6.1 Rate of reaction

Name: _____ Class: _____ Date: _____ **Total: 49 marks**

KEY WORDS rate of reaction • collision theory • catalyst

Objective

Calculate rates, explain the five factors by collision theory, and describe catalysts. Original questions for spec statements 4.6.1.1–4.6.1.4.

1 Key knowledge and worked examples

Rate = quantity used or formed ÷ time – g/s or cm^3/s (HT: mol/s). On a product–time graph the curve is steepest at the start and flattens as the reactant runs out; draw a **tangent** and use its gradient as the rate at that time (HT).

Five factors: concentration ↑, pressure ↑ (gases), surface area ↑ (smaller pieces), temperature ↑, catalyst – all raise the rate.

Collision theory: react only on collision with **at least the activation energy**:

- concentration/pressure ↑ at fixed temperature: more particles per volume, more frequent collisions;
- surface area ↑ – more exposed particles, more frequent collisions;
- temperature ↑ – more collisions AND **more exceed the activation energy** (the big effect).

Catalyst – a different pathway with **lower activation energy**; **not used up overall**; absent from the overall equation; **enzymes** are biological catalysts.

Worked example (3 marks, both tiers). 48 cm^3 of hydrogen forms in 20 s from magnesium + acid. Give the mean rate. Magnesium remains during the measured interval: explain why the rate falls as acid is consumed.

Mean rate = 48 ÷ 20 = 2.4 cm^3/s; For a trial with magnesium remaining, the rate falls as acid is consumed; acid concentration falls, so successful collisions become less frequent.

2 Practice

2.1 0.48 g of magnesium disappears in 60 s. Give the mean rate in g/s. [2]

2.2 Why does powdered marble react faster than the same mass as one lump? [2]

2.3 What does a catalyst change on a reaction profile? [1]

3 Exam-style questions

3.1 Explain, using collision theory, why a reaction mixture kept at 60 °C reacts faster than the same mixture at 20 °C. [4]

3.2 In RP5, sodium thiosulfate and acid turn cloudy. Describe how to compare different thiosulfate concentrations using this change, and give two control variables. [4]

3.3 *(HT)* Sketch a product–time graph and explain how to find the rate at 30 s. [3]

3.4 Gas volume is 12 cm^3 at 10 s and 36 cm^3 at 30 s. Find the mean collection rate over this interval. Explain why 36 ÷ 30 does not answer this interval question. [3]

3.5 *(HT)* At 30 s a gas-volume curve is at 45 cm^3. Its tangent passes through (10 s, 25 cm^3) and (50 s, 65 cm^3). (a) Calculate the instantaneous rate and the mean rate from zero to 30 s. (b) A separate reactant amount falls from 0.018 to 0.006 mol in 40 s. Calculate its mean consumption rate in mol/s. [6]

3.6 Plan the gas-volume part of RP5 to test the hypothesis that higher acid concentration makes magnesium react faster. Include collection, timing, controlled quantities, repeats and how to compare results. Explain one way a lower final collected volume could occur without proving a slower reaction. [7]

3.7 In the cloudy-reaction method, 40 g/dm^3 thiosulfate stock is diluted with water to 50 cm^3 before the same 10 cm^3 acid is added. (a) Find the pre-acid concentration for 30 cm^3 stock plus 20 cm^3 water. (b) Endpoint times are 48, 50 and 52 s. Find the mean and reciprocal-time proxy; compare with a second dilution taking 25 s. (c) Explain why disappearing-cross time is not completion time, and identify one respiratory hazard/control. [7]

3.8 A reacting gas is compressed at fixed temperature. Separately, a solid cube of side 2 cm is cut into eight cubes of side 1 cm without losing material. Explain both rate changes by collisions. Calculate the surface-area-to-volume ratio before and after cutting. Does compression at fixed temperature give each gas particle more kinetic energy? [6]

3.9 Adding substance X changes the initial gas-formation rate from 0.40 to 1.20 cm^3/s. Both trials finish at 60 cm^3. X has mass 1.00 g before and after, and is absent from the overall equation. Explain the evidence for catalytic action, the change on an energy profile and one limit on concluding that X catalyses every reaction. [4]

Solutions

2.1 $0.48 \div 60 = 0.008$ g/s.

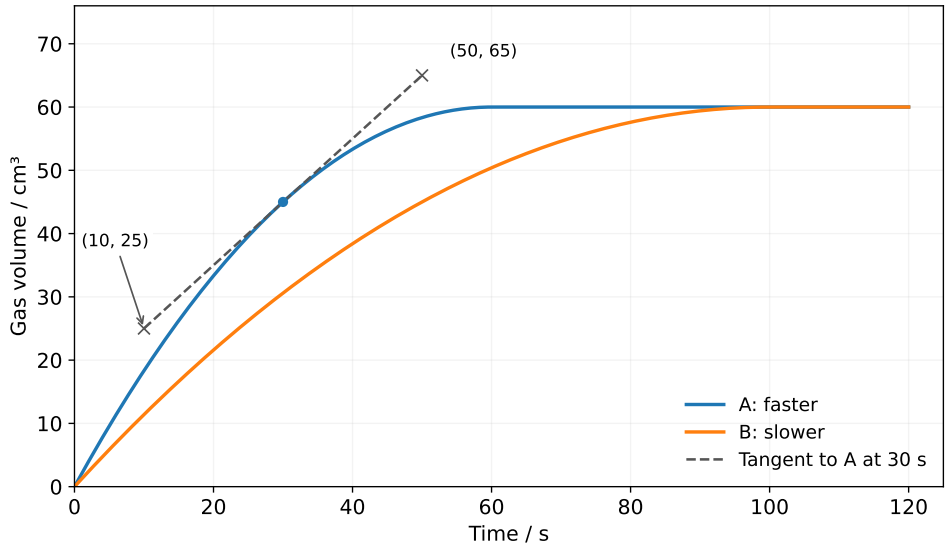
2.2 Powder has a larger surface-area-to-volume ratio; more particles are exposed to the acid, so collisions are more frequent.

2.3 It lowers the activation energy – a lower hump; the overall energy change is unchanged.

3.1 At 60 °C particles have more kinetic energy; they collide more frequently; and a greater proportion of collisions exceeds the activation energy; both effects raise the successful-collision rate, so the reaction is faster.

3.2 Time from adding acid until the same cross is no longer visible; a shorter time to the same threshold indicates a faster relative rate, not completion of reaction. Any two controls: acid volume/concentration, total mixture volume/viewing depth, temperature, flask or consistent lighting/observer. Thiosulfate concentration is varied and cannot be a control.

3.3



Curve rises then flattens; draw a tangent touching the curve at 30 s; the gradient of the tangent = change in quantity ÷ change in time along it, giving the rate at 30 s.

3.4 Volume change $= 36 - 12 = 24$ cm^3 and time interval $= 30 - 10 = 20$ s; rate $= 24/20 = 1.2$ cm^3/s. 36/30 uses the whole period from zero, not the specified 10–30 s interval.

3.5 (a) Tangent gradient $= (65 - 25)/(50 - 10) = 1.0$ cm^3/s. Mean from zero $= 45/30 = 1.5$ cm^3/s; mean over an interval need not equal the rate at its end. (b) Amount used $= 0.018 - 0.006 = 0.012$ mol; divide by 40 s $= 0.00030$ mol/s. The positive consumption amount is required; a chord is not a tangent.

3.6 Original guidance, any six method points: matched cleaned magnesium mass/length; fixed acid volume and starting temperature; two provided acid concentrations with an unobstructed collection path into a gas syringe or inverted water-filled measuring cylinder; fit bung/start timer consistently and promptly; record gas volumes at fixed intervals until constant; repeat and plot comparable volume–time curves, comparing initial slopes. Eye protection, clamping and no ignition sources near hydrogen; never seal without an outlet. Lower recorded final volume can result from a leak, delayed bung, unequal limiting magnesium or insufficient acid, so it alone cannot establish slower chemistry. Actual supervised practical remains separate.

3.7 (a) Pre-acid concentration $= 40 \times 30/50 = 24$ g/dm^3. This is not the final 60 cm^3 mixture concentration. (b) Mean $= (48 + 50 + 52)/3 = 50$ s; proxy $= 1/50 = 0.020$ per second; second $= 1/25 = 0.040$ per second, twice the proxy at the same visual threshold. (c) The cross disappears at a chosen cloudiness threshold while reaction may continue. Sulfur dioxide can aggravate breathing difficulties; use approved ventilation/teacher controls and do not inhale fumes. Repeats cannot fix a systematic viewing-depth difference.

3.8 Compression increases particles per volume and collision frequency; temperature and average kinetic energy are unchanged. Original cube area $= 6 \times 2^2 = 24$ cm^2 and volume $= 2^3 = 8$ cm^3, ratio 3 per cm. Eight small cubes have total area $8 \times 6 = 48$ cm^2 and the same volume 8 cm^3, ratio 6 per cm. More exposed solid surface permits more frequent reactant–surface collisions, increasing rate under matched conditions. Total mass is unchanged; no exact universal factor-of-two rate law follows from the geometry alone.

3.9 X increases the rate and is not used up overall, consistent with catalytic action; absence from the overall equation supports the identification together with those observations. The alternative pathway has lower activation energy, with unchanged reactant/product levels and overall energy change. Different reactions need suitable catalysts: this experiment alone cannot establish action in another reaction. The equal final gas volumes show faster production, not extra product from the same limiting amount. Mass alone would not prove chemical identity unchanged.