changes:

- $CoCl_2 + 6H_2O \rightleftharpoons CoCl_2.6H_2O$ **Blue to pink**
- $CuSO_4 + 5H_2O \rightleftharpoons CuSO_4.5H_2O$ **White to blue**

Extended Only

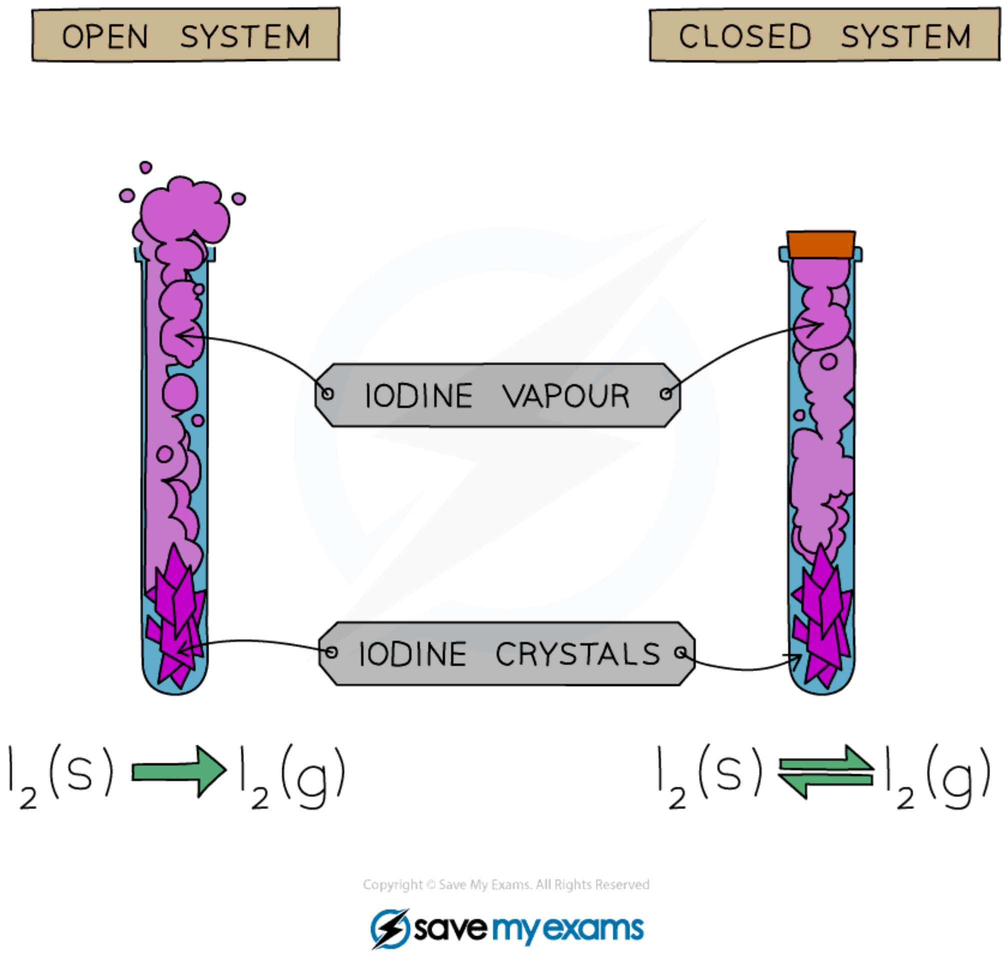
The Concept of Equilibrium

Reversible reactions and equilibrium

- We have already seen that a reversible reaction is one that occurs in **both** directions
- When during the course of reaction, the rate of the forward reaction equals the rate of the reverse reaction, then the overall reaction is said to be in a state of **equilibrium**

Characteristics of a reaction at equilibrium

- It is **dynamic** eg the molecules on the left and right of the equation are **changing** into each other by chemical reactions constantly and at the same rate
- The concentration of reactants and products remains **constant** (given there is no other change to the system such as temperature and pressure)
- It only occurs in a **closed system** so that none of the participating chemical species are able to leave the reaction vessel



Equilibrium can only be reached in a closed vessel which prevents reactants or products from escaping system

The reaction between H_2 and N_2 in the Haber process

- When only nitrogen and hydrogen are present at the beginning of the reaction, the rate of the forward reaction is at its **highest**, since the **concentrations** of hydrogen and nitrogen are at their **highest**
- As the reaction proceeds, the concentrations of hydrogen and nitrogen gradually **decrease**, so the rate of the forward reaction will decrease
- However, the concentration of ammonia is gradually increasing and so the rate of the **backward** reaction will increase (ammonia will decompose to reform hydrogen and nitrogen)
- Since the two reactions are interlinked and none of the gas can escape, the rate of the forward reaction and the rate of the backward reaction will eventually become **equal** and equilibrium is reached:

$$3H_2\ (g) + N_2\ (g) \rightleftharpoons 2NH_3\ (g)$$

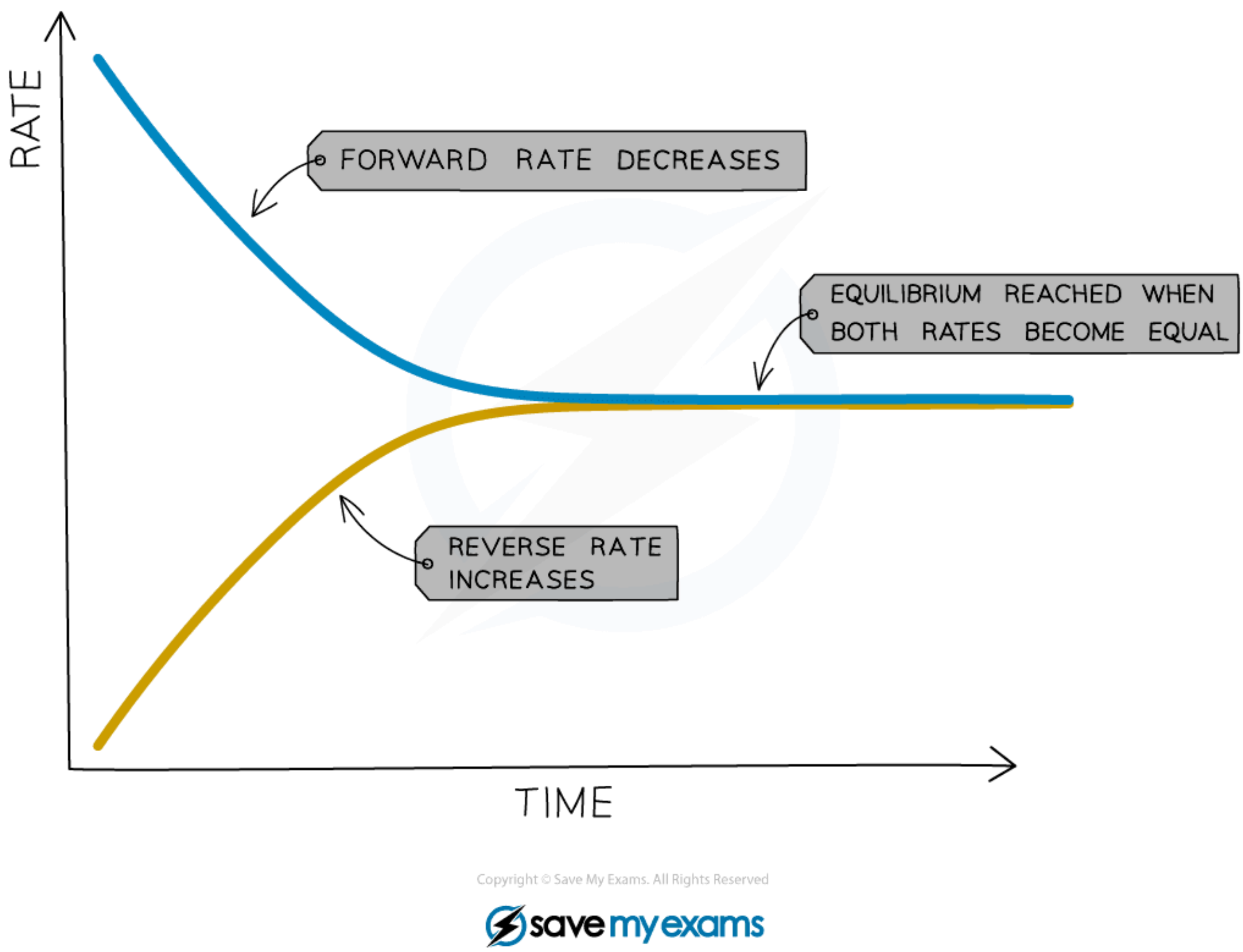


Diagram showing when the rates of forward and backward reactions become equal

The position of equilibrium

- Equilibrium position refers to the relationship between the concentration of reactants and products at the equilibrium state
- When the position of equilibrium shifts to the **left**, it means the concentration of **reactant** increases
- When the position of equilibrium shifts to **right**, this means the concentration of **product** increases

Effect of catalyst on equilibrium position

- The presence of a catalyst does **not** affect the position of equilibrium but it does increase the rate at which equilibrium is reached
- This is because the catalyst increases the rate of **both** the forward and backward reactions by the same amount (by providing an alternative pathway requiring lower activation energy)
- As a result, the **concentration** of reactants and products is nevertheless the **same** at equilibrium as it would be without the catalyst

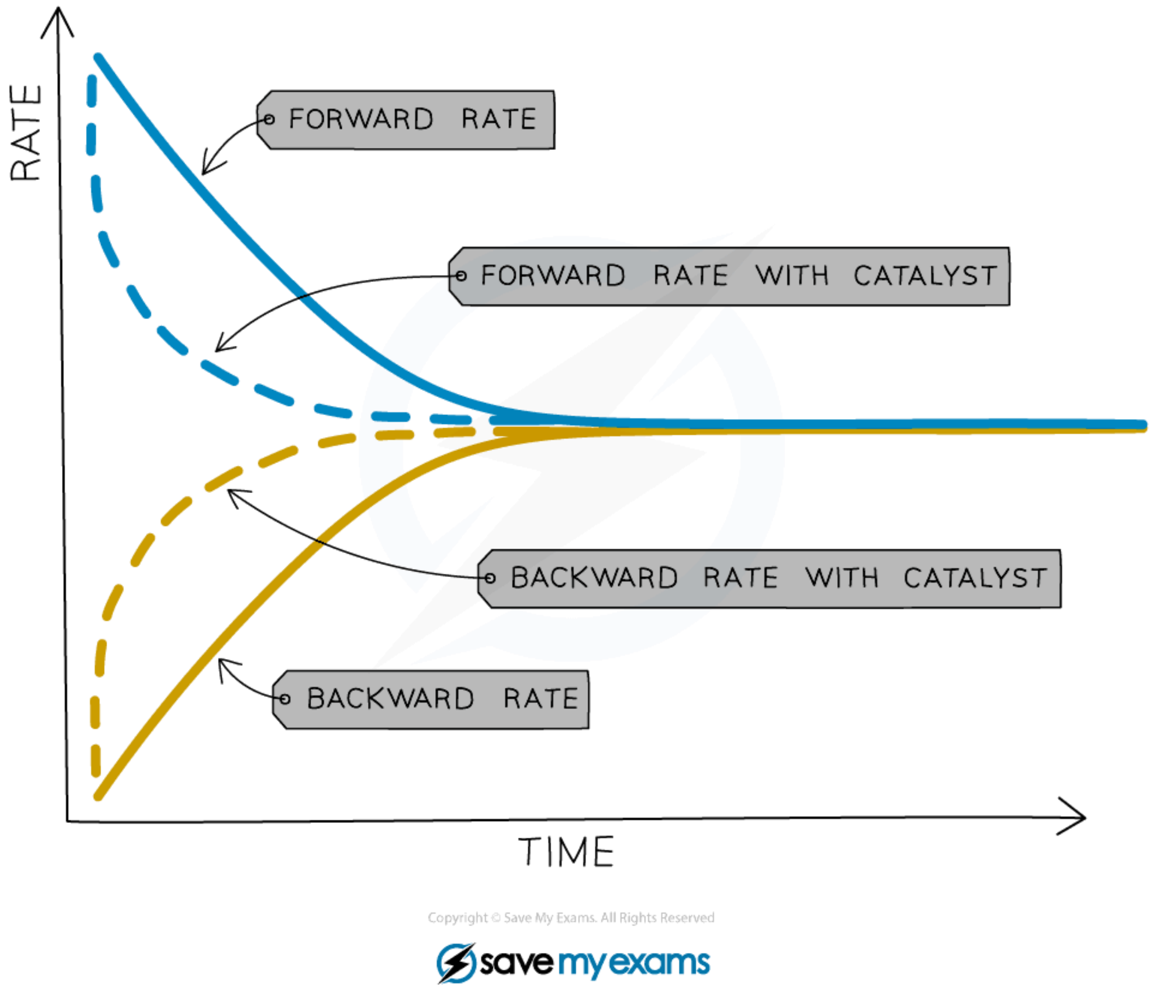


Diagram showing the effect of catalyst on equilibrium position

Extended Only

Le Chatelier's Principle

- **Le Chatelier's Principle** states that when a change is made to the conditions of a system at equilibrium, the system automatically moves to **oppose** the change.
- The principle is used to predict changes to the position of equilibrium when there are changes in temperature, pressure or concentration.

Effects of temperature

CHANGE	HOW THE EQUILIBRIUM SHIFTS
INCREASE IN TEMPERATURE	EQUILIBRIUM MOVES IN THE **ENDOTHERMIC** DIRECTION TO REVERSE THE CHANGE
DECREASE IN TEMPERATURE	EQUILIBRIUM MOVES IN THE **EXOTHERMIC** DIRECTION TO REVERSE THE CHANGE

-

Example:

Iodine Monochloride reacts reversibly with Chlorine to form Iodine Trichloride

$ICl \quad + \quad Cl_2 \quad \rightleftharpoons \quad ICl_3$

Dark Brown Yellow

When the equilibrium mixture is heated, it becomes **dark brown** in colour. Explain whether the backward reaction is exothermic or endothermic:

- Equilibrium has shifted to the left as the colour dark brown means that more of ICI is produced
- Increasing temperature moves the equilibrium in the endothermic direction
- **So the backward reaction is endothermic**

Effects of pressure

CHANGE	HOW THE EQUILIBRIUM SHIFTS
INCREASE IN PRESSURE	EQUILIBRIUM SHIFTS IN THE DIRECTION THAT PRODUCES THE **SMALLER** NUMBER OF MOLECULES OF GAS TO DECREASE THE PRESSURE AGAIN
DECREASE IN PRESSURE	EQUILIBRIUM SHIFTS IN THE DIRECTION THAT PRODUCES THE **LARGER** NUMBER OF MOLECULES OF GAS TO INCREASE THE PRESSURE AGAIN

Nitrogen Dioxide can form Dinitrogen Tetroxide, a colourless gas

$2NO_2 \rightleftharpoons N_2O_4$

Brown Gas Colourless Gas

Predict the effect of an increase in pressure on the position of equilibrium:

-
    - Number of molecules of gas on the left = 2
    - Number of molecules of gas on the right = 1

- An increase in pressure will cause equilibrium to shift in the direction that produces the smaller number of molecules of gas
- **So equilibrium shifts to the right**

-

Effects of concentration

CHANGE	HOW THE EQUILIBRIUM SHIFTS
INCREASE IN CONCENTRATION	EQUILIBRIUM SHIFTS TO THE **RIGHT** TO REDUCE THE EFFECT OF INCREASE IN THE CONCENTRATION OF A REACTANT
DECREASE IN CONCENTRATION	EQUILIBRIUM SHIFTS TO THE **LEFT** TO REDUCE THE EFFECT OF A DECREASE IN REACTANT (OR AN INCREASE IN THE CONCENTRATION OF PRODUCT)

Example:

Iodine Monochloride reacts reversibly with Chlorine to form Iodine Trichloride

$$ICl + Cl_2 \rightleftharpoons ICl_3$$

Dark Brown Yellow

Predict the effect of an increase in concentration on the position of equilibrium:

- An increase in the concentration of ICl or Cl_2 causes the equilibrium to shift to the **right** so more of the **yellow** product is formed
- A decrease in the concentration of ICl or Cl_2 causes the equilibrium to shift to the **left** so more of the **dark brown** reactant is formed

Exam Tip

When the conditions at equilibrium are changed, the system always responds by doing the **opposite**.

For example if the concentration is increased the system tries to reduce it by changing the direction of the reaction or if the temperature is increased the system will try to reduce the temperature by absorbing the extra heat.

7.3 Redox Reactions

Oxidation & Reduction

- Oxidation and reduction take **place together** at the **same time** in the **same reaction**
- These are called **redox** reactions
- There are **three** definitions of **oxidation**. It is a reaction in which:
  - **Oxygen** is **added** to an element or a compound
  - An element, ion or compound **loses electrons**
  - The **oxidation state** of an element is **increased**
- There are **three** definitions of **reduction**. It is a reaction in which:
  - **Oxygen** is **removed** from an element or a compound
  - An element, ion or compound **gains electrons**
  - The **oxidation state** of an element is **decreased**

Oxidation state

- The oxidation state (also called oxidation number) is a number assigned to an atom or ion in a compound which indicates the **degree of oxidation** (or reduction)
- The oxidation state helps you to keep track of the movement of electrons in a redox process
- It is written as a +/- sign followed by a number.
- Eg O^{2-} means that it is an atom of oxygen that has an oxidation state of **-2**. It is not written as O^{2-} as this refers to the ion and its charge

Assigning the oxidation number

- Oxidation number refers to a **single atom** or **ion** only
- The oxidation number of a **compound** is **0** and of an **element** (for example Br in Br_2) is also **0**
- The oxidation number of oxygen in a compound is always **-2** (except in peroxide R-O-O-R, where it is **-1)**
- For example in FeO, oxygen is -2 then Fe must have an oxidation number of **+2** as the overall oxidation number for the **compound** must be **0**

Ionic Equations

- Ionic equations are used to show only the particles that actually take part in a reaction
- These equations show only the ions that change their status during a chemical process, i.e: their bonding or physical state changes
- The other ions present are not involved and are called spectator ions

Writing ionic equations

- For the neutralisation reaction between hydrochloric acid and sodium hydroxide:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

- If we write out all of the ions present in the equation and include the state symbols, we get:

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

- The **spectator ions** are thus Na^+ and Cl^-. Removing these from the previous equation leaves the overall net ionic equation:

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

- This ionic equation is the **same for all acid-base neutralisation** reactions

Example redox equation: oxygen loss/gain

Zinc oxide + carbon → zinc + carbon monoxide

$$ZnO + C \rightarrow Zn + CO$$

- In this reaction the zinc oxide has been reduced since it has **lost** The carbon atom has been oxidised since it has gained oxygen

Extended Only

Redox & Electron Transfer

Example redox equation: electron loss/gain and oxidation state

Zinc + copper sulphate → zinc sulphate + copper

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- Writing this as an ionic equation:

$$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$$

- By analysing the ionic equation, it becomes clear that zinc has become oxidised as its oxidation state has **increased** and it has **lost** electrons:

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

- Copper has been reduced as its oxidation state has **decreased** and it has **gained** electrons:

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

Exam Tip

Use the mnemonic **OIL-RIG** to remember oxidation and reduction in terms of the movement of electrons: **O**xidation **I**s **L**oss – **R**eduction **I**s **G**ain.

Extended Only

Oxidising & Reducing Agents

Oxidising agent

- A substance that **oxidises** another substance, in so doing becoming itself **reduced**
- Common examples include hydrogen peroxide, fluorine and chlorine

Reducing agent

- A substance that **reduces** another substance, in so doing becoming itself **oxidised**
- Common examples include carbon and hydrogen
- The process of reduction is very important in the chemical industry as a means of extracting metals from their ores

Example

$$CuO + H_2 \rightarrow Cu + H_2O$$

- In the above reaction, hydrogen is reducing the CuO and is itself oxidised, so the **reducing agent** is therefore **hydrogen**
- The CuO is reduced to Cu and has oxidised the hydrogen, so the **oxidising agent** is therefore **copper oxide**

Extended Only

Redox Reactions

Identifying redox reactions

- Redox reactions can be identified by the changes in the **oxidation states** when a reactant goes to a product

Example

Chlorine + potassium iodide → potassium chloride + iodine

$$Cl_2 + 2KI \rightarrow 2KCl + I_2$$

- Chlorine has become reduced as its **oxidation state** has **decreased** from 0 to -1 on changing from the chlorine molecule to chloride ions:

$$Cl_2(g) \rightarrow 2Cl^-(aq)$$

- Iodine has been oxidised as its **oxidation state** has **increased** from -1 to 0 on changing from iodide ions to the iodine molecule:

$$2I^-(aq) \rightarrow I_2(s)$$

Identifying redox reactions by colour changes

- The tests for redox reactions involve the observation of a colour change in the solution being analyse
- Two common examples are acidified potassium manganate(VII), and potassium iodide
- Potassium manganate (VII), $KMnO_4$, is an oxidising agent which is often used to test for the presence of reducing agents
- When acidified potassium manganate (VII) is added to a reducing agent its colour changes from pink-purple to colourless
- Potassium iodide, KI, is a reducing agent which is often used to test for the presence of oxidising agents
- When added to an acidified solution of an oxidising agent such as aqueous chlorine or hydrogen peroxide, the solution turns a brown colour due to the formation of iodine

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