PCFGSTUDY

AQA GCSE Combined Science: Trilogy — Chemistry

Topic 13

The rate and extent of chemical change

Specification 8464 · AQA

Free edition

Revision notes for one topic, free to download.

These are the summary notes only. The full edition of this booklet also carries the worked examples, the required practicals and the formula sheets, with every topic in one place.

Colour Codes

Definitions

Diagrams

Notes and warnings

Sub-topic headings

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Define Rate of reaction

- how quickly a reactant is used up, or how quickly a product is formed.

Define Instantaneous rate

- the rate at one moment, found from the gradient of a tangent to the curve at that time.

Define Activation energy

- the minimum amount of energy that particles must have in order to react.

Define Catalyst

- a substance that changes (speeds up) the rate of a reaction but is not used up during the reaction.

Define Turbidity

- cloudiness of a solution caused by a precipitate forming in it.

Define Equilibrium

- the state reached in a closed system when the forward and reverse reactions occur at exactly the same rate.

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Define HT ONLY Equilibrium position

- the relative amounts of reactants and products present at equilibrium.

Define Mean rate of reaction

- the quantity of reactant used, or product formed, divided by the time taken.

Define Collision theory

- the idea that chemical reactions can occur only when reacting particles collide with each other **and** with sufficient energy.

Define Surface area to volume ratio

- surface area divided by volume; it increases as the pieces of solid get smaller.

Define Enzyme

- a biological catalyst; enzymes act as catalysts in living systems.

Define Reversible reaction

- a reaction in which the products can react to produce the original reactants, shown by the symbol ⇌.

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Define HT ONLY Le Chatelier's Principle

- if a system is at equilibrium and a change is made to any of the conditions, the system responds to counteract the change.

13.1 Rate of reaction

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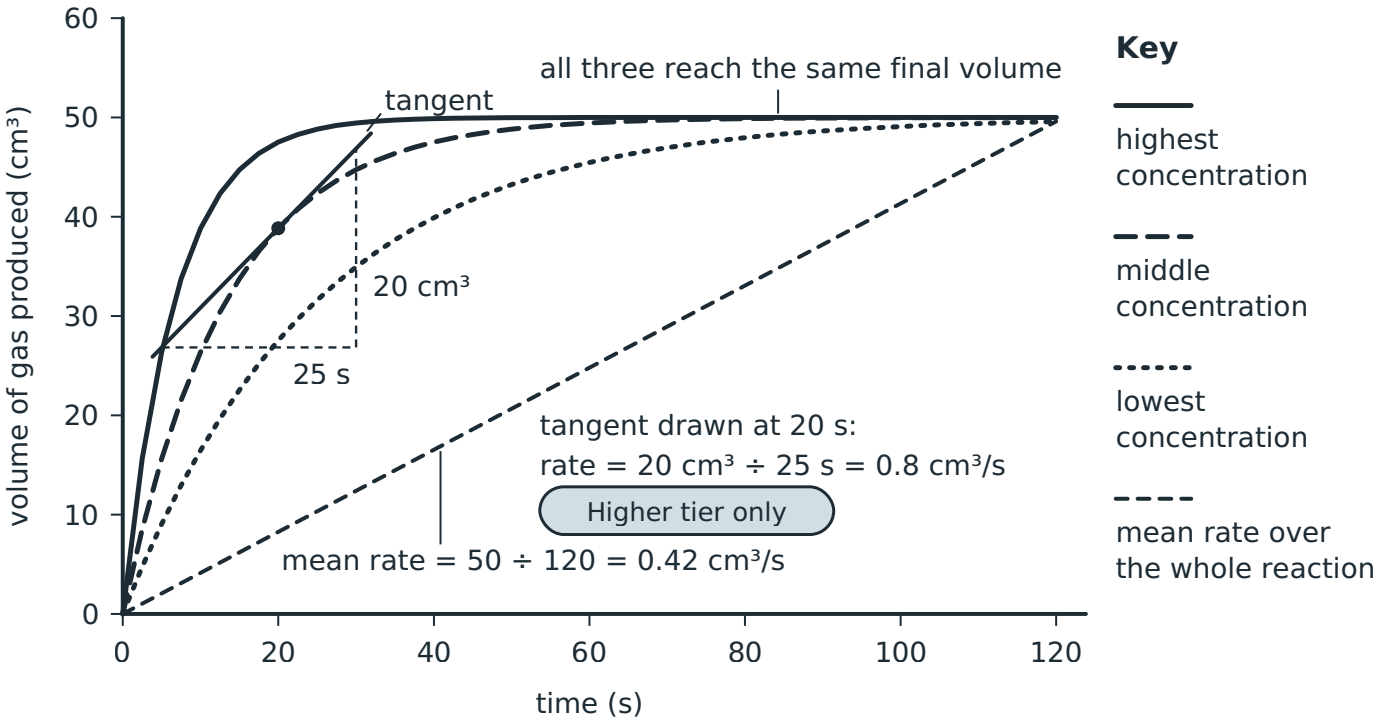
13.1.1 Calculating rates of reactions

The rate of a reaction is found by measuring **the quantity of a reactant used** or **the quantity of product formed** over time. The quantity is measured as a **mass in grams** or a **volume in cm³**, so rate is given in **g/s** or **cm³/s**. HT ONLY You must also use quantities in **moles**, giving a rate in **mol/s**.

Finding the rate from a graph using a tangent



The steeper the curve, the faster the reaction. The highest concentration gives the steepest gradient at the start, but all three produce the same final volume of gas because the same amount of the limiting reactant is used up. The gradient of a tangent gives the rate at that moment.

Rate graphs. Plot the quantity of product formed (or reactant used up) on the *y*-axis against time on the *x*-axis and draw a smooth curve of best fit.

- The **steeper** the curve, the **faster** the reaction. It is **steepest at the start**, when the concentration of reactants is highest, and gets shallower as they are used up.
- The curve becomes **horizontal** when the reaction has finished and the limiting reactant is used up — a flat line means the rate is **zero**.

To find the rate **at a particular moment**, draw a **tangent** to the curve at that time and find its slope.

13.1.2 Factors which affect the rates of chemical reactions

The factors which affect the rates of chemical reactions include the **concentration** of reactants in solution, the **pressure** of reacting gases, the **surface area** of solid reactants, the **temperature**, and the presence of **catalysts**. Increasing any of the first four increases the rate; a catalyst increases the rate without being used up.

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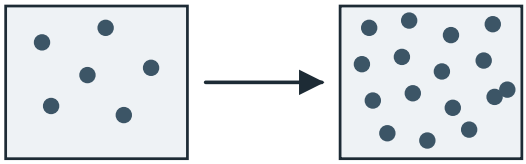
13.1.3 Collision theory and activation energy

Collision theory explains how the various factors affect rates of reaction. According to this theory, chemical reactions can occur only when reacting particles collide with each other and with sufficient energy. The minimum amount of energy that particles must have to react is called the **activation energy**.

Collision theory and activation energy

Concentration and pressure



dilute concentrated

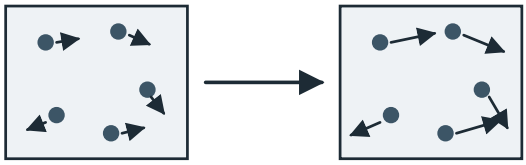
More particles in the same volume, so they collide more often. Increasing the pressure of a gas does the same thing: more frequent collisions, faster rate.

Surface area of a solid

one large lump the same mass, in pieces

Smaller pieces have a larger surface area to volume ratio, so more particles are exposed. The frequency of collisions increases and so the rate increases.

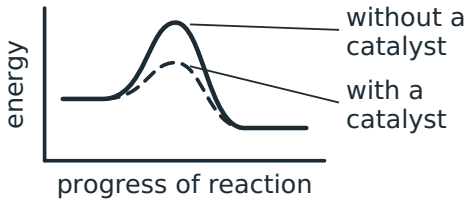
Temperature



cooler hotter

Particles move faster, so collisions are more frequent and more energetic. A greater proportion of collisions have energy above the activation energy.

Catalyst



A catalyst provides a different pathway with a lower activation energy, so more collisions have enough energy to react. The catalyst is not used up.

Collision theory: particles react only when they collide with at least the activation energy.

ANALOGY

A reaction is like a game of dodgeball: for a "hit" the ball must **reach** someone (a collision) **and** be thrown **hard enough** to count (the activation energy). Crowding more players in gives more hits per second; throwing harder makes more throws count. A gentle throw that lands is wasted — just like a collision with less than the activation energy.

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Factor	What you change	Collision-theory explanation
Concentration (solutions)	Increase it	More particles in the same volume, so collisions are **more frequent**
Pressure (gases)	Increase it	The same number of gas particles in a smaller volume, so collisions are **more frequent**
Surface area (solids)	Break the solid up / powder it	Larger **surface area to volume ratio**, so more particles are exposed, so collisions are **more frequent**
Temperature	Increase it	Particles move faster, so collisions are **more frequent and more energetic**, so a greater proportion have at least the activation energy — two reasons, so a large effect
Catalyst	Add one	Provides a different pathway with a lower activation energy, so more collisions are successful; the frequency of collisions is unchanged

Simple proportionality. Doubling the concentration roughly **doubles** the frequency of collisions and so roughly doubles the rate; cutting a 1 cm cube into eight 0.5 cm cubes doubles the surface area for the same volume.

13.1.4 Catalysts

- Catalysts change the rate of chemical reactions but are not used up during the reaction — the mass of catalyst is the same at the end as at the start.
- **Different reactions need different catalysts. Enzymes act as catalysts in biological systems.**
- Catalysts increase the rate by providing a different pathway for the reaction that has a lower activation energy.
- **Identify a catalyst** by its effect on the rate and because it does not appear in the chemical equation (it is usually written above the arrow). You do not need to know the names of catalysts other than those named elsewhere in the course.

Reaction profile with and without a catalyst — energy on the *y*-axis, progress of reaction on the *x*-axis.

- Both curves start at the **same reactants level** and end at the **same products level** — the catalyst does **not** change the overall energy change.
- The uncatalysed curve rises to a **high hump**; the height from the reactants line to the peak is the **activation energy**.
- The catalysed curve is a **second, lower hump** between the same two levels, with a **smaller activation energy**. Label both activation energies and the overall energy change.

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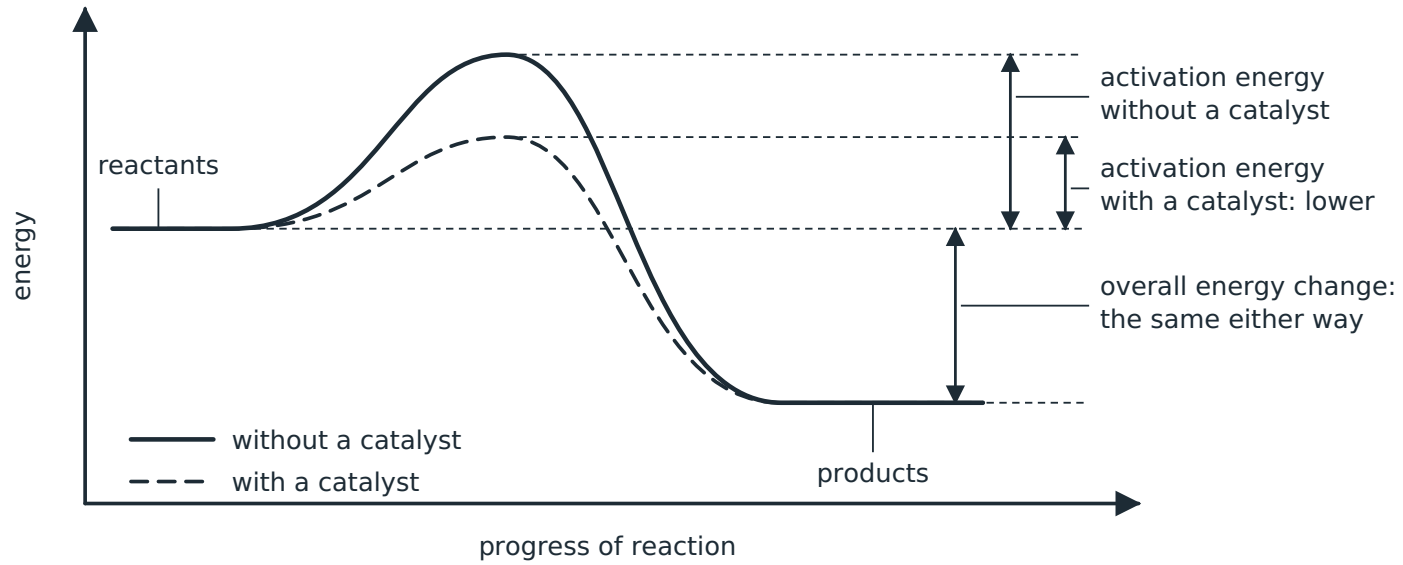
How a catalyst changes the reaction profile



A catalyst provides a different pathway for the reaction that has a lower activation energy. The reactants and the products are unchanged, so the overall energy change is the same.

ANALOGY

A catalyst is a **tunnel through a mountain** instead of the pass over the top: same start valley, same end valley, but far less climbing, so many more travellers make it. The tunnel is not worn away by the traffic — which is why the catalyst is not used up.

WATCH OUT

A catalyst **is not used up** and does not change the position of equilibrium — it just gets you there faster. Common errors to avoid:

- "It is used up" — no; it can be filtered off, dried and reused, and its mass is unchanged.
- "It lowers the activation energy of the reaction" — more precisely, it provides an alternative pathway with a lower activation energy.
- "It gives a bigger yield at equilibrium" — no. It speeds up the forward **and** reverse reactions equally, so equilibrium is reached **sooner** with exactly the same amounts present.

13.2 Reversible reactions and dynamic equilibrium

13.2.1 Reversible reactions

- In some reactions, the products of the reaction can react to produce the original reactants. These are **reversible reactions**, written A + B ⇌ C + D.
- The direction of a reversible reaction can be changed by changing the conditions.
- Examples: hydrated copper(II) sulfate ⇌ anhydrous copper(II) sulfate + water (blue crystals ⇌ white powder), and ammonium chloride ⇌ ammonia + hydrogen chloride.

13.2.2 Energy changes and reversible reactions

- If a reversible reaction is exothermic in one direction, it is endothermic in the opposite direction, and the same amount of energy is transferred in each case.
- Example: hydrated → anhydrous copper(II) sulfate + water is **endothermic** (you must heat it); adding water back is **exothermic** by exactly the same amount.

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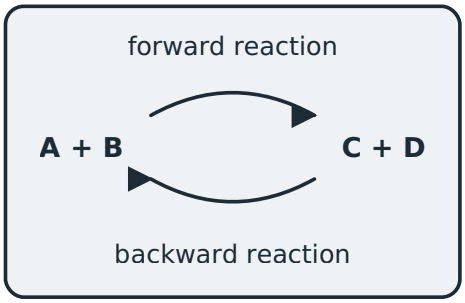
13.2.3 Equilibrium

- When a reversible reaction occurs in apparatus which prevents the escape of reactants and products (a **closed system**), equilibrium is reached when the forward and reverse reactions occur at exactly the same rate.
- Equilibrium is **dynamic**: both reactions are still happening, but at the same rate, so there is **no overall change** in the amounts of reactants and products.
- Those amounts are usually **not equal** — equilibrium means equal **rates**, not equal amounts.

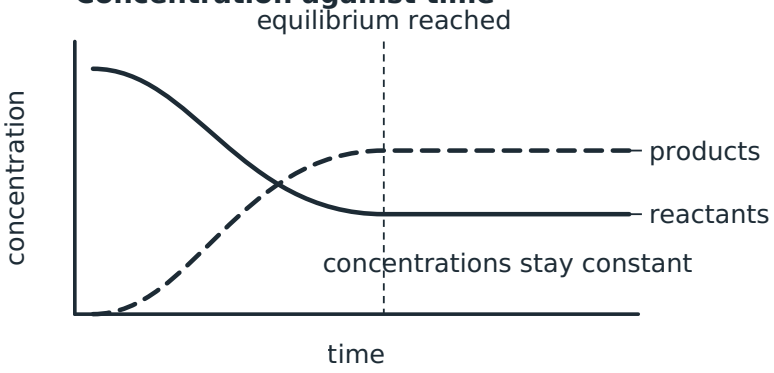
Dynamic equilibrium

Dynamic equilibrium



closed system: nothing can escape

Concentration against time



At equilibrium the forward and backward reactions are still happening, but at exactly the same rate, so the concentrations of the reactants and the products stop changing. They are not necessarily equal.

Le Chatelier's principle

Higher tier only

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

shift right: more products

reactants ⇌ **products**

shift left: more reactants

Concentration

Increase the concentration of a reactant: the position of equilibrium shifts to the right, so more products form. Decreasing a product also shifts it to the right.

Temperature

Increase the temperature: the equilibrium shifts in the endothermic direction. Decrease the temperature: it shifts in the exothermic direction.

Pressure (gases)

Increase the pressure: the equilibrium shifts to the side with fewer molecules of gas. Decrease it and it shifts to the side with more molecules of gas.

ANALOGY

In a shopping centre the up and down escalators each carry 30 people a minute. Everyone is moving, but the number of shoppers on each floor never changes. That is **dynamic** equilibrium: constant activity both ways, no overall change.

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13.2.4 HT ONLY The effect of changing conditions on equilibrium

- HT ONLY The relative amounts of all the reactants and products at equilibrium **depend on the conditions** of the reaction.
- HT ONLY If a system is at equilibrium and a change is made to any of the conditions, the system responds to counteract the change.
- HT ONLY These effects can be predicted using **Le Chatelier's Principle**. You must make **qualitative predictions** when given appropriate information (the balanced equation, and whether the forward reaction is exothermic or endothermic).

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13.2.6 HT ONLY The effect of temperature changes on equilibrium

HT ONLY If the temperature of a system at equilibrium is **increased**:

- the relative amount of **products increases** for an **endothermic** reaction
- the relative amount of **products decreases** for an **exothermic** reaction.

HT ONLY If the temperature of a system at equilibrium is **decreased**:

- the relative amount of products decreases for an endothermic reaction
- the relative amount of products increases for an exothermic reaction.
- HT ONLY In one line: raising the temperature favours the endothermic direction, lowering it favours the exothermic direction, because the system counteracts the change.
- HT ONLY Interpret given data to predict the effect of a temperature change on a given equilibrium.

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ANALOGY

HT ONLY Le Chatelier's Principle is the chemistry of a stubborn toddler: whatever you do, it does the opposite. Add reactant and it makes product; heat it and it absorbs the heat by going the endothermic way; squeeze it and it shifts to the side with fewer gas molecules. "The system opposes the change you made" is the exam sentence.

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13.2.5 HT ONLY The effect of changing concentration

- HT ONLY If the concentration of one of the reactants or products is changed, the system is no longer at equilibrium and the concentrations of all the substances will change until equilibrium is reached again.
- HT ONLY If the concentration of a **reactant is increased**, **more products will be formed** until equilibrium is reached again.
- HT ONLY If the concentration of a **product is decreased**, **more reactants will react** until equilibrium is reached again.
- HT ONLY Interpret **given data** to predict the effect of a concentration change on a given reaction at equilibrium.

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13.2.7 HT ONLY The effect of pressure changes on equilibrium

HT ONLY For **gaseous** reactions at equilibrium:

- an **increase in pressure** causes the equilibrium position to shift towards the side with the **smaller number of molecules**, as shown by the symbol equation for that reaction
- a **decrease in pressure** causes the equilibrium position to shift towards the side with the **larger number of molecules**, as shown by the symbol equation for that reaction.
- HT ONLY Count gas molecules using the **balancing numbers**. In $\mathbf{N_2 + 3H_2 \rightleftharpoons 2NH_3}$ there are 4 molecules on the left and 2 on the right, so increasing the pressure increases the yield of ammonia.
- HT ONLY If both sides have the **same number of gas molecules**, pressure has **no effect** on the equilibrium position.
- HT ONLY Interpret given data to predict the effect of pressure changes on given equilibria.

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