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GCSE CHEMISTRY || LESSON NOTE

Rates of Reaction

Collision theory, the factors that change the rate, the two core practicals, reading rate graphs, catalysts and enzymes.

Specification	Points 7.1 to 7.8 (including Core Practical 7.1)
Maths skills	Rates from data, surface area to volume ratio, gradients and tangents
Inside	5 concepts || 31 worked examples || 29 practice questions || full worked answer key

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Before we start

THE BIG IDEA

Some reactions are very fast (an explosion), some are very slow (rusting). The **rate of reaction** tells us how quickly reactants are used up or products are made. Particles can only react when they **collide** with enough energy, so anything that makes successful collisions happen more often makes the reaction faster.

Learning objectives

Spec	I can...
7.1	carry out the core practical: measure the gas made by marble chips and hydrochloric acid, and time the colour change of sodium thiosulfate and hydrochloric acid
7.2	suggest practical methods for measuring the rate of a given reaction
7.3	explain how reactions happen when particles collide, and why more frequent or more energetic collisions increase the rate
7.4	explain the effects of temperature, concentration, surface area to volume ratio and pressure on rate
7.5	interpret graphs of mass, volume or concentration against time
7.6	describe a catalyst
7.7	explain how a catalyst increases rate in terms of activation energy
7.8	recall that enzymes are biological catalysts used to make alcoholic drinks

Key words

Word	Meaning
Rate of reaction	How quickly a reactant is used up or a product is formed.
Successful collision	A collision between particles that has enough energy to cause a reaction.
Activation energy	The minimum energy particles need when they collide for a reaction to happen.
Frequency of collisions	How many collisions happen every second.
Surface area to volume ratio	Surface area divided by volume. Smaller pieces have a bigger ratio.
Catalyst	A substance that speeds up a reaction without being changed chemically or in mass at the end.
Enzyme	A biological catalyst, made of protein, found in living things.
Tangent	A straight line that just touches a curve at one point. Its gradient gives the rate at that moment.

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CONCEPT 1

Collision Theory

Specification point 7.3

For a reaction to happen, two things must be true:

COLLISION THEORY

1. The reactant particles must **collide** with each other.
2. They must collide with **enough energy**: at least the **activation energy**.

A collision with enough energy is a **successful collision**. Most collisions are **not** successful: the particles just bounce off each other.

So the rate of reaction depends on **how many successful collisions happen every second**. There are two ways to increase this:

Increase the...	What it means
Frequency of collisions	More collisions every second, so more of them are successful every second.
Energy of collisions	A bigger fraction of the collisions have energy equal to or greater than the activation energy.

ANALOGY: THE SCHOOL CORRIDOR

Picture students rushing along a corridor at break time. The more students there are (concentration), and the faster they rush (temperature), the more often they bump into each other, and the harder the bumps. A reaction is like a bump hard enough to knock books out of someone's hands.

Rate of reaction

DEFINITION

$$\text{Mean rate of reaction} = \frac{\text{amount of reactant used or product formed}}{\text{time taken}}$$

A reaction is **fastest at the start**, when there are the most reactant particles. As reactants are used up, collisions become less frequent and the reaction slows down. It **stops** when one reactant (the limiting reactant) is used up.

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COMMON MISTAKES

- Saying "particles collide more". Always say **more frequent collisions** or **more collisions per second**.
- Forgetting that particles need enough energy. A collision alone is not enough.
- Saying particles "have more energy" when the concentration increases. Concentration changes frequency only, not energy.

Worked examples

Worked Example 1.1 [2 marks]

State the two conditions needed for particles to react.

Step 1: Think about the particles

They must meet each other.

They must hit hard enough to break bonds.

ANSWER Particles must collide, with energy equal to or greater than the activation energy

Worked Example 1.2 [2 marks]

Explain why not every collision between reactant particles leads to a reaction.

Step 1: Compare the energy

Many collisions have **less** energy than the activation energy.

Step 2: Conclude

These particles just bounce apart without reacting, so the collision is not successful.

ANSWER Most collisions do not have enough energy (less than the activation energy), so they are not successful

Worked Example 1.3 [1 mark]

What is meant by the rate of a reaction?

Step 1: Give the definition

The rate is how quickly a reactant is used up or a product is formed.

ANSWER The amount of reactant used up or product formed per unit time

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Worked Example 1.4 [3 marks]

Explain, in terms of particles, why a reaction slows down as it goes on.

Step 1: What happens to the reactants

Reactant particles are used up as the reaction goes on.

Step 2: Effect on collisions

There are fewer reactant particles in the same volume, so the concentration falls.

Step 3: Effect on rate

Collisions become **less frequent**, so there are fewer successful collisions per second.

ANSWER Reactants are used up, so there are fewer collisions per second and the rate falls

Worked Example 1.5 [2 marks]

Temperature, concentration, pressure and surface area all increase the rate. Which factor increases the **energy** of collisions as well as their frequency?

Step 1: Check each factor

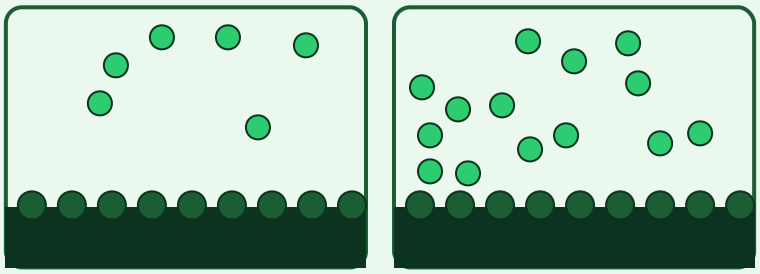
Concentration, pressure and surface area only make collisions **more frequent**.

Temperature makes particles move faster: collisions are more frequent **and** more energetic.

ANSWER Temperature

Worked Example 1.6 [2 marks]

The two diagrams show acid particles (light) near a solid. In which container will the reaction be faster? Explain.



Container A Container B

Step 1: Count the acid particles

Container A has 6 acid particles. Container B has 14 in the same volume.

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Step 2: Link to collisions

B has a higher concentration, so acid particles hit the solid **more frequently**.

More successful collisions per second means a faster reaction.

ANSWER Container B, because more particles in the same volume give more frequent collisions

Practice questions: Concept 1

1.1 [1 mark]
What is meant by a successful collision?

1.2 [2 marks]
Two particles collide with less energy than the activation energy. What happens? Explain.

1.3 [2 marks]
Explain why a reaction is fastest at the start.

1.4 [2 marks]
Explain why a reaction stops when one of the reactants is used up.

1.5 [3 marks]
Sort these factors into two groups: those that only increase the **frequency** of collisions, and those that increase **frequency and energy**: temperature, concentration, pressure, surface area.

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CONCEPT 2

Factors that Affect the Rate

Specification point 7.4

Change	What happens to the particles	Effect on collisions
Increase temperature	Particles move **faster** and have more energy.	More **frequent** collisions **and** more collisions have energy equal to or greater than the activation energy.
Increase concentration (solutions)	More particles in the **same volume**.	More **frequent** collisions.
Increase pressure (gases)	The same number of gas particles squeezed into a **smaller volume**.	More **frequent** collisions.
Increase surface area to volume ratio (solids)	Cutting a solid into smaller pieces or a powder exposes **more particles on the surface**.	More **frequent** collisions.

EXAM SENTENCE BUILDER

Increasing the _____ means there are more particles in the same volume, or particles move faster, or more particles are exposed, so there are **more frequent collisions** (and, for temperature only, **more collisions with energy greater than or equal to the activation energy**), so there are **more successful collisions per second** and the rate increases.

Surface area to volume ratio

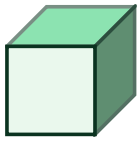
Only particles on the **surface** of a solid can be hit by the other reactant. Breaking a solid into smaller pieces increases the surface area to volume ratio.

FORMULAS FOR A CUBE OF SIDE A

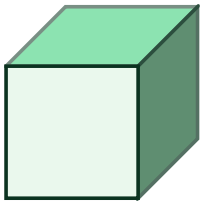
Surface area = $6 \times a \times a$ || Volume = $a \times a \times a$ || Ratio = $\frac{\text{surface area}}{\text{volume}}$



1 cm cube



2 cm cube



3 cm cube

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COMMON MISTAKES

- Saying a powder has a "bigger surface area" without linking it to **more particles exposed** and **more frequent collisions**.
- Saying higher temperature only makes particles "collide more". You must also mention **energy** for full marks.
- Using pressure for liquids or solids. Pressure only affects reactions involving **gases**.

Worked examples

Worked Example 2.1 [4 marks]

Explain why increasing the temperature increases the rate of reaction.

Step 1: Effect on particles

The particles gain kinetic energy and move faster.

Step 2: Frequency

They collide **more frequently**.

Step 3: Energy

A greater proportion of collisions have energy equal to or greater than the activation energy.

Step 4: Conclude

So there are more successful collisions per second, and the rate increases.

ANSWER Particles move faster: more frequent collisions and more collisions with enough energy, so more successful collisions per second

Worked Example 2.2 [3 marks]

Marble chips react faster with 2.0 mol dm^{-3} hydrochloric acid than with 0.5 mol dm^{-3} acid. Explain why.

Step 1: Compare the acids

The 2.0 mol dm^{-3} acid has more acid particles in the same volume.

Step 2: Link to collisions

Acid particles collide with the marble **more frequently**.

Step 3: Conclude

More successful collisions per second, so the rate is higher.

ANSWER Higher concentration: more particles in the same volume, more frequent collisions, faster rate

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Worked Example 2.3 [3 marks]

Explain why increasing the pressure speeds up a reaction between two gases.

Step 1: Effect on particles

The same number of gas particles are in a smaller volume, so they are closer together.

Step 2: Link to collisions

They collide more frequently.

Step 3: Conclude

More successful collisions per second, so the rate increases.

ANSWER Particles are closer together, so collisions are more frequent and the rate increases

Worked Example 2.4 [3 marks]

Calculate the surface area to volume ratio of a cube with sides of 2 cm.

Step 1: Surface area

A cube has 6 square faces.

Surface area = $6 \times 2 \times 2$

Surface area = $24\ cm^2$

Step 2: Volume

Volume = $2 \times 2 \times 2$

Volume = $8\ cm^3$

Step 3: Ratio

Ratio = $\frac{24}{8}$

Ratio = 3 (written 3 : 1)

ANSWER Surface area to volume ratio = 3 : 1

Worked Example 2.5 [4 marks]

The 2 cm cube from Worked Example 2.4 is cut into eight 1 cm cubes. Calculate the new surface area to volume ratio and explain the effect on the rate of reaction with acid.

Step 1: Surface area of one small cube

$6 \times 1 \times 1 = 6\ cm^2$

Step 2: Total surface area

$8 \times 6 = 48\ cm^2$

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Step 3: Volume (unchanged)

$8 \times (1 \times 1 \times 1) = 8\ cm^3$

Step 4: Ratio

$$\text{Ratio} = \frac{48}{8} = 6\ (6 : 1)$$

Step 5: Compare and explain

The ratio has doubled from 3 : 1 to 6 : 1.

More particles are exposed to the acid, so collisions are more frequent and the rate increases.

ANSWER Ratio = 6 : 1 (doubled), so the reaction is faster

Worked Example 2.6 [3 marks]

A student reacts 5 g of large marble chips and 5 g of powdered marble with the same acid. Predict and explain which reacts faster.

Step 1: Compare the solids

Powder is made of much smaller pieces, so it has a much bigger surface area to volume ratio.

Step 2: Link to collisions

More marble particles are exposed to the acid, so collisions are more frequent.

Step 3: Conclude

More successful collisions per second: the powder reacts faster.

ANSWER The powder reacts faster because it has a larger surface area to volume ratio

Practice questions: Concept 2

2.1 [4 marks]

Magnesium ribbon reacts with hydrochloric acid at 20 °C and at 40 °C. Explain why the reaction is faster at 40 °C.

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2.2 [3 marks]
Calculate the surface area to volume ratio of a cube with sides of 3 cm.

2.3 [3 marks]
Calculate the surface area to volume ratio of a cuboid measuring 4 cm by 2 cm by 1 cm.

2.4 [3 marks]
Explain why increasing the pressure of a gas reaction increases the rate.

2.5 [3 marks]
A student dilutes some acid with water before adding zinc. Explain what happens to the rate of reaction.

2.6 [2 marks]
Fine flour dust in a mill can explode if it meets a spark, but a bag of flour burns slowly. Suggest why.

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CONCEPT 3

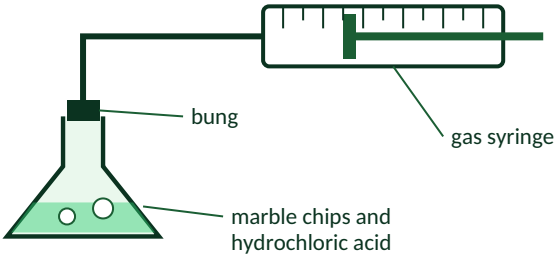
Measuring the Rate: Core Practical

Specification points 7.1 and 7.2

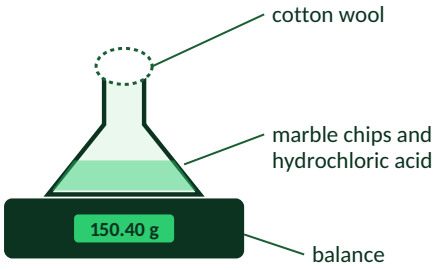
To measure a rate, we follow something that **changes** during the reaction: the volume of gas, the mass of the flask, or how cloudy the solution becomes.

Method A: measuring the gas produced (marble chips and acid)

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$



Collecting carbon dioxide in a gas syringe

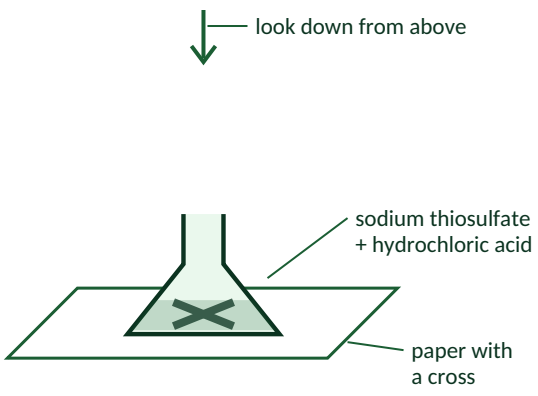


Measuring mass loss on a balance

- **Gas syringe:** read the volume of carbon dioxide every 10 seconds. (An upturned measuring cylinder full of water works too.)
- **Mass loss:** the carbon dioxide escapes, so the mass on the balance **falls**. The cotton wool lets gas out but stops acid spraying out.
- Change one factor (for example concentration of acid or size of chips) and keep everything else the same.

Method B: the disappearing cross (sodium thiosulfate and acid)

$$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + S + SO_2 + H_2O$$



The sulfur makes the mixture cloudy

1. Draw a black cross on paper and stand a conical flask on it.
2. Add sodium thiosulfate solution, then hydrochloric acid, and start the stopwatch.
3. Look down through the flask. Stop timing when the cross **disappears**.
4. Repeat at different temperatures or concentrations.

Solid **sulfur** forms, which makes the mixture cloudy (a colour change from clear to pale yellow).

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RATE FORMULAS

$$\text{Mean rate} = \frac{\text{volume of gas or mass lost}}{\text{time}} \quad || \quad \text{Rate (disappearing cross)} = \frac{1}{\text{time}}$$

Units: $cm^3\ s^{-1}$ for gas volume, $g\ s^{-1}$ for mass, s^{-1} for 1 divided by time. A **shorter** time means a **faster** rate.

Variables

Type	Meaning	Example (changing acid concentration)
Independent	What you change	Concentration of acid
Dependent	What you measure	Volume of gas, or time for cross to disappear
Control	What you keep the same	Temperature, volume of acid, mass and size of chips

COMMON MISTAKES

- Using the mass loss method for hydrogen. Hydrogen is so light that the change in mass is too small to measure accurately. Use a gas syringe.
- Saying a longer time means a faster rate. It is the opposite: rate = 1 divided by time.
- Different people judge "cross disappeared" differently. The same person should watch every time.

Worked examples

Worked Example 3.1 [2 marks]

Marble chips react with acid and 48 cm^3 of carbon dioxide is collected in the first 60 seconds. Calculate the mean rate of reaction.

Step 1: Write down the values

Amount = 48 cm^3

Time = 60 s

Step 2: Use the formula

$$\text{Mean rate} = \frac{\text{amount}}{\text{time}}$$

$$\text{Mean rate} = \frac{48}{60}$$

$$\text{Mean rate} = 0.8\ cm^3\ s^{-1}$$

ANSWER Mean rate = 0.8 $cm^3\ s^{-1}$

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Worked Example 3.2 [3 marks]

In a mass loss experiment, the mass falls from 150.40 g to 149.68 g in 120 seconds. Calculate the mean rate of reaction.

Step 1: Find the mass lost

Mass lost = 150.40 − 149.68

Mass lost = 0.72 g

Step 2: Use the formula

$$\text{Mean rate} = \frac{\text{mass lost}}{\text{time}}$$

$$\text{Mean rate} = \frac{0.72}{120}$$

$$\text{Mean rate} = 0.006 \text{ g s}^{-1}$$

ANSWER Mean rate = 0.006 g s^{-1}

Worked Example 3.3 [2 marks]

In the disappearing cross experiment the cross disappears after 40 seconds. Calculate the rate.

Step 1: Use the formula

$$\text{Rate} = \frac{1}{\text{time}}$$

$$\text{Rate} = \frac{1}{40}$$

$$\text{Rate} = 0.025 \text{ s}^{-1}$$

ANSWER Rate = 0.025 s^{-1}

Worked Example 3.4 [4 marks]

Results for the disappearing cross at different temperatures are shown. Calculate the rate at each temperature and describe the pattern.

Temperature (°C)	20	30	40
Time (s)	80	40	20

Step 1: 20 °C

$$\text{Rate} = \frac{1}{80} = 0.0125 \text{ s}^{-1}$$

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Step 2: 30 °C

$$\text{Rate} = \frac{1}{40} = 0.025\ s^{-1}$$

Step 3: 40 °C

$$\text{Rate} = \frac{1}{20} = 0.05\ s^{-1}$$

Step 4: Describe the pattern

As temperature increases, the time decreases and the rate increases.

Here, every 10 °C rise **doubles** the rate (0.0125 → 0.025 → 0.05).

ANSWER 0.0125, 0.025 and 0.05 s^{-1}; the rate doubles for every 10 °C rise

Worked Example 3.5 [3 marks]

A student investigates how the concentration of hydrochloric acid affects the rate of reaction with marble chips, using a gas syringe. Identify the independent, dependent and two control variables.

Step 1: Independent (what is changed)

Concentration of hydrochloric acid

Step 2: Dependent (what is measured)

Volume of carbon dioxide collected (in a set time)

Step 3: Control (kept the same)

Temperature, volume of acid, mass of marble chips, size of marble chips

ANSWER IV: acid concentration. DV: volume of CO_2. CVs: temperature, volume of acid, mass and size of chips

Worked Example 3.6 [3 marks]

Magnesium reacts with hydrochloric acid to make hydrogen gas. Suggest the best method to measure the rate, and explain why the mass loss method is not suitable.

Step 1: Choose a method

Collect the hydrogen in a **gas syringe** and record the volume every 10 seconds.

Step 2: Explain the problem with mass loss

Hydrogen has a very small mass, so the balance reading would barely change.

The change would be too small to measure accurately.

ANSWER Use a gas syringe; hydrogen is too light for mass loss to be measured accurately

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Practice questions: Concept 3

3.1 [2 marks]
60 cm^3 of gas is collected in 40 seconds. Calculate the mean rate of reaction.

3.2 [3 marks]
The mass of a flask falls from 85.30 g to 84.40 g in 150 seconds. Calculate the mean rate of reaction.

3.3 [2 marks]
In a disappearing cross experiment the cross disappears after 25 seconds. Calculate the rate.

3.4 [3 marks]
A student investigates the effect of temperature on the reaction between sodium thiosulfate and hydrochloric acid. Name the independent variable, the dependent variable and one control variable.

3.5 [3 marks]
Zinc reacts with sulfuric acid to produce hydrogen. Suggest a method to measure the rate of this reaction and name the equipment you would use.

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3.6 [4 marks]

Calculate the rate for each concentration, then describe the trend.

Concentration of thiosulfate	Low	Medium	High
Time for cross to disappear (s)	50	25	20

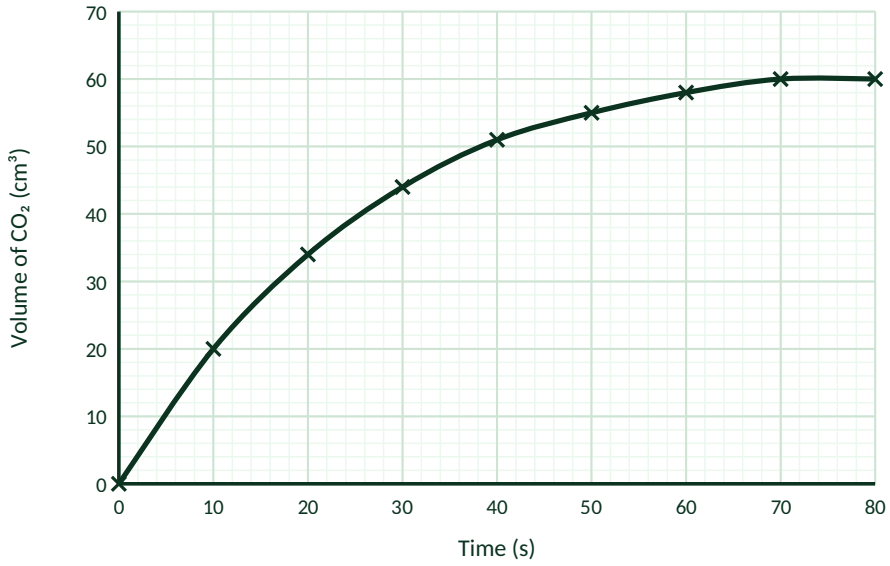
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CONCEPT 4

Interpreting Rate Graphs

Specification point 7.5

When we plot the amount of product (or reactant) against time, the **shape** of the graph tells the story of the reaction.



Volume of carbon dioxide made when marble chips react with acid

Part of the graph	What it tells you
Steep at the start	The rate is fastest: lots of reactant particles, frequent collisions.
Gets less steep	The rate is slowing down as reactants are used up.
Flat (levels off)	The reaction has **stopped**: one reactant is used up.
Final height	The total amount of product. This depends on the **amount** of limiting reactant, not on the rate.

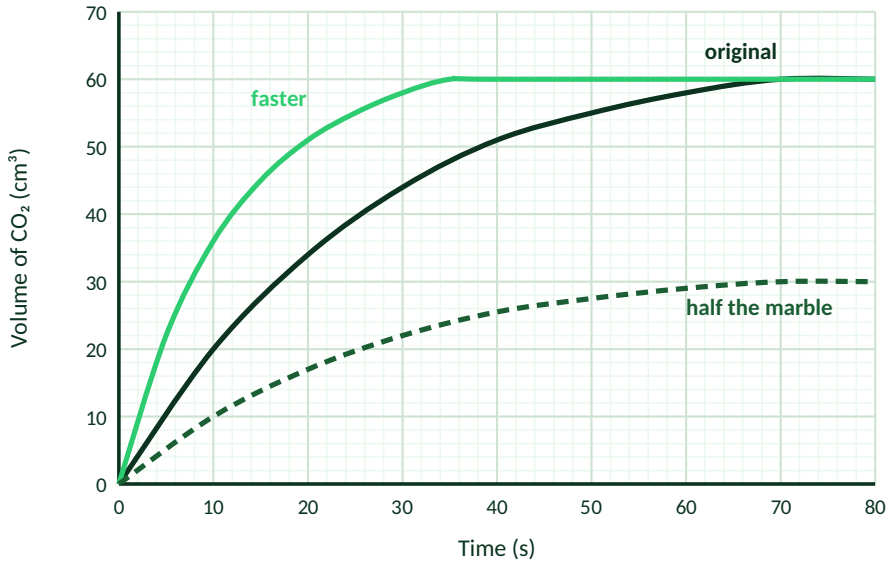
RATE FROM A GRAPH

Mean rate between two times = $\frac{\text{change in volume}}{\text{change in time}}$

Rate at one moment = gradient of the **tangent** at that time = $\frac{\text{change in y}}{\text{change in x}}$ (pick two points far apart on the tangent)

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Comparing curves



A faster reaction is steeper and finishes sooner. Half the limiting reactant gives half the gas.

- **Faster reaction** (higher temperature, more concentrated acid, powder, catalyst): curve is **steeper** and levels off **sooner**, but at the **same final volume** if the amounts are the same.
- **Less limiting reactant**: curve levels off at a **lower** final volume.
- **Mass loss graphs** go **down** and then level off, because gas escapes from the flask.

COMMON MISTAKES

- Saying a faster reaction makes **more** product. It makes the same amount, just more quickly.
- Reading the scale wrongly. Check what each small square is worth before reading values.
- Drawing a tangent that cuts through the curve. It should just touch the curve at the point.

Worked examples

Worked Examples 4.1 to 4.5 use the first graph in this concept.

Worked Example 4.1 [2 marks]

Use the graph to find (a) the total volume of gas made (b) the time when the reaction stopped.

Step 1: Find where the graph levels off

The graph becomes flat at 60 cm^3.

Step 2: Read the time

The graph first reaches 60 at 70 s.

ANSWER (a) 60 cm^3 (b) 70 s

Worked Example 4.2 [2 marks]

Calculate the mean rate of reaction during the first 20 seconds.

Step 1: Read the graph

At 0 s: 0 cm^3. At 20 s: 34 cm^3.

Step 2: Use the formula

$$\text{Mean rate} = \frac{34 - 0}{20 - 0}$$

$$\text{Mean rate} = \frac{34}{20}$$

$$\text{Mean rate} = 1.7 \text{ cm}^3 \text{ s}^{-1}$$

ANSWER 1.7 cm^3 s^{-1}

Worked Example 4.3 [2 marks]

Calculate the mean rate between 20 s and 40 s. Compare it with your answer to 4.2.

Step 1: Read the graph

At 20 s: 34 cm^3. At 40 s: 51 cm^3.

Step 2: Use the formula

$$\text{Mean rate} = \frac{51 - 34}{40 - 20}$$

$$\text{Mean rate} = \frac{17}{20}$$

$$\text{Mean rate} = 0.85 \text{ cm}^3 \text{ s}^{-1}$$

Step 3: Compare

0.85 is half of 1.7: the reaction has slowed down because reactants are being used up.

ANSWER 0.85 cm^3 s^{-1}, slower than in the first 20 s

Worked Example 4.4 [2 marks]

Calculate the mean rate for the whole reaction. Give your answer to 2 decimal places.

Step 1: Use the totals

Total volume = 60 cm^3, time taken = 70 s

Step 2: Use the formula

$$\text{Mean rate} = \frac{60}{70}$$

Mean rate = 0.857142...

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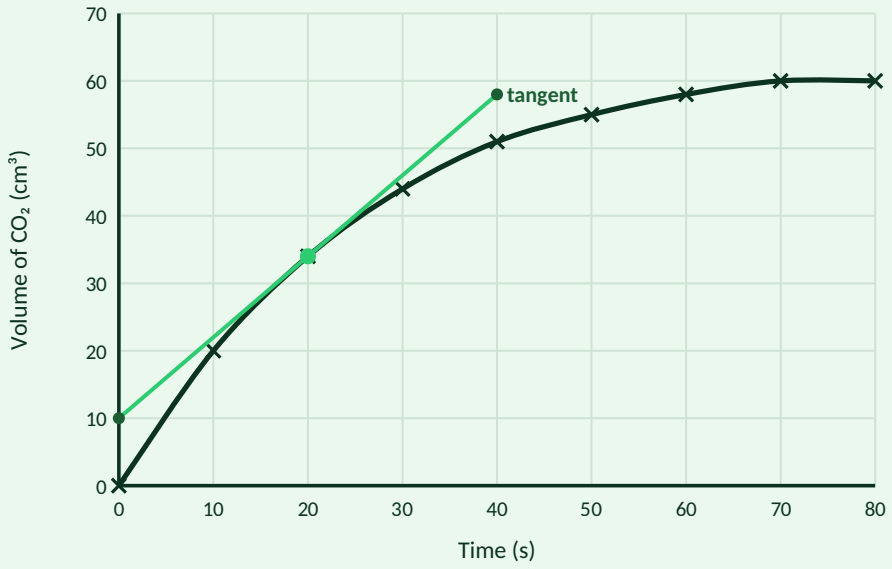
Step 3: Round

Mean rate = 0.86 $cm^3\ s^{-1}$ (2 d.p.)

ANSWER 0.86 $cm^3\ s^{-1}$ (2 d.p.)

Worked Example 4.5 [3 marks]

A tangent has been drawn on the curve at 20 s. Use it to find the rate of reaction at 20 s.



Step 1: Choose two points on the tangent, far apart

(0, 10) and (40, 58)

Step 2: Find the changes

Change in volume = 58 − 10 = 48

Change in time = 40 − 0 = 40

Step 3: Gradient

Rate = $\frac{48}{40}$

Rate = 1.2 $cm^3\ s^{-1}$

ANSWER Rate at 20 s = 1.2 $cm^3\ s^{-1}$

Worked Example 4.6 [4 marks]

Look at the comparing curves graph. (a) Why does the "faster" curve level off at the same volume as the original? (b) Why does the "half the marble" curve level off at 30 cm^3?

Step 1: Faster curve

The same amounts of reactants were used, only the rate changed (for example using powder).

So the same total volume of gas is made, just in less time.

Step 2: Half the marble

Marble is the limiting reactant.

Half the marble can only make half the gas: half of 60 = 30 cm^3.

ANSWER (a) Same amount of reactant, so same amount of product (b) Half the limiting reactant gives half the product

Worked Example 4.7 **[3 marks]**

Sketch and explain the shape of a mass against time graph for marble chips and acid on a balance.

Step 1: Starting point

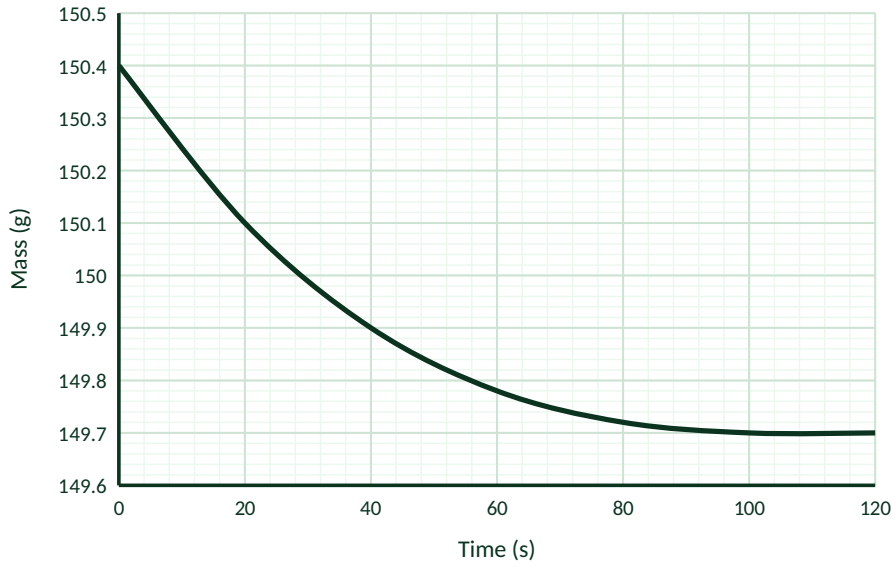
The graph starts at the total starting mass.

Step 2: Shape

Mass falls quickly at first (steep), then more slowly, then levels off.

Step 3: Explain

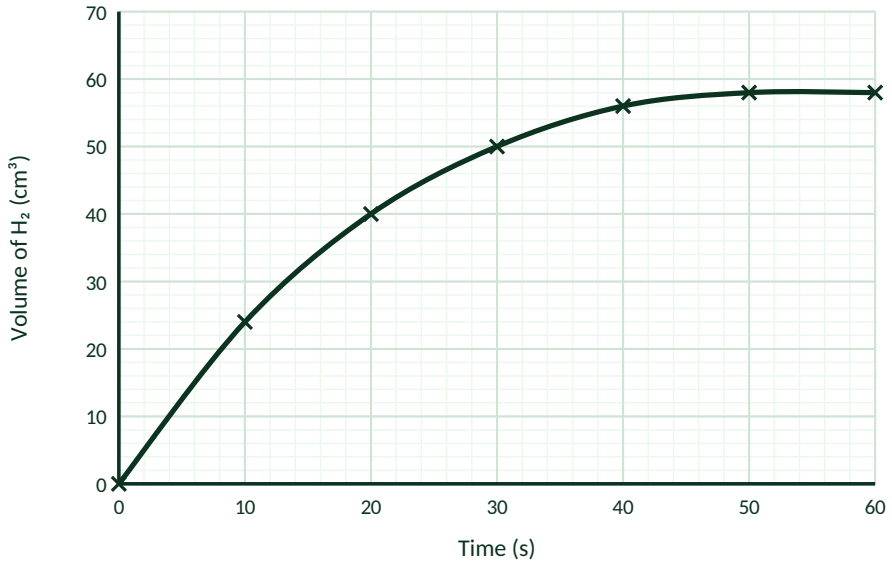
Carbon dioxide escapes, so the mass falls. The reaction slows as reactants are used up, and stops when the limiting reactant is used up.



ANSWER A curve going down, steep at first, then levelling off when the reaction stops

Practice questions: Concept 4

Questions 4.1 to 4.4 use this graph for magnesium reacting with acid.



Volume of hydrogen against time

4.1 [2 marks]

Use the graph to find (a) the total volume of hydrogen made (b) the time when the reaction stopped.

4.2 [2 marks]

Calculate the mean rate of reaction in the first 10 seconds.

4.3 [2 marks]

Calculate the mean rate between 10 s and 30 s.

4.4 [3 marks]

A tangent drawn at 20 s passes through (0, 14) and (40, 66). Calculate the rate at 20 s.

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4.5 [3 marks]

The experiment is repeated with the same amount of magnesium but at a higher temperature. Describe how the curve would be different, and explain why.

4.6 [3 marks]

In a mass loss experiment the mass falls by 0.80 g in the first 60 seconds. Calculate the mean rate. Give your answer to 2 significant figures.

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CONCEPT 5

Catalysts and Enzymes

Specification points 7.6, 7.7 and 7.8

CATALYST

A **catalyst** is a substance that **speeds up** the rate of a reaction **without altering the products**, and is **unchanged chemically and in mass** at the end of the reaction.

Because it is not used up, a catalyst is **not** written as a reactant in the equation (it may be written above the arrow), and a small amount can be used again and again.

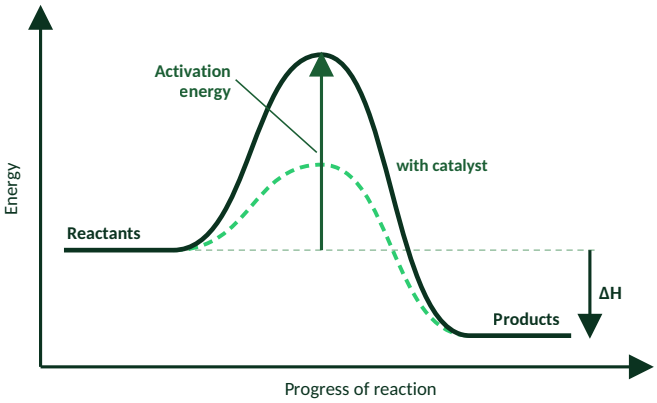
How does a catalyst work?

A catalyst provides a **different pathway** for the reaction with a **lower activation energy**.

This means a **greater proportion** of particles have energy equal to or greater than the activation energy when they collide.

So there are **more successful collisions per second** and the rate increases.

The reactants and products are the same, so the overall energy change (ΔH) is **unchanged**.



The dashed curve shows the lower hump with a catalyst

Enzymes

Enzymes are **biological catalysts**. They are proteins made by living cells and they speed up reactions in living things, usually at fairly low temperatures.

Use	How enzymes help
Alcoholic drinks (beer, wine)	Enzymes in **yeast** convert sugar (glucose) into ethanol and carbon dioxide. This is **fermentation**.
Bread making	The same yeast enzymes make carbon dioxide, which makes the dough rise.
Biological washing powders	Enzymes break down food stains at low temperatures.

Fermentation: glucose → ethanol + carbon dioxide || $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

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COMMON MISTAKES

- Saying a catalyst "gives particles more energy". It does not. It **lowers the activation energy**.
- Saying a catalyst is "not involved" in the reaction. It takes part, but is turned back into its original form by the end.
- Saying a catalyst changes ΔH or the amount of product. It changes neither.

Worked examples

Worked Example 5.1 [2 marks]

Describe what is meant by a catalyst.

Step 1: Give the three key points

It speeds up the reaction.

It does not change the products.

It is unchanged chemically and in mass at the end.

ANSWER A substance that speeds up a reaction without changing the products, and is unchanged chemically and in mass at the end

Worked Example 5.2 [3 marks]

Explain how adding a catalyst increases the rate of reaction.

Step 1: Pathway

The catalyst provides an alternative pathway with a **lower activation energy**.

Step 2: Energy of particles

More of the colliding particles have energy equal to or greater than this lower activation energy.

Step 3: Conclude

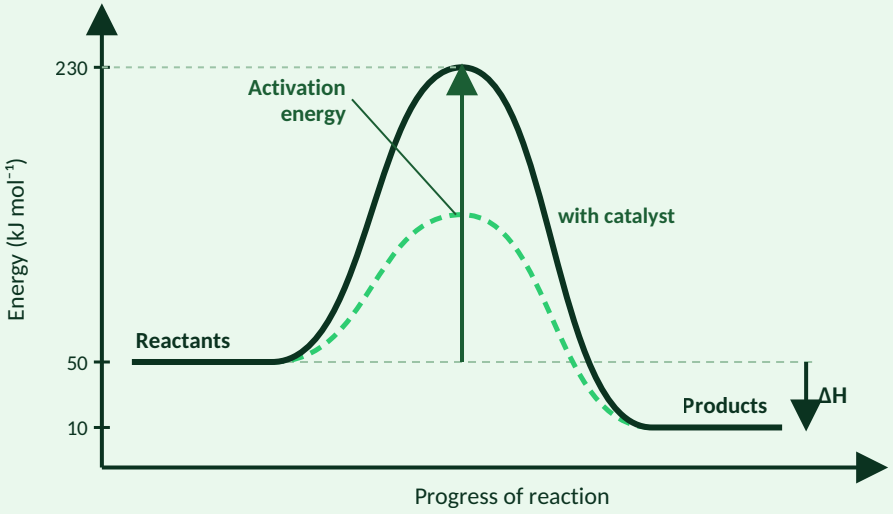
More successful collisions per second, so the rate increases.

ANSWER Lower activation energy, so more collisions are successful each second

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Worked Example 5.3 [4 marks]

Use the profile to find (a) the activation energy without a catalyst (b) the activation energy with a catalyst (c) ΔH.



Step 1: Read the values

Reactants = 50, uncatalysed top = 230, catalysed top = 140, products = 10 (all in kJ mol^{-1})

Step 2: Without catalyst

Ea = 230 − 50 = 180 kJ mol^{-1}

Step 3: With catalyst

Ea = 140 − 50 = 90 kJ mol^{-1}

Step 4: Overall energy change

ΔH = 10 − 50 = −40 kJ mol^{-1} (the same with or without a catalyst)

ANSWER (a) 180 kJ mol^{-1} (b) 90 kJ mol^{-1} (c) −40 kJ mol^{-1}

Worked Example 5.4 [3 marks]

Manganese(IV) oxide speeds up the decomposition of hydrogen peroxide. A student uses 0.50 g of the black powder, then filters, dries and weighs it at the end: it is still 0.50 g. What does this show?

Step 1: Compare the masses

Mass at start = 0.50 g. Mass at end = 0.50 g. No change.

Step 2: Conclude

The powder speeds up the reaction but is not used up.

This is evidence that manganese(IV) oxide is a **catalyst**.

ANSWER The mass is unchanged, so manganese(IV) oxide acts as a catalyst

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Worked Example 5.5 [3 marks]

Explain how enzymes are used in the production of alcoholic drinks.

Step 1: What enzymes are

Enzymes are biological catalysts found in living things.

Step 2: Where they come from

Yeast contains enzymes.

Step 3: What they do

They catalyse fermentation: glucose → ethanol + carbon dioxide. The ethanol is the alcohol in the drink.

ANSWER Yeast enzymes (biological catalysts) convert sugar into ethanol and carbon dioxide by fermentation

Worked Example 5.6 [2 marks]

Suggest two reasons why catalysts are useful in industry.

Step 1: Speed

Products are made faster, so more product is made each day.

Step 2: Energy and cost

Reactions can run at lower temperatures, which saves energy and money, and the catalyst can be reused.

ANSWER Faster production; lower temperatures save energy and money (and the catalyst is reused)

Practice questions: Concept 5

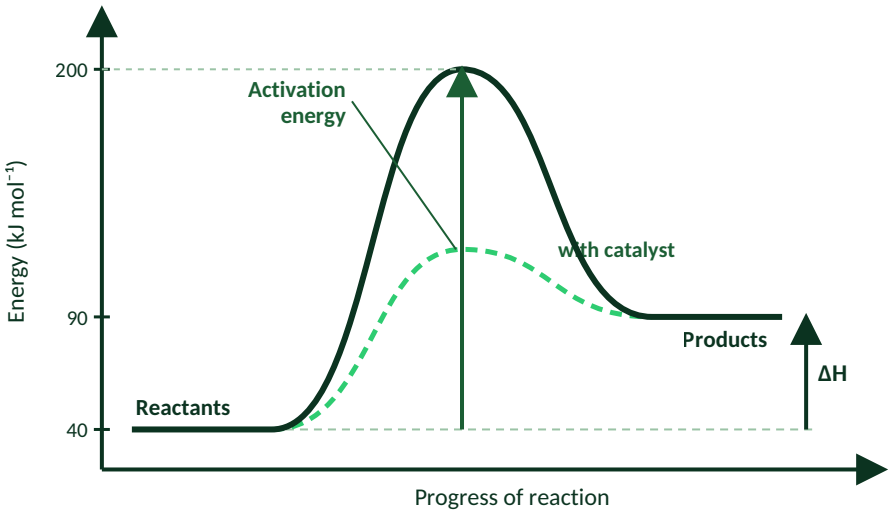
5.1 [2 marks]
Describe what is meant by a catalyst.

5.2 [3 marks]
Explain, in terms of activation energy, how a catalyst increases the rate of a reaction.

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5.3 [4 marks]

Use the profile to find (a) the activation energy without a catalyst (b) the activation energy with a catalyst (c) ΔH, and state if the reaction is exothermic or endothermic.



5.4 [2 marks]

A student adds 2.00 g of a powder to a reaction. At the end, 2.00 g of the same powder is recovered. The reaction was faster with the powder. What can you conclude?

5.5 [3 marks]

What are enzymes? Name the process in which enzymes are used to make alcoholic drinks, and give its word equation.

5.6 [3 marks]

Hydrogen peroxide decomposes. Without a catalyst, 12 cm^3 of oxygen is made in 30 s. With a catalyst, 45 cm^3 is made in 30 s. Calculate both mean rates and how many times faster the catalysed reaction is.

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WRAP UP

Summary and Exam Checklist

Spec	What you must know
7.1	Core practical: gas syringe or mass loss for marble chips and acid; disappearing cross for sodium thiosulfate and acid. Rate = $\frac{\text{amount}}{\text{time}}$, or $\frac{1}{\text{time}}$ for the cross.
7.2	Pick a method that follows something that changes: gas volume, mass, colour or cloudiness. Use a gas syringe for light gases like hydrogen.
7.3	Particles must collide with energy at least equal to the activation energy. More frequent or more energetic collisions give a faster rate.
7.4	Temperature: frequency and energy. Concentration, pressure and surface area to volume ratio: frequency only.
7.5	Steep = fast, flat = finished, final height = amount of product. Mean rate from two points; rate at a moment from a tangent.
7.6	Catalyst: speeds up, same products, unchanged chemically and in mass.
7.7	Catalyst gives a pathway with lower activation energy, so more successful collisions per second.
7.8	Enzymes are biological catalysts; yeast enzymes ferment sugar into ethanol for alcoholic drinks.

Self check

I can...	Not yet	Nearly	Got it
describe both core practicals and identify variables			
calculate mean rates, including 1 divided by time			
explain all four factors using collision theory			
calculate a surface area to volume ratio			
read rate graphs and find a gradient from a tangent			
explain how catalysts work and where enzymes are used			

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Answer Key

Answer Summary first, then full worked solutions for every practice question

Answer Summary

Q	Final answer	Q	Final answer
1.1	A collision with energy at least equal to the activation energy, causing a reaction	**3.5**	Gas syringe, read volume of H_2 every 10 s
1.2	No reaction; particles bounce apart	**3.6**	0.02, 0.04, 0.05 s^{-1}; rate rises with concentration
1.3	Most reactant particles, so most frequent collisions	**4.1**	(a) 58 cm^3 (b) 50 s
1.4	No particles of that reactant left to collide	**4.2**	2.4 cm^3 s^{-1}
1.5	Frequency only: concentration, pressure, surface area. Both: temperature	**4.3**	1.3 cm^3 s^{-1}
2.1	Faster particles: more frequent and more energetic collisions	**4.4**	1.3 cm^3 s^{-1}
2.2	2 : 1	**4.5**	Steeper, levels off sooner, same final volume (58)
2.3	3.5 : 1 (7 : 2)	**4.6**	0.013 g s^{-1} (2 s.f.)
2.4	Particles closer together, more frequent collisions	**5.1**	Speeds up reaction; products unchanged; unchanged chemically and in mass
2.5	Rate decreases: fewer particles per volume	**5.2**	Lower activation energy pathway, more successful collisions
2.6	Dust has a huge surface area to volume ratio	**5.3**	(a) 160 (b) 80 (c) +50 kJ mol^{-1}, endothermic
3.1	1.5 cm^3 s^{-1}	**5.4**	The powder is a catalyst
3.2	0.006 g s^{-1}	**5.5**	Biological catalysts; fermentation: glucose → ethanol + carbon dioxide
3.3	0.04 s^{-1}	**5.6**	0.4 and 1.5 cm^3 s^{-1}; 3.75 times faster
3.4	IV: temperature; DV: time for cross to disappear; CV: e.g. volumes and concentrations		

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Worked solutions: Concept 1

Question 1.1 [1 mark]

What is a successful collision?

Step 1: Definition

A collision between particles with energy equal to or greater than the activation energy, so they react.

ANSWER A collision with enough energy (at least the activation energy) to cause a reaction

Question 1.2 [2 marks]

Collision with less energy than the activation energy.

Step 1: Compare the energy

The energy is less than the activation energy, so bonds cannot break.

Step 2: Conclude

The particles bounce apart: no reaction.

ANSWER No reaction, because the collision does not have enough energy

Question 1.3 [2 marks]

Why is the rate fastest at the start?

Step 1: Number of particles

At the start, the concentration of reactant particles is highest.

Step 2: Collisions

So collisions are most frequent, giving the most successful collisions per second.

ANSWER Most reactant particles, so the most frequent collisions

Question 1.4 [2 marks]

Why does a reaction stop?

Step 1: Limiting reactant

When one reactant is used up, there are none of its particles left.

Step 2: Collisions

No collisions between the reactants can happen, so no more product forms.

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ANSWER No particles of the used up reactant remain to collide

Question 1.5 [3 marks]

Sort the factors.

Step 1: Frequency only

Concentration, pressure and surface area put more particles where they can collide, but do not change their energy.

Step 2: Frequency and energy

Temperature makes particles move faster, so it changes both.

ANSWER Frequency only: concentration, pressure, surface area. Frequency and energy: temperature

Worked solutions: Concept 2

Question 2.1 [4 marks]

Why is the reaction faster at 40 °C?

Step 1: Particles

At 40 °C the particles have more kinetic energy and move faster.

Step 2: Frequency

They collide more frequently.

Step 3: Energy

More collisions have energy equal to or greater than the activation energy.

Step 4: Conclude

More successful collisions per second, so the rate is faster.

ANSWER More frequent collisions and more collisions with enough energy, so more successful collisions per second

Question 2.2 [3 marks]

Cube with sides of 3 cm.

Step 1: Surface area

$6 \times 3 \times 3 = 54\ cm^2$

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Step 2: Volume

$3 \times 3 \times 3 = 27\ cm^3$

Step 3: Ratio

$\frac{54}{27} = 2$

ANSWER Surface area to volume ratio = 2 : 1

Question 2.3 [3 marks]

Cuboid 4 cm by 2 cm by 1 cm.

Step 1: Surface area

Two faces of $4 \times 2 = 8$, two faces of $4 \times 1 = 4$, two faces of $2 \times 1 = 2$

Surface area = $2 \times (8 + 4 + 2) = 2 \times 14 = 28\ cm^2$

Step 2: Volume

$4 \times 2 \times 1 = 8\ cm^3$

Step 3: Ratio

$\frac{28}{8} = \frac{7}{2} = 3.5$

ANSWER Surface area to volume ratio = 3.5 : 1 (or 7 : 2)

Question 2.4 [3 marks]

Why does higher pressure increase the rate?

Step 1: Particles

The gas particles are pushed into a smaller volume, so they are closer together.

Step 2: Collisions

They collide more frequently.

Step 3: Conclude

More successful collisions per second, so the rate increases.

ANSWER Closer particles, more frequent collisions, faster rate

Question 2.5 [3 marks]

Effect of diluting the acid.

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Step 1: Concentration

Adding water lowers the concentration: fewer acid particles in the same volume.

Step 2: Collisions

Collisions with the zinc become less frequent.

Step 3: Conclude

Fewer successful collisions per second, so the rate **decreases**.

ANSWER The rate decreases because collisions are less frequent

Question 2.6 [2 marks]

Why does flour dust explode?

Step 1: Surface area

Fine dust has a huge surface area to volume ratio, so many particles are exposed to oxygen.

Step 2: Conclude

Collisions with oxygen are extremely frequent, so the reaction is so fast it becomes an explosion. The bag of flour has a small exposed surface.

ANSWER The dust has a very large surface area to volume ratio, so it reacts extremely quickly

Worked solutions: Concept 3

Question 3.1 [2 marks]

60 cm^3 of gas in 40 s.

Step 1: Write down the values

Amount = 60 cm^3

Time = 40 s

Step 2: Use the formula

$$\text{Mean rate} = \frac{\text{amount}}{\text{time}}$$

$$\text{Mean rate} = \frac{60}{40}$$

Mean rate = 1.5 $cm^3\ s^{-1}$

ANSWER Mean rate = 1.5 $cm^3\ s^{-1}$

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Question 3.2 [3 marks]

Mass falls from 85.30 g to 84.40 g in 150 s.

Step 1: Mass lost

85.30 − 84.40 = 0.90 g

Step 2: Use the formula

$$\text{Mean rate} = \frac{0.90}{150}$$

Mean rate = 0.006 g s^{-1}

ANSWER Mean rate = 0.006 g s^{-1}

Question 3.3 [2 marks]

Cross disappears after 25 s.

Step 1: Use the formula

$$\text{Rate} = \frac{1}{25}$$

Rate = 0.04 s^{-1}

ANSWER Rate = 0.04 s^{-1}

Question 3.4 [3 marks]

Variables in the thiosulfate temperature investigation.

Step 1: Independent

Temperature

Step 2: Dependent

Time for the cross to disappear

Step 3: Control (any one)

Volume and concentration of sodium thiosulfate, volume and concentration of acid, same cross, same person watching

ANSWER IV: temperature. DV: time for cross to disappear. CV: e.g. concentration of thiosulfate

Question 3.5 [3 marks]

Method for zinc and sulfuric acid.

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Step 1: Choose what to measure

Hydrogen gas is produced, so measure the volume of gas.

Step 2: Equipment and method

Connect a conical flask to a **gas syringe**. Add zinc to the acid, start a stopwatch and record the volume every 10 seconds.

Step 3: Why not mass loss

Hydrogen is too light for the mass change to be measured accurately.

ANSWER Collect the hydrogen in a gas syringe and record the volume at regular time intervals

Question 3.6 [4 marks]

Rates for times of 50, 25 and 20 s.

Step 1: Low

$\frac{1}{50} = 0.02\ s^{-1}$

Step 2: Medium

$\frac{1}{25} = 0.04\ s^{-1}$

Step 3: High

$\frac{1}{20} = 0.05\ s^{-1}$

Step 4: Trend

As the concentration increases, the rate increases.

ANSWER 0.02, 0.04, 0.05 s^{-1}; higher concentration gives a faster rate

Worked solutions: Concept 4

Question 4.1 [2 marks]

Total volume and stopping time.

Step 1: Where it levels off

The graph becomes flat at 58 cm^3.

Step 2: Time

It first reaches 58 at 50 s.

ANSWER (a) 58 cm^3 (b) 50 s

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Question 4.2 [2 marks]

Mean rate in the first 10 s.

Step 1: Read the graph

At 0 s: 0 cm^3. At 10 s: 24 cm^3.

Step 2: Formula

Mean rate = $\frac{24}{10}$ = 2.4 $cm^3\ s^{-1}$

ANSWER 2.4 $cm^3\ s^{-1}$

Question 4.3 [2 marks]

Mean rate between 10 s and 30 s.

Step 1: Read the graph

At 10 s: 24 cm^3. At 30 s: 50 cm^3.

Step 2: Formula

Mean rate = $\frac{50 - 24}{30 - 10}$

Mean rate = $\frac{26}{20}$ = 1.3 $cm^3\ s^{-1}$

ANSWER 1.3 $cm^3\ s^{-1}$

Question 4.4 [3 marks]

Tangent through (0, 14) and (40, 66).

Step 1: Changes

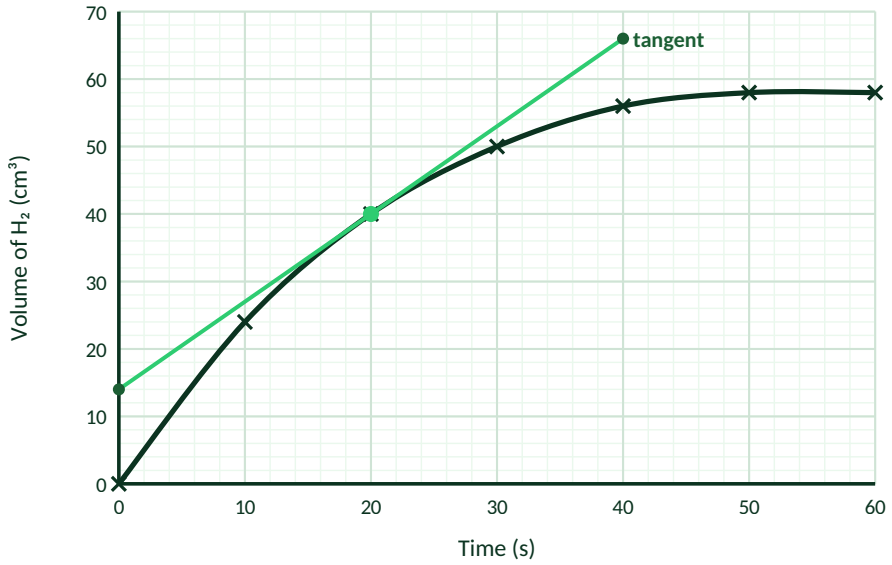
Change in volume = 66 − 14 = 52

Change in time = 40 − 0 = 40

Step 2: Gradient

Rate = $\frac{52}{40}$ = 1.3 $cm^3\ s^{-1}$

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ANSWER Rate at 20 s = 1.3 cm^3 s^{-1}

Question 4.5 [3 marks]

Higher temperature, same amount of magnesium.

Step 1: Shape

The curve is steeper at the start and levels off sooner.

Step 2: Final volume

It levels off at the same volume, 58 cm^3, because the same amount of magnesium is used.

Step 3: Reason

Particles move faster, so there are more frequent and more energetic collisions.

ANSWER Steeper curve, levels off earlier, same final volume of 58 cm^3

Question 4.6 [3 marks]

0.80 g lost in 60 s, answer to 2 s.f.

Step 1: Formula

$$\text{Mean rate} = \frac{0.80}{60}$$

Mean rate = 0.013333...

Step 2: Round

Mean rate = 0.013 g s^{-1} (2 s.f.)

ANSWER 0.013 g s^{-1} (2 s.f.)

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Worked solutions: Concept 5

Question 5.1 [2 marks]

Describe a catalyst.

Step 1: Key points

Speeds up the rate of reaction.

Does not change the products.

Unchanged chemically and in mass at the end.

ANSWER A substance that speeds up a reaction without changing the products and is unchanged chemically and in mass at the end

Question 5.2 [3 marks]

How does a catalyst work?

Step 1: Pathway

It provides an alternative pathway with a lower activation energy.

Step 2: Particles

A greater proportion of particles collide with energy equal to or greater than the activation energy.

Step 3: Conclude

More successful collisions per second, so the rate increases.

ANSWER Lower activation energy, so more collisions are successful per second

Question 5.3 [4 marks]

Reactants 40, uncatalysed top 200, catalysed top 120, products 90.

Step 1: Without catalyst

Ea = 200 − 40 = 160 kJ mol^{-1}

Step 2: With catalyst

Ea = 120 − 40 = 80 kJ mol^{-1}

Step 3: Energy change

ΔH = 90 − 40 = +50 kJ mol^{-1}

Step 4: Type

Products are higher than reactants: endothermic.

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ANSWER (a) 160 kJ mol^{-1} (b) 80 kJ mol^{-1} (c) +50 kJ mol^{-1}, endothermic

Question 5.4 [2 marks]

2.00 g in, 2.00 g out, faster reaction.

Step 1: Compare

The powder sped up the reaction and its mass did not change.

Step 2: Conclude

It was not used up, so it is a catalyst.

ANSWER The powder is a catalyst

Question 5.5 [3 marks]

Enzymes and alcoholic drinks.

Step 1: Enzymes

Enzymes are biological catalysts (proteins made by living cells).

Step 2: Process

Fermentation, using enzymes in yeast.

Step 3: Word equation

glucose → ethanol + carbon dioxide

ANSWER Biological catalysts; fermentation: glucose → ethanol + carbon dioxide

Question 5.6 [3 marks]

12 cm^3 and 45 cm^3 of oxygen in 30 s.

Step 1: Without catalyst

Mean rate = $\frac{12}{30}$ = 0.4 $cm^3\ s^{-1}$

Step 2: With catalyst

Mean rate = $\frac{45}{30}$ = 1.5 $cm^3\ s^{-1}$

Step 3: How many times faster

$\frac{1.5}{0.4}$ = 3.75

ANSWER 0.4 $cm^3\ s^{-1}$ and 1.5 $cm^3\ s^{-1}$; the catalysed reaction is 3.75 times faster

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