Rates of reaction

Science, grade 11 - by Claude Opus 5.5 - status: written (unreviewed)

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Grade 11 Science — Rates of reaction: what makes chemistry faster or slower

> **Unreviewed AI draft.** Written by Claude Opus 5.5 (an AI agent) on 2026-11-04. No teacher has checked it yet. Read it fully, adapt it to your class and local rules, and test any practical work yourself before using it.

Back to all units

Big idea

Some reactions are over in a flash and others take years. The rate of a reaction is how quickly reactants are used up or products are made, and it can be measured. Particles only react when they collide with enough energy (the activation energy), so anything that makes energetic collisions more frequent speeds a reaction up: higher concentration, higher temperature, more surface area, or a catalyst that offers a lower-energy route. These ideas are tested here with measured data, then used to explain real choices in industry, cars, kitchens and factory safety.

Outcomes

By the end of the unit, students can:

1. Measure the rate of a reaction by following a change over time (mass lost, gas made, or time for a solution to go cloudy), plot the results and explain why the rate is fastest at the start.
2. Use the disappearing-cross method to test how concentration affects rate, and use 1/time as a measure of rate.
3. Test how temperature affects rate, and explain the result using collision theory and activation energy, including the rough rule that rate increases by a factor of about 2 to 3 for every 10 °C rise.
4. Compare the rate for lumps and powder of the same mass, and explain the difference using surface area.
5. Show that a catalyst speeds up a reaction without being used up, including an enzyme (catalase) and what happens when the enzyme is boiled.
6. Use rate ideas to explain real examples: catalytic converters, the iron catalyst in the Haber process, and dust explosions.

Prior knowledge

- Particles in solids, liquids and gases; dissolving; mass is conserved (grades 5–8).
- Acids react with carbonates to make carbon dioxide; neutralisation; safe handling of acids (grade 8).
- Enzymes and respiration in living things (grade 7).
- Fair tests, repeats, medians, line graphs and gradients (grades 3–9; grade 9 trolley experiment).
- Word and symbol equations; moles (grade 10 chemistry, if taught; the mole calculation in lesson 1 is a stretch task).

Lessons (about 50 minutes each)

#	Lesson	Practical?
1	Measuring rate: marble chips and acid	Mass lost every 30 s on a balance; graph
2	Concentration: the disappearing cross	Sodium thiosulfate + dilute acid at four concentrations
3	Temperature, collisions and activation energy	Disappearing cross at 20, 30, 40 and 50 °C
4	Surface area and catalysts	Chips vs powder; hydrogen peroxide with manganese dioxide and potato
5	Rates in the real world + unit check	Case studies; written assessment

Materials (whole unit)

- Eye protection for everyone in every practical lesson.
- Lesson 1: top-pan balances (reading to 0.01 g); 100 ml conical flasks; cotton wool; marble chips (calcium carbonate), about 10 g per group; dilute hydrochloric acid, 1 mol/L; 50 ml measuring cylinders; stopwatches or phones.
- Lessons 2 and 3: sodium thiosulfate solutions at 40, 30, 20 and 10 g/L (made up by a technician); dilute hydrochloric acid, 1 mol/L; 100 ml conical flasks; paper with a bold black cross; 10 ml and 50 ml measuring cylinders; thermometers; a water bath or kettle-filled beakers (lesson 3, up to 50 °C only); stopwatches; a well-ventilated room.
- Lesson 4: large marble chips, small marble chips and calcium carbonate powder (10 g each per group); dilute hydrogen peroxide, about 3% (no stronger); manganese(IV) oxide (manganese dioxide) powder, about 0.5 g per group, handled with a spatula; raw potato, cut into 1 cm cubes and grated; one batch of boiled potato; gas syringes, or upturned measuring cylinders in water troughs; filter paper and funnels for recovering the catalyst.
- Lesson 5: case-study cards (given in lesson 5); student copies of the unit check.
- Student record sheets (tables are given in each lesson).

Assessment

- Ongoing: closing check questions at the end of each lesson (answers included).
- Lesson 1: the mass-loss table, graph and rate calculations are a practical skills task.
- Lesson 3: the conclusion explaining the temperature result with collision theory is a written-explanation task.
- Lesson 5: a 10-question unit check (20 marks) with an answer guide and marking notes. No pass mark is set; use it to decide what to reteach.

Sources

These are general references that a teacher can use to check facts. Links were checked to load on 2026-11-04.

- Wikipedia, "Reaction rate": https://en.wikipedia.org/wiki/Reaction_rate
- Wikipedia, "Collision theory": https://en.wikipedia.org/wiki/Collision_theory
- Wikipedia, "Activation energy": https://en.wikipedia.org/wiki/Activation_energy
- Wikipedia, "Arrhenius equation": https://en.wikipedia.org/wiki/Arrhenius_equation
- Wikipedia, "Catalysis": https://en.wikipedia.org/wiki/Catalysis
- Wikipedia, "Enzyme": https://en.wikipedia.org/wiki/Enzyme

- Wikipedia, "Catalase": https://en.wikipedia.org/wiki/Catalase
- Wikipedia, "Hydrogen peroxide": https://en.wikipedia.org/wiki/Hydrogen_peroxide
- Wikipedia, "Manganese dioxide": https://en.wikipedia.org/wiki/Manganese_dioxide
- Wikipedia, "Sodium thiosulfate": https://en.wikipedia.org/wiki/Sodium_thiosulfate
- Wikipedia, "Calcium carbonate": https://en.wikipedia.org/wiki/Calcium_carbonate
- Wikipedia, "Hydrochloric acid": https://en.wikipedia.org/wiki/Hydrochloric_acid
- Wikipedia, "Iodine clock reaction": https://en.wikipedia.org/wiki/Iodine_clock_reaction
- Wikipedia, "Catalytic converter": https://en.wikipedia.org/wiki/Catalytic_converter
- Wikipedia, "Haber process": https://en.wikipedia.org/wiki/Haber_process
- Wikipedia, "Dust explosion": https://en.wikipedia.org/wiki/Dust_explosion

Facts used from these pages include: the factors that affect rate (Reaction rate); that only collisions with the right energy and orientation succeed, and that higher concentration gives more collisions (Collision theory); that rate rises by a factor of about 2 to 3 for every 10 °C rise for common activation energies (Arrhenius equation), and that the ratio of rates 10 °C apart is usually on the order of 2 (Reaction rate); that catalysts are not consumed and remain unchanged, with manganese dioxide speeding up hydrogen peroxide decomposition (Catalysis); 2 H2O2 → 2 H2O + O2 and that even 3% hydrogen peroxide can be dangerous for the eye (Hydrogen peroxide); that one catalase molecule can convert millions of hydrogen peroxide molecules each second, and that human catalase works best near pH 7 (Catalase); that enzymes denature when heated (Enzyme); that dilute acid decomposes thiosulfate to sulfur, sulfur dioxide and water (Sodium thiosulfate); CaCO3 + 2 H+ → Ca2+ + CO2 + H2O (Calcium carbonate); that hydrochloric acid is corrosive (Hydrochloric acid); oxidation of carbon monoxide and unburned hydrocarbons in catalytic converters (Catalytic converter); the iron catalyst (Haber process); and the flour-dust explosion at the Washburn A Mill, Minnesota, on 2 May 1878 (Dust explosion).

Known gaps and cautions

- **All results tables are invented example data**, chosen to be realistic and internally consistent. Real classes will get different numbers; the point is the pattern and the reasoning.
- The disappearing-cross reaction gives off a little sulfur dioxide. Use small volumes, a well-ventilated room and prompt disposal, and follow your local chemical-safety guidance and technician's advice, especially for students with asthma.
- Using 1/time as "rate" is an approximation: it treats every run as reaching the same amount of cloudiness. Rate equations, orders of reaction and the full Arrhenius equation are left for later study.
- Collision theory is taught qualitatively. The energy-distribution graph in lesson 3 is a sketch, not a calculated curve.
- Equilibrium and Le Chatelier's principle (why the Haber process uses a compromise temperature) are mentioned only briefly in lesson 5.
- This unit has not been checked by a teacher or a chemistry technician. Concentrations and quantities should be checked against your school's practical guidance before use.

Lesson 1 — Measuring rate: marble chips and acid

Back to unit

Goal

Students follow one reaction over time by measuring the mass lost as carbon dioxide escapes, plot a graph, and explain why the reaction is fastest at the start and then slows to a stop.

Materials

- Per group: a top-pan balance reading to 0.01 g; a 100 ml conical flask; a loose plug of cotton wool; about 10 g of marble chips (calcium carbonate), more than enough to react with all the acid; 50 ml of dilute hydrochloric acid, 1 mol/L; a 50 ml measuring cylinder; a stopwatch.
- Eye protection for everyone. Paper towels for drips.

Steps

✎ Opening (5 min)

Ask: "Which is faster, a firework or rusting? How could you measure how fast a reaction is, not just say it?" Collect ideas. Steer towards: watch something change over time — a reactant being used up or a product being made.

Introduce the word **rate**: how much reactant is used up, or product made, per second.

✎ Main activity (30 min)

1. Write the reaction on the board: calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide. Symbol form for the reacting ions: CaCO3 + 2 H+ → Ca2+ + CO2 + H2O.
2. Ask: "If carbon dioxide gas escapes from an open flask, what happens to the reading on the balance?" (It goes down.)
3. Each group puts 50 ml of acid in the flask, places the flask and the cotton-wool plug together on the balance and zeroes it (or notes the reading).
4. Add the marble chips, put the cotton wool loosely in the neck (it lets gas out but stops acid spray escaping), and start the stopwatch.
5. Record the mass lost every 30 seconds for 5 minutes.

Example results (invented, but realistic):

Time (s)	0	30	60	90	120	150	180	210	240	270	300
Mass lost (g)	0.00	0.30	0.54	0.72	0.85	0.94	1.00	1.04	1.07	1.09	1.10

1. Plot mass lost (y-axis) against time (x-axis) and draw a smooth curve.

2. Work out the mean rate in three time windows:
3. 0–30 s: 0.30 g ÷ 30 s = **0.010 g/s**
4. 120–150 s: (0.94 − 0.85) g ÷ 30 s = 0.09 ÷ 30 = **0.003 g/s**
5. 270–300 s: (1.10 − 1.09) g ÷ 30 s = 0.01 ÷ 30 ≈ **0.0003 g/s**
6. Stretch (moles): 50 ml of 1 mol/L acid contains 0.050 mol of H+. The equation needs 2 H+ for every CO_2, so at most 0.025 mol of CO_2 can form. CO_2 has a molar mass of 44 g/mol, so the most mass that can be lost is 0.025 × 44 = **1.10 g**. The reaction stopped at 1.10 g because the acid ran out (the marble chips were in excess).

✓ Discussion questions (10 min)

1. **Where is the graph steepest, and what does that mean?** Example answer: at the start. The steeper the curve, the more mass is lost each second, so the reaction is fastest at the start.
2. **Why does the reaction slow down?** Example answer: the acid is being used up, so there are fewer acid particles to hit the marble, and fewer reactions happen each second.
3. **Why did it stop at 1.10 g, even though there was still marble left?** Example answer: all the acid had reacted; the acid was the limiting reactant.
4. **Why is the cotton wool loose, not a tight bung?** Example answer: the gas must escape for the mass to drop, and a sealed flask could build up pressure.

✓ Closing check (5 min)

1. A curve on a mass-lost graph becomes flat at 4 minutes. What does that tell you? Answer: the reaction has stopped (one reactant has run out).
2. A reaction makes 24 cm^3 of gas in the first 20 seconds. What is the mean rate? Answer: 24 ÷ 20 = 1.2 cm^3/s.
3. Give one other way to follow this reaction, apart from mass. Answer: collect the gas in a gas syringe or upturned measuring cylinder and measure its volume over time.

Vocabulary

rate of reaction, reactant, product, limiting reactant, in excess, mean rate, gradient

Differentiation

- **Support:** give a pre-drawn axes grid and the first three points plotted; students compare "steep" and "flat" parts in words.
- **Stretch:** draw a tangent at 60 s and estimate the rate at that instant from its gradient; do the stretch mole calculation at the end of the main activity.

⚠ Safety

- Eye protection throughout. Dilute hydrochloric acid at 1 mol/L is an irritant; stronger hydrochloric acid is corrosive. Wipe up spills and wash skin with water.
- Keep the flask on the balance pan and do not lean over the neck while it fizzes.

- Pour the used liquid away as directed by the technician; leftover marble chips can be rinsed and reused.

Teacher notes and misconceptions

- "The reaction stops because the marble runs out." Here it does not; the acid runs out. Show the leftover chips.
- "Rate means how much product you get in total." Rate is how fast; the total amount depends on how much reactant there was. Two reactions can make the same total at different rates (lesson 4 shows this).
- The balance reading may drift slightly while the flask is still warm or wet; tell students to read it at the moment each 30 s ends, not to wait for it to settle.

Lesson 2 — Concentration: the disappearing cross

Back to unit

Goal

Students carry out a fair test of how the concentration of sodium thiosulfate affects the rate of its reaction with dilute acid, use 1/time as a measure of rate, and explain the pattern with collision theory.

Materials

- Per group: a 100 ml conical flask; a sheet of white paper with a bold black cross drawn on it; sodium thiosulfate solutions at 40, 30, 20 and 10 g/L (50 ml of each per run, made up by a technician); dilute hydrochloric acid, 1 mol/L; a 50 ml and a 10 ml measuring cylinder; a stopwatch.
- Eye protection for everyone. A well-ventilated room.

Steps

✎ Opening (5 min)

Show the reaction once as a demonstration. Put the flask of thiosulfate on the cross, add the acid, swirl once and start the stopwatch. The liquid slowly turns cloudy and yellow-white until the cross can no longer be seen from above.

Explain: dilute acid makes thiosulfate break down into sulfur, sulfur dioxide and water. The sulfur forms as tiny solid specks, which is what makes the liquid cloudy. The time until the cross disappears is a measure of how fast the sulfur forms.

✎ Main activity (30 min)

1. Ask for a prediction: "If the thiosulfate is more concentrated, will the cross disappear sooner or later? Why?"
2. Agree what to keep the same: volume of thiosulfate (50 ml), volume and concentration of acid (5 ml of 1 mol/L), the same flask, the same cross, the same person looking down from the same height, room temperature.
3. For each concentration: put 50 ml of thiosulfate in the flask on the cross, add 5 ml of acid, swirl once, start timing, and stop when the cross is no longer visible. Do three runs and take the median.
4. Rinse the flask between runs straight away (see Safety).

Example results (invented, but realistic):

Thiosulfate (g/L)	Run 1 (s)	Run 2 (s)	Run 3 (s)	Median time (s)	Rate = 1000 ÷ time (1/s × 1000)
40	21	22	20	21	47.6
30	28	27	30	28	35.7
20	43	45	42	43	23.3
10	86	84	112	86	11.6

1. Plot rate (1000 ÷ time) against concentration. The points lie close to a straight line through the origin: rate ÷ concentration is 1.19, 1.19, 1.16 and 1.16.
2. Discuss the 112 s run at 10 g/L: it is far from the other two, so it is probably a mistake (for example the timer was stopped late or the person looked from a different angle). Using the median means one odd run does not drag the answer.

✓ Discussion questions (10 min)

1. **Why do we use 1 ÷ time as the rate?** Example answer: a shorter time means a faster reaction, so 1 ÷ time gets bigger as the reaction gets faster. Each run makes about the same amount of cloudiness before the cross vanishes.
2. **What does the graph show?** Example answer: doubling the concentration roughly doubles the rate (20 g/L gives 23.3, 40 g/L gives 47.6).
3. **Explain the result using particles.** Example answer: a more concentrated solution has more thiosulfate particles in the same volume, so they collide with acid particles more often, and there are more successful collisions each second.
4. **What is the weakest part of this method?** Example answer: deciding exactly when the cross disappears is a judgement, and different people see it at different times. Keeping the same observer helps.

✓ Closing check (5 min)

1. At 25 g/L, between which two measured times would you expect the cross to disappear? Answer: between 28 s (30 g/L) and 43 s (20 g/L).
2. Why must the volume of acid be kept the same? Answer: it is a control variable; changing it would also change how fast the reaction goes, so the test would not be fair.
3. Using collision theory, explain why the 10 g/L solution took longest. Answer: fewest thiosulfate particles per ml, so the fewest collisions with acid particles each second.

Vocabulary

concentration, independent variable, dependent variable, control variable, median, collision theory, successful collision

Differentiation

- **Support:** give the results table with medians filled in; students plot time (not 1 ÷ time) against concentration and describe the curve in words.
- **Stretch:** compare the line through the origin with the iodine clock reaction, another timed reaction in which more thiosulfate makes the blue colour appear later. Ask why.

⚠ Safety

- Eye protection throughout. Dilute hydrochloric acid is an irritant.
- The reaction gives off a little sulfur dioxide, which has a sharp smell and irritates the airways. Use only the small volumes given, keep the room well ventilated, do not lean over the flask to sniff it, and pour each finished mixture away promptly with plenty of water, or as your technician directs.
- Students with asthma may be more sensitive. Seat them near an open window or let them be the timer and recorder.
- Follow your local chemical-safety guidance and technician's advice.

Teacher notes and misconceptions

- "A faster reaction makes more product." The cross disappears at roughly the same amount of sulfur each time; the concentrated run just gets there sooner.
- "More concentrated means the particles move faster." The particles move at the same speed at the same temperature; there are just more of them in each ml, so collisions are more frequent.
- The sulfur sticks to glass as a yellow film; rinse flasks between runs or the cross starts to look faint before the reaction does anything.

Lesson 3 — Temperature, collisions and activation energy

Back to unit

Goal

Students test how temperature affects the rate of the thiosulfate reaction and explain the result using collision theory and activation energy, including why a 10 °C rise has a much bigger effect than students usually expect.

Materials

- As lesson 2, using only the 20 g/L sodium thiosulfate solution.
- Thermometers; a water bath, or beakers of warm water from a kettle mixed with cold water, set to about 30, 40 and 50 °C. No hotter than 50 °C.
- Eye protection for everyone. A well-ventilated room.

Steps

✎ Opening (5 min)

Ask: "Why do we keep milk in the fridge?" Collect ideas, then ask: "Is there something general about reactions going slower when it is cold?"

Recap collision theory from lesson 2: particles must collide to react. Then add the new idea: **not every collision works.** Particles must hit hard enough, with at least a minimum amount of energy called the **activation energy**, and in a suitable orientation. Collisions that lead to a reaction are called successful collisions.

✎ Main activity (30 min)

1. Predict: "If we warm the thiosulfate from 20 °C to 30 °C, will the time halve, change a little, or stay the same?"
2. For each temperature: warm 50 ml of the 20 g/L thiosulfate in the flask in the water bath, measure its temperature, stand it on the cross, add 5 ml of 1 mol/L acid (at room temperature), swirl once and time until the cross disappears. Measure the temperature again at the end and record the mean of the start and end temperatures.
3. The 20 °C run is the same as the 20 g/L result from lesson 2; it can be reused if the room was at about 20 °C.

Example results (invented, but realistic; median of three runs each):

Mean temperature (°C)	Median time (s)	Rate = 1000 ÷ time
20	43	23.3
30	22	45.5

Mean temperature (°C)	Median time (s)	Rate = 1000 ÷ time
40	12	83.3
50	7	142.9

1. Work out how the rate changes for each 10 °C step: 45.5 ÷ 23.3 ≈ 1.95; 83.3 ÷ 45.5 ≈ 1.83; 142.9 ÷ 83.3 ≈ 1.72. Each 10 °C rise roughly doubles the rate.
2. Compare with chemistry references: for many reactions the rate rises by a factor of about 2 to 3 for every 10 °C rise, and the ratio of rates 10 °C apart is usually on the order of 2. The class result fits that rough rule.
3. Note the limit of the data: 7 s is hard to time by hand, so the 50 °C result is the least reliable.

Explaining the result (sketch)

Draw a sketch (not to scale) of the spread of particle energies at a low and a higher temperature, with a vertical line marking the activation energy. At the higher temperature the curve is flatter and spreads to higher energies, so a bigger share of particles lies to the right of the line.

Two things happen when the temperature rises:

1. Particles move faster, so they collide a little more often.
2. More importantly, a much bigger fraction of collisions have at least the activation energy, so a much bigger fraction are successful.

The second effect is why a 10 °C rise can double the rate, even though the particles only move a little faster.

✓ Discussion questions (10 min)

1. **Why does the rate roughly double when the temperature only goes up a little in kelvin (from 293 K to 303 K)?** Example answer: the number of collisions only rises a little, but the share of collisions with enough energy to react rises a lot.
2. **Why do we record the mean of the start and end temperatures?** Example answer: the warm flask cools while we time it, and the cold acid cools it too, so the start temperature alone would be too high.
3. **Why is the 50 °C result the least trustworthy?** Example answer: 7 s is very short, so a reaction-time error of half a second is a large fraction of the result.
4. **Why do fridges slow down food going off?** Example answer: lower temperature means fewer collisions with enough energy, so the reactions that spoil food (including those in microbes) go more slowly.

✓ Closing check (5 min)

1. What is activation energy? Answer: the minimum energy colliding particles need for the collision to lead to a reaction.
2. A reaction takes 80 s at 25 °C. Using the "about 2 per 10 °C" rule, estimate the time at 35 °C. Answer: about 40 s.

3. Give two reasons a higher temperature increases the rate. Answer: particles collide more often, and a bigger fraction of collisions have at least the activation energy.

Vocabulary

activation energy, successful collision, energy distribution, kelvin, rule of thumb

Differentiation

- **Support:** give the sketch of the energy distribution already drawn; students shade the part above the activation energy for each temperature and compare the areas.
- **Stretch:** plot rate against temperature and describe the curve (it gets steeper). Ask whether the rule "about 2 per 10 °C" can hold for ever (no: at high temperatures other limits appear, and in living things enzymes denature, as lesson 4 shows).

⚠ Safety

- Everything in lesson 2 applies, including ventilation and prompt disposal because of sulfur dioxide. Warm solutions give off sulfur dioxide faster, so the 50 °C limit matters.
- Water no hotter than 50 °C. Teachers handle any kettle; students only use the water bath or pre-mixed warm water.
- Eye protection throughout.

Teacher notes and misconceptions

- "Heat gives the reaction energy, so the products have more energy." The point is collision energy, not that heat is stored in the products.
- "Higher temperature means more particles." The number of particles is the same; they move faster and more collisions are energetic enough.
- "Rate doubles every 10 °C exactly." It is a rough rule for many common reactions, not a law; real factors vary.

Lesson 4 — Surface area and catalysts

Back to unit

Goal

Students show that breaking a solid into smaller pieces speeds up its reaction without changing how much product it makes, and that a catalyst (manganese dioxide, or the enzyme catalase in potato) speeds up a reaction without being used up.

Materials

- **Part A, per group:** 10 g each of large marble chips (the same size as lesson 1), small marble chips and calcium carbonate powder; 3 × 50 ml of 1 mol/L hydrochloric acid; a 100 ml conical flask for the chips and a 250 ml conical flask for the powder (it froths); cotton wool; a balance; a stopwatch.
- **Part B, per group:** dilute hydrogen peroxide, about 3%, 25 ml per run; a gas syringe (or an upturned measuring cylinder in a trough of water) connected to a conical flask with a bung and delivery tube; about 0.5 g of manganese dioxide powder on a weighed filter paper; 5 g of raw potato in 1 cm cubes; 5 g of grated raw potato; 5 g of grated potato that has been boiled for 5 minutes and cooled (prepared by the technician); a funnel, filter paper and a warm place to dry the paper.
- Eye protection for everyone.

Steps

✎ Opening (5 min)

Hold up a sugar cube and a spoonful of granulated sugar. Ask: "Which dissolves faster in tea, and why?" (Grade 5 dissolving.) Then: "Would the same idea work for a reaction?"

Show a sheet of paper and the same sheet torn into eight pieces. Ask how much more edge and surface there is to touch. Introduce **surface area**: particles can only react where the solid touches the liquid, so more surface means more collisions each second.

✎ Main activity, part A: surface area (12 min)

Repeat the lesson 1 method with each size, recording the mass lost at 60 s and at the end.

Example results (invented, but realistic):

Solid (10 g)	Mass lost at 60 s (g)	Mass lost at end (g)
Large chips	0.54	1.10
Small chips	0.86	1.10
Powder	1.08	1.10

All three end at 1.10 g, because in every case the same 50 ml of acid is used up (lesson 1). Smaller pieces change how fast, not how much.

✎ Main activity, part B: catalysts (18 min)

Hydrogen peroxide slowly breaks down into water and oxygen: 2 H2O2 → 2 H2O + O2. On its own this is very slow. For each run, put 25 ml of peroxide in the flask, add the solid, put the bung in at once and record the volume of oxygen collected after 60 s.

Example results (invented, but realistic):

Added to 25 ml peroxide	Oxygen after 60 s (cm³)
Nothing	1
0.5 g manganese dioxide	52
5 g potato cubes	9
5 g grated potato	24
5 g grated boiled potato	0

Then filter off the manganese dioxide, rinse it, let the paper dry and weigh it again. Example: 0.50 g before, 0.49 g after; the small loss is powder stuck to the glass and paper.

✓ Discussion questions (10 min)

1. **Why did the powder react fastest even though it was the same mass?** Example answer: it has much more surface touching the acid, so more acid particles hit the solid each second.
2. **How do we know manganese dioxide is a catalyst and not a reactant?** Example answer: it speeds the reaction up a lot, but almost all of it is still there afterwards, unchanged. Catalysts are not used up and can be reused.
3. **Why does potato work at all?** Example answer: potato cells contain the enzyme catalase, a biological catalyst that breaks down hydrogen peroxide. One catalase molecule can break down millions of peroxide molecules each second.
4. **Why did grated potato beat cubes, and boiled potato do nothing?** Example answer: grating exposes more cells, so more enzyme meets the peroxide (surface area again). Boiling denatures the enzyme: its shape unfolds and it stops working.
5. **How does a catalyst work, in collision-theory terms?** Example answer: it gives the reaction a different route with a lower activation energy, so a bigger fraction of collisions are successful at the same temperature.

✓ Closing check (5 min)

1. 10 g of lumps and 10 g of powder react with the same acid until the acid runs out. Which makes more gas in total? Answer: neither; they make the same total, but the powder makes it faster.
2. Give two pieces of evidence that a substance is a catalyst. Answer: the reaction is much faster with it; its mass is (almost) unchanged at the end.

3. Why are enzymes in your body sometimes said to work best at "about pH 7 and body temperature"? Answer: outside their best conditions enzymes can lose their shape (denature) and work less well. For example, human catalase works best near pH 7.

Vocabulary

surface area, catalyst, enzyme, catalase, denature, alternative route, decomposition

Differentiation

- **Support:** give part B as a teacher demonstration with the gas syringe projected; students fill in the table and answer questions 2 and 4.
- **Stretch:** 25 ml of 3% hydrogen peroxide can give about 250 cm³ of oxygen in total (around ten times its own volume). Ask: what percentage had broken down after 60 s with manganese dioxide? (52 ÷ 250 ≈ 21%.)

⚠ Safety

- Eye protection throughout. Even 3% hydrogen peroxide can be dangerous in the eye, and it irritates skin. Do not use stronger peroxide for this lesson. Wash splashes off with plenty of water.
- Never seal the flask tightly without a delivery tube to a gas syringe or trough; oxygen must have somewhere to go.
- Handle manganese dioxide powder with a spatula, avoid making dust, and dispose of it as your technician directs.
- No large "elephant toothpaste" demonstrations; they use much stronger peroxide.
- Do not eat any potato used in the lesson.
- Dilute hydrochloric acid is an irritant (part A). The powder froths up quickly; use the larger flask.

Teacher notes and misconceptions

- "A catalyst makes more product." It only changes how fast; the total amount of oxygen is set by how much peroxide there is.
- "Catalysts are not involved in the reaction." They take part, but come out unchanged at the end.
- "Boiling kills the enzyme." Enzymes are not alive; boiling changes their shape so they no longer work (denaturing).
- The "nothing" run may give 0–2 cm³ just from the gas warming or the bung pushing air in; that is a useful point about measurement error.

Lesson 5 — Rates in the real world + unit check

Back to unit

Goal

Students use concentration, temperature, surface area and catalysts to explain real situations where people want reactions faster or slower, then complete the unit check.

Materials

- Four case-study cards (below), one set per group.
- Student copies of the unit check (10 questions, 20 marks).

Steps

✎ Opening (5 min)

Ask: "Name a reaction people want to speed up, and one they want to slow down." Examples: making ammonia for fertiliser or cleaning car exhaust (faster); food going off or a dust cloud catching fire (slower, or not at all).

✎ Main activity: case-study cards (20 min)

Groups read each card and answer: Which factor or factors are involved, and how would you explain them using collisions?

1. **Catalytic converters.** A car exhaust contains carbon monoxide and unburned fuel. In a catalytic converter, a ceramic honeycomb coated with metal catalysts (such as platinum and palladium) helps these react with oxygen to form carbon dioxide and water. "Three-way" converters also reduce oxides of nitrogen; Volvo first sold them on 1977 California-specification 240 cars. Factors: catalyst (lower activation energy route), large surface area (honeycomb), hot exhaust (temperature).
2. **The Haber process.** Ammonia for fertiliser is made by reacting nitrogen and hydrogen on the surface of a solid catalyst, usually iron. Many plants run at about 500 °C and 200 bar. The iron catalyst needs at least about 400 °C to be efficient. Factors: catalyst; temperature; high pressure (gas particles closer together, like higher concentration). Teacher note: lower temperatures would give more ammonia at equilibrium but too slowly, so the temperature is a compromise. Equilibrium is studied later.
3. **Dust explosions.** Flour, sugar, grain, cocoa and coffee can all burn explosively as a fine dust cloud. On 2 May 1878 a flour-dust explosion destroyed the Washburn A Mill in Minnesota, killing 14 workers there and four more in nearby buildings. Factor: surface area. A bag of flour burns slowly; the same flour spread through the air as dust has a huge surface touching oxygen, so it can burn almost all at once.
4. **Keeping food cool.** Fridges and freezers slow down the reactions that make food go off, including those in microbes. Factor: temperature. Fewer collisions with at least the

activation energy; using the "about 2 per 10 °C" rule, food at 5 °C spoils several times more slowly than at 25 °C, as a rough estimate.

Share answers as a class. Draw out that the same four factors explain every case.

✓ Unit check (25 min)

Students work alone. Mark out of 20. No pass mark is set; use the results to decide what to reteach.

1. What is meant by the "rate" of a reaction? Give one way to measure it. **(2 marks)**
2. A reaction gives these gas volumes:

Time (s)	**0**	**10**	**20**	**30**	**40**	**50**	**60**
Gas (cm^3)	0	30	50	62	68	70	70

(a) Calculate the mean rate in the first 10 s. (b) When did the reaction stop? **(2 marks)** 3. In a disappearing-cross test, three runs at one concentration took 18 s, 19 s and 31 s. Which value would you report, and why? **(2 marks)** 4. Use collision theory to explain why a more concentrated acid reacts faster with magnesium ribbon. **(2 marks)** 5. What is activation energy? **(2 marks)** 6. A reaction takes 60 s at 20 °C. Using the rough rule that the rate about doubles for each 10 °C rise, estimate how long it takes at 40 °C. Show your working. **(2 marks)** 7. 10 g of marble lumps and 10 g of marble powder each react with 50 ml of the same acid, which runs out in both cases. Compare (a) how fast and (b) how much carbon dioxide each makes. **(2 marks)** 8. A student adds 1.00 g of a black powder to hydrogen peroxide. Bubbles of oxygen appear quickly. Afterwards she filters, dries and weighs the powder: 0.99 g. What does this suggest about the powder? Explain. **(2 marks)** 9. Raw grated potato makes hydrogen peroxide fizz, but boiled grated potato does not. Explain why. **(2 marks)** 10. A flour mill keeps its floors and machines free of dust. Use rates of reaction to explain why. **(2 marks)**

Answer guide and marking notes

1. How fast reactants are used up or products are formed (1); any valid method, e.g. mass lost over time, gas volume over time, time for a cross to disappear (1).
2. (a) 30 ÷ 10 = 3.0 cm^3/s (1). (b) At 50 s, when the volume stops rising (1). Accept "between 40 and 50 s" only if the student explains the curve is still rising at 40 s.
3. 19 s, the median (1), because 31 s is far from the others and is probably a mistake, so it should not drag the result (1). Accept "the mean of 18 and 19" with the same reason.
4. More acid particles in the same volume (1), so more frequent collisions with the magnesium and more successful collisions per second (1). "Particles move faster" scores 0 for the second mark.
5. The minimum energy (1) that particles must have when they collide for the collision to lead to a reaction (1).
6. 20 °C → 30 °C: rate doubles, time halves to 30 s (1); 30 °C → 40 °C: time halves again to 15 s (1). Accept "about 15 s" with working.
7. (a) Powder reacts faster because it has a bigger surface area (1). (b) The same amount, because the same amount of acid is used up in each (1).
8. It is a catalyst (1): it sped up the reaction but its mass was almost unchanged, so it was not used up; the 0.01 g difference is powder lost in filtering (1).

9. Potato contains the enzyme catalase, which catalyses the breakdown of hydrogen peroxide (1). Boiling denatures the enzyme, so its shape changes and it no longer works (1). "Boiling kills the enzyme" scores 1 only if "changes shape" or "denatures" is also given.
10. Fine dust has a very large surface area (1), so a dust cloud can burn extremely fast, causing an explosion (1). Accept any reference to the Washburn A Mill or similar as supporting detail, not as a replacement for the explanation.

Vocabulary

catalytic converter, Haber process, dust explosion, compromise conditions, pressure

Differentiation

- **Support:** give each case-study card with the factor names listed underneath; students choose and explain one.
- **Stretch:** for the Haber card, research what "equilibrium" means and why a lower temperature would give more ammonia but more slowly.

⚠ Safety

- No practical work in this lesson. Do not attempt dust-explosion demonstrations in class.

Teacher notes and misconceptions

- "Catalytic converters remove all pollution." They change some harmful gases into less harmful ones; carbon dioxide is still produced.
- "Higher pressure only matters for gases because they get hotter." In this unit, higher pressure is treated as packing gas particles closer together, like a higher concentration.
- Students often write "more collisions" without "per second". Insist on the rate of collisions, not just the number.