6 · The rate and extent of chemical change

AQA · GCSE · Chemistry

The rate and extent of chemical change

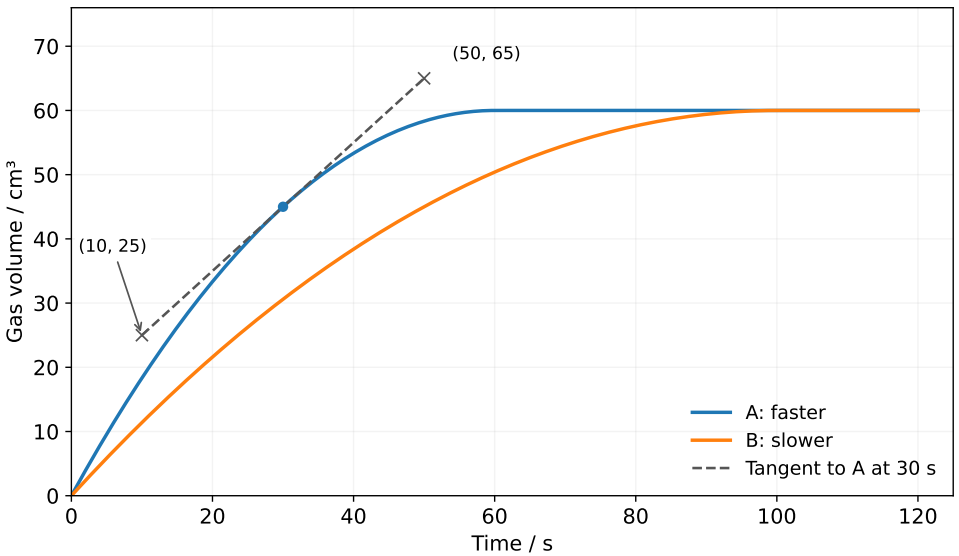
A crushed tablet can react quickly without making more gas than the same tablet left whole. A factory can speed a reaction without increasing its equilibrium yield. Separate **rate of reaction**, how fast change happens, from the amount finally obtained.

This reference covers AQA 8462 sections 4.6.1.1–4.6.2.7. Mean mass/volume rates, drawing/interpreting tangents and equilibrium are both-tier. Numerical tangent gradients, mol/s rates and equilibrium-shift predictions are Higher Tier (HT). Required practical 5 needs both supervised methods as well as written understanding.

Measure a rate (4.6.1.1)

Mean rate is quantity of reactant used divided by time, or quantity of product formed divided by time. These are alternative measurements, not necessarily equal numerical quantities. State the quantity and unit: g/s or cm^3/s; HT also uses mol/s. If gas volume changes from 12 to 36 cm^3 between 10 and 30 s, use $(36 - 12)/(30 - 10) = 1.2$ cm^3/s. Dividing 36 by 30 describes a different interval.

A product–time curve normally rises and flattens as a reactant runs out. A remaining-reactant curve falls and flattens. Compare steepness for rate and plateaux for final quantity. A lower plateau alone does not prove slower chemistry: a different limiting amount or collection loss can change the plateau.



A **tangent** follows the curve's local direction at the chosen time. Draw a straight line touching the curve there; extend it to read two separated points on the tangent. It is not a line joining two nearby points on the curve. Its gradient represents the **instantaneous rate** at that instant; steeper means faster.

HT worked example. The illustrated tangent passes through (10 s, 25 cm^3) and (50 s, 65 cm^3). Its gradient is $(65 - 25)/(50 - 10) = 1.0$ cm^3/s at 30 s. The mean rate from zero to 30 s is instead $45/30 = 1.5$ cm^3/s. The tangent can extend above the plateau: its extension is a local construction, not the later gas-volume curve.

HT molar example. Reactant amount falls from 0.018 to 0.006 mol in 40 s. Amount used is 0.012 mol, giving $0.012/40 = 0.00030$ mol/s. Use the positive amount consumed rather than a negative reaction rate from the falling graph.

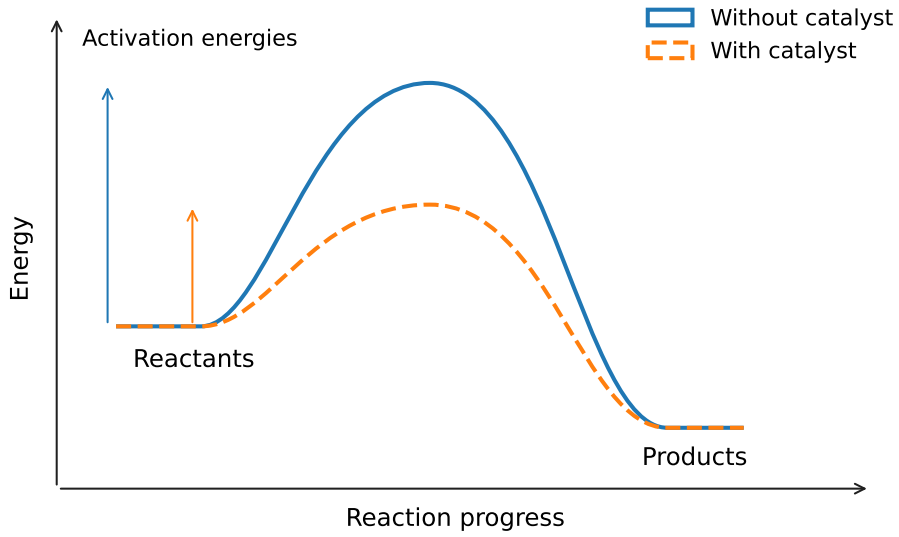
Explain rate changes (4.6.1.2–4.6.1.4)

Collision theory says reacting particles must collide with enough energy.
Activation energy is the minimum energy needed for reaction.

Change	Explanation under comparable conditions
Higher solution concentration	More reactant particles per volume; more frequent collisions.
Higher reacting-gas pressure at fixed temperature	Particles closer together; more frequent collisions.
Smaller pieces of the same solid mass	Larger surface-area-to-volume ratio; more particles exposed for collisions.
Higher temperature	Faster particles collide more frequently; a greater proportion has enough energy.
Add a suitable catalyst	An alternative pathway has lower activation energy.

At fixed temperature, higher concentration does not give each particle more kinetic energy. Breaking a lump into pieces does not increase its mass. The **surface-area-to-volume ratio** compares exposed area with solid volume. Use proportional reasoning only when justified: doubling particles per volume can explain more collisions, but does not prove every measured rate exactly doubles.

A **catalyst** changes rate and is not used up overall. Different reactions need suitable catalysts; **enzymes** are biological catalysts. Identify a catalyst from its effect on rate together with absence from the overall equation, not absence alone. A solvent may also be absent without being a catalyst.



The catalysed curve has a lower maximum relative to the reactants. Both start and finish at the same levels, so overall energy change is unchanged. A catalyst is not an energy source. Faster reaction does not mean more product from the same limiting amount; do not explain catalytic action by saying particles become hotter.

Required practical 5: two concentration methods (4.6.1.2)

Develop a hypothesis before measuring: higher acid concentration should give a steeper gas-volume curve with matched magnesium; higher thiosulfate concentration should reach the same cloudiness threshold sooner. Both approaches are required. Follow the approved school procedure; written answers do not certify participation.

Gas-volume method. Adopt the source-reviewed handbook activity: equal cleaned 3 cm magnesium ribbons react with 50 cm^3 teacher-prepared dilute hydrochloric acid at two concentrations. Collect hydrogen through a delivery tube into an inverted water-filled measuring cylinder; a gas syringe is an allowed alternative. Keep ribbon mass/length, oxide removal, acid volume, starting temperature and apparatus the same. Add magnesium, fit the bung promptly and start timing consistently. Record volume, for example every 10 s, until constant. Repeat and plot both curves on the same axes.

Check for leaks and an unobstructed gas path; never seal a gas-producing reaction without an outlet. Use eye protection, a stable collector and no ignition source near hydrogen. A delayed bung loses early gas. Water collection needs suitable gas solubility and consistent readings; a syringe can stick or leak. Equal final volumes require equal limiting magnesium and enough acid in both trials.

Colour/turbidity method. Turbidity is cloudiness from suspended particles. The reviewed activity uses 40 g/dm^3 sodium thiosulfate stock. Mix stock and water to a constant 50 cm^3, then add the same 10 cm^3 dilute hydrochloric acid. Stock/water volumes 10/40, 20/30, 30/20, 40/10 and 50/0 cm^3 give pre-acid concentrations 8, 16, 24, 32 and 40 g/dm^3. These labels describe the dilution before adding acid, not the final 60 cm^3 mixture.

Place the same flask on a black cross, add acid, swirl consistently and time until the cross disappears. Sulfur makes the mixture cloudy. Keep temperature, total volume, flask/viewing depth, acid amount, lighting and endpoint judgement consistent. Repeat each dilution and compare mean times. At a comparable threshold, 1/mean time is a relative-rate proxy: 60 and 30 s give 0.0167 and 0.0333 per second, so the second proxy is twice as large. This is not a measured mass rate or proof that the whole reaction has finished.

Sulfur dioxide is released and can aggravate breathing difficulties. Use ventilation, eye protection and teacher-selected controls; do not lean over the flask or inhale fumes. Record raw repeats and justify any anomaly exclusion. Averaging cannot remove systematic viewing differences. Reducing stock volume without replacing water changes viewing depth as well as concentration.

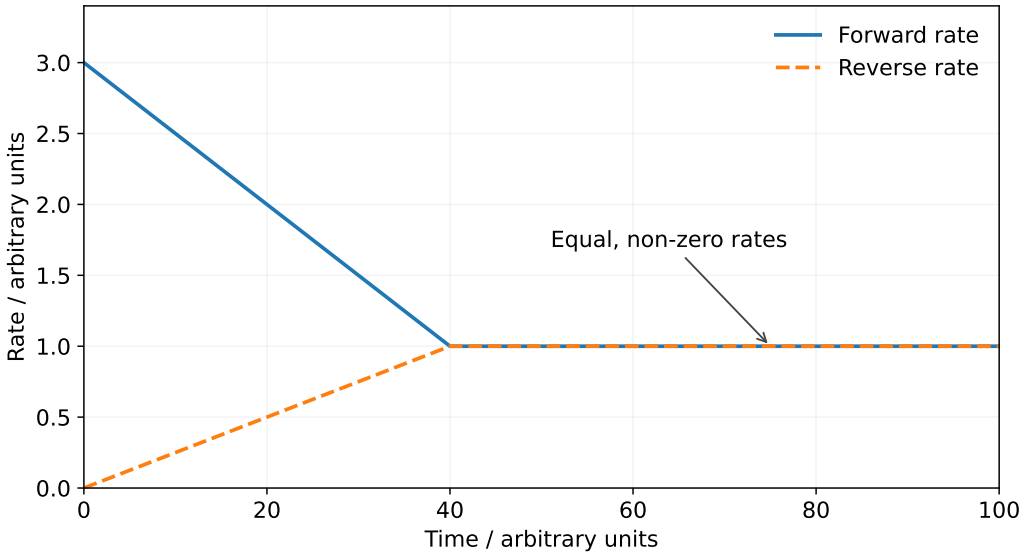
Reverse a reaction and its energy transfer (4.6.2.1–4.6.2.2)

In a **reversible reaction**, products can form the original reactants. Use $\rightleftharpoons$; changing conditions can change the direction favoured. Heating ammonium chloride forms ammonia and hydrogen chloride; cooling allows the gases to form ammonium chloride again. Preserve the substances in both directions.

Hydrated copper sulfate is blue. Heating removes water to form white anhydrous copper sulfate; adding water reverses this change. Dehydration is endothermic and hydration exothermic. For the same amount, energy transfers have equal magnitude and opposite directions. If dehydration takes in 72 kJ, reverse hydration releases 72 kJ for the corresponding amount; it does not release 144 kJ.

Dynamic equilibrium (4.6.2.3)

A **closed system** prevents reactants/products escaping. **Dynamic equilibrium** means forward and reverse reactions continue at equal rates. Concentrations remain constant but need not be equal. A sealed vessel with constant 0.60 mol reactant and 0.20 mol product can be at equilibrium if the rates are equal.



Closing a vessel allows equilibrium but does not establish it instantly. Starting with reactants, forward rate can fall while reverse rate rises as product forms. At equilibrium particles still react both ways; a flat concentration trace does not mean reaction stopped. An escaping product prevents this closed-system description from applying.

Higher Tier: predict a changed equilibrium (4.6.2.4–4.6.2.7)

Equilibrium position describes relative reactant/product amounts. **Le Chatelier principle** predicts a response opposing a condition change until a new equilibrium is reached. Use qualitative predictions here; no equilibrium-constant calculation is required.

Concentration. Adding a reactant immediately raises its concentration. Subsequent net forward reaction uses some added reactant and forms more product. Removing a reactant favours reverse; adding product favours reverse; removing product favours forward. Concentrations change until equilibrium returns. The disturbed concentration need not return exactly to its original value; rates are unequal during adjustment.

Temperature. Heating favours the endothermic direction; cooling favours exothermic. With an exothermic forward reaction, heating reduces equilibrium product amount and cooling increases it. With an endothermic forward reaction, heating increases product amount and cooling reduces it. Read the given forward energy change. Faster reaction after heating does not decide which equilibrium mixture contains more product.

Gas pressure. Compression at fixed temperature favours fewer gas molecules; expansion favours more. Count gaseous coefficients, not chemical names or solids. In $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, four gaseous molecules become two, so compression favours ammonia. In $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, both sides have two, giving no predicted position shift. This rule concerns changing volume, not adding a reactant and changing composition.

A suitable catalyst speeds both directions and reaches equilibrium sooner, without changing final position at the same conditions. In an industrial exothermic process, low temperature may favour yield but make production slow. A moderate temperature can balance rate, yield, energy use and cost. Pressure choices also need equipment/safety costs. Use supplied data: maximum yield alone is not an economic judgement.

Check yourself

- Choose quantity and interval; draw/interpret a tangent. HT calculates tangent and molar rates.
- Explain all five factors; distinguish faster reaction from final amount.
- Plan both RP5 methods with hypotheses, controls, repeats, observation limits and hazards.
- Explain reversibility, opposite energy transfers and equal continuing rates in a closed system.
- HT separates disturbance from adjustment, counts gas coefficients and predicts heating/cooling for either energy direction.