EXAM PAPERS PRACTICE

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Detailed mark scheme

Suitable for all boards

Designed to test your ability and thoroughly prepare you

Level: CIE AS and A Level (9701)
Subject: Chemistry
Topic: CIE Chemistry
Type: Mark Scheme

2002

XVIII

1583

Chemistry CIE AS & A Level
To be used for all exam preparation for 2025+

CHEMISTRY AS and A

This to be used by all students studying CIE AS and A level Chemistry (9701) But students of other boards may find it useful

Mark Scheme

Answer 1.

a) According to collision theory, the **two** conditions that particles need for an effective collision are:

- Sufficient energy
  OR
  Energy above the activation energy / $E > E_a$; [1 mark]
- (To be in the) correct orientation; [1 mark]

[Total: 2 marks]

- For a chemical reaction to occur, the particles must collide in the correct orientation and with enough energy

b) In terms of collisions, increasing the concentration increases the rate of reaction by:

- Increasing the number of (effective / successful) collisions per unit time
  OR
  Increase the frequency of (effective / successful) collisions
  OR
  Having particles collide (effectively / successfully) more often; [1 mark]

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[Total: 1 mark]

- The idea of the frequency of (effective / successful) collisions is one of the most commonly missed marks in an exam
- **TIP:** Adding the phrase ***more frequent effective collisions*** should be sufficient to score 2 common exam marks when the rate of reaction increases
  - Obviously, you should write ***less*** *frequent effective collisions* if that is appropriate to the question you are answering

c)

i) The role of the manganese dioxide in this reaction is:

- Catalyst
  OR
  To increase the rate of reaction / to speed the reaction up; [1 mark]

ii) The average rate of reaction, in $cm^3\ s^{-1}$, for the first 100 seconds of the reaction is:

- $\left(\frac{7}{100}\right)$ = 0.07 ($cm^3\ s^{-1}$); [1 mark]

[Total: 2 marks]

- The manganese dioxide is not involved in the chemical equation so it is not a reactant or an oxidising / reducing agent
- Manganese dioxide is not an acid as it does not have hydrogen ions
- Manganese dioxide is not an alkali as it does not dissolve in water to release hydroxide ions
- The only realistic role left for manganese dioxide is as a catalyst
- It is rare that you would be asked to use results directly from a table to calculate the rate, you would usually be asked to plot the graph and then use that to calculate the rate
- **Remember:** If you aren't sure what to calculate, the units help you $cm^3\ s^{-1}$ or cm^3 per second
  - This is telling you to have a value for the volume and divide it (per) by the time

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d) To find the rate of the reaction, in $cm^3\ s^{-1}$, at t = 300 s

- Volume at 200 s = 14.5 (cm^3)
  AND
  Volume at 400 s = 21.5 (cm^3); [1 mark]
- Difference in volume = 21.5 - 14.5 = 7 (cm^3); [1 mark]
- Rate $\left(= \frac{7}{200}\right)$ = 0.035 ($cm^3\ s^{-1}$); [1 mark]

[Total: 3 marks]

- If your final answer is 0.035 $cm^3\ s^{-1}$ then all three marks would be awarded, even without working
  - For the second mark, other values can be read off the tangent as long as they lead to the same final answer
  - Rates in the range of 0.0325 - 0.0375 $cm^3\ s^{-1}$ would be accepted for the final mark
- In an exam, you could be expected to plot the graph, add the tangent and then calculate the rate
- **Tip:** Extend your tangent to points that will be easy to read off
- Examiners will usually have a range of acceptable answers because the position of the tangent will be slightly different across all students

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Answer 2.

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a)

i) The definition of the term 'heterogeneous catalyst' is:

- The catalyst is in) a different phase / state to the reactants; [1 mark]

ii) A catalyst has no effect on the yield of the products in the reaction because:

- (A catalyst) increases the rate of the forward and reverse reactions to the same extent; [1 mark]

[Total: 2 marks]

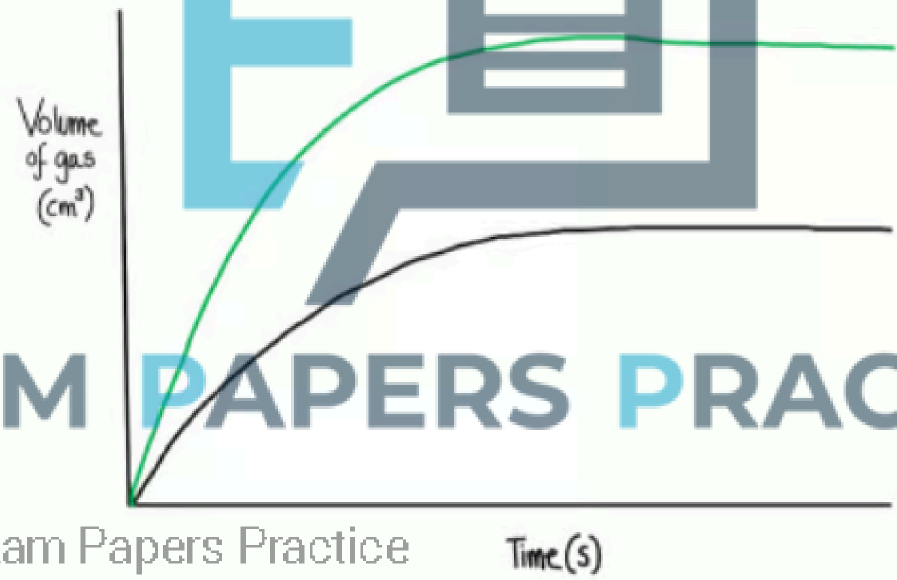
- A catalyst only speeds up the rate of the reaction therefore the equilibrium is met faster
- However, the same net amount of products is produced, just at a quicker rate

b)

i) The letter **R** on the students' graph represents:

- The rate of reaction at this point is zero
  OR
  The reaction is complete / has stopped / no more product will be produced
  OR
  All of the reactants have been used up / reacted; [1 mark]

ii) The sketch of the graph if students used 0.5 g of **A** and 50 cm^3 of 2.0 mol dm^{-3} of **B**:



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- The curve is steeper than the original curve and still starts at the origin; [1 mark]
- The curve levels / plateaus at double the height of the original curve; [1 mark]

[Total: 3 marks]

- Once all of the reactants have reacted together, no more product can be produced, meaning that the graph will plateau
- The faster the rate of the reaction, the steeper the gradient/slope of the graph
  - In the second experiment, the concentration of **B** is doubled, but **nothing else is changed**
  - So, the reaction happens faster and double the amount of product in this case is formed

c) The gradient of the curve decreases as the time of the reaction progresses because:

- The concentration of reactants decreases
  OR
  The reactants start to be used up; [1 mark]
- The rate of the reaction / rate of the formation of the gas decreases
  OR
  The frequency of successful collisions decreases / falls / is slower; [1 mark]

[Total: 2 marks]

- The faster the rate of the reaction, the steeper the gradient/slope of the graph
  - As reactants start to be used up, the gradient of the graph starts to decrease
  - Once all of the reactants have been used up, the graph plateaus and no more product is produced
  - The rate of the reaction at this point is zero; the reaction has finished

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d) A small increase in temperature has a large effect on the rate of a chemical reaction because:

- Many / significantly more molecules / particles will have energy greater than or equal to the activation energy
  OR
  A much higher proportion of molecules / particles will have energy greater than or equal to the activation energy; [1 mark]
- The proportion of collisions which are effective / successful will increase
  AND
  The frequency of successful collisions will increase; [1 mark]

[Total: 2 marks]

- For a chemical reaction to take place, molecules / particles must collide in the correct orientation and with sufficient energy
  - They must have energy greater than or equal to the activation energy
- Increasing the temperature gives the particles more energy, which means that a greater proportion of the particles will have the activation energy
- Particles will also be moving faster and colliding more frequently
- Since the proportion of particles with the activation energy is higher, and the particles are colliding more frequently, the frequency of effective collisions will increase significantly
- Increasing the temperature has the greatest effect on increasing the rate of the reaction of all of the possible factors

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Answer 3.

a)

i) The term effective collision means:

- A collision that results in a reaction / products being formed; [1 mark]

ii) **Two** factors that could cause an ineffective collision are:

- The (reactant) particles may not collide in the correct orientation; [1 mark]
- The (reactant) particle may not have an energy greater than the activation energy (of the reaction); [1 mark]

[Total: 3 marks]

- These are some of the key terms for this particular topic that you need to be able to define and understand

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b) Discussing the statement about how the rate of reaction is affected by temperature:

- (As temperature increases) there are more collisions (so the statement is true); [1 mark]
- But the rate depends on the frequency of effective collisions (not just the number of collisions); [1 mark]
- As the temperature increases the (average) kinetic energy / speed (of the particles) increases; [1 mark]
- This increases the frequency of collisions; [1 mark]
- There is a higher proportion of collisions that exceeds E_a resulting in more effective collisions; [1 mark]

[Total: 5 marks]

- You should be aware that when explaining increases in the rate of reaction, you should always refer to the increased frequency of collisions, not just stating there are more collisions
- This is because the rate is a measure of the changes in the amount of reactant/product in a given amount of time - so there need to be more collisions per unit time for this to occur
- When temperature increases, there are two reasons why the rate occurs
    - As the particles are moving quicker due to an increase in kinetic energy, they will collide more frequently
    - A greater proportion of these collisions are effective as more particles have energy that exceeds E_a
- The second factor is the most important in increasing the rate of reaction

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c) To describe and explain the effect of increasing the pressure on this reaction:

- The rate increases; [1 mark]
- As there are more molecules / particles in a given volume, so more frequent collisions; [1 mark]

[Total: 2 marks]

- There are keywords that you must use in explanations of rate changes
- As seen in part **(b)**, it is important not just to state 'more collisions', you must give a time factor, such as more collisions per second or more collisions per unit time
- It is also insufficient to state just 'more molecules' - you must explain that there are more molecules in a given volume, or per dm^3, or other unit of volume
  - **Careful**: Do **not** state there are more molecules in a given **area**
- You could also state that the molecules are closer together for the second mark

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d) A highly exothermic reaction is at more risk of exploding than a less exothermic reaction because:

- When the reaction starts the temperature / kinetic energy of the (reactant) particles / speed of the (reactant) particles increases (more than in a less exothermic reaction); [1 mark]
- (This means that) **many** / **many** more particles have an energy greater than E_a, which increases the rate of reaction (more rapidly than in a less exothermic reaction); [1 mark]

[Total: 2 marks]

- The 'suggest' command word indicates that you may not have specifically studied this but you need to apply your knowledge to answer it
- The question implies that as a highly exothermic reaction is more at risk of exploding, then it is more likely that its rate of reaction will increase suddenly
- You need to explain why the rate of reaction of a highly exothermic reaction will increase more suddenly than the rate of a reaction of a less exothermic reaction
- **Remember**: The more molecules with an energy exceeding E_a, the quicker the rate of reaction, so for a reaction to have a sudden increase in the rate of reaction, there must be a sudden increase in the energy of the particles
- Highly exothermic reactions will release more energy than less exothermic reactions
  - After the reaction starts, the particles in the highly exothermic reaction will suddenly gain more energy, causing the rate of reaction to increase as more particles have an energy > E_a

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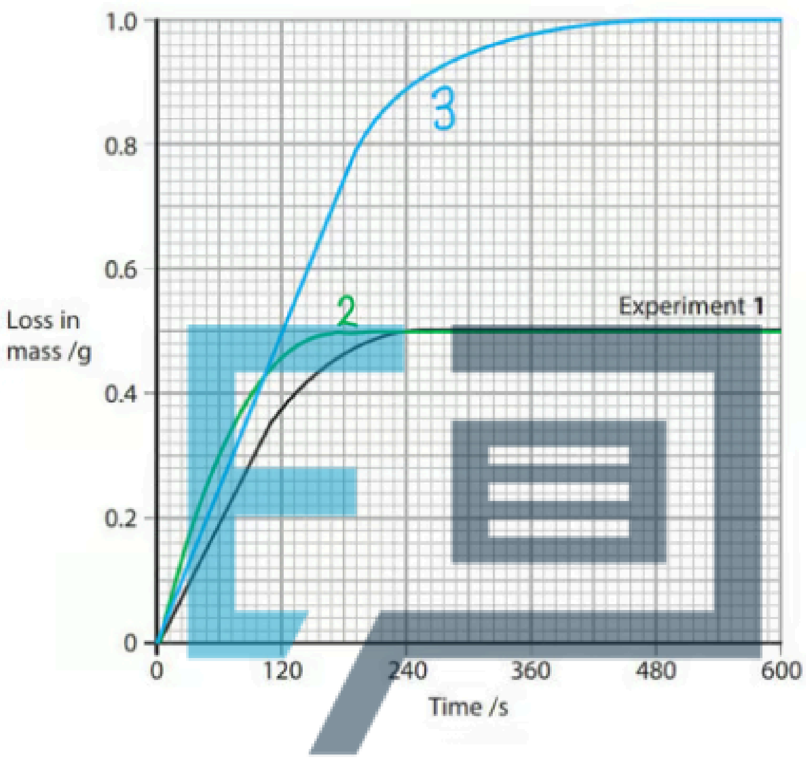
Answer 4.

a) The graph with curves for experiments **2** and **3** is:



- Experiment **2**: Curve steeper than Experiment **1**
  AND
  Curve finishes at the same mass loss as Experiment **1** / 0.5 g; [1 mark]
- Experiment **3**: Curve steeper than Experiment **1**
  AND
  Curve finishes at a mass loss of 1.0 g; [1 mark]

[Total: 2 marks]

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- This question requires you to compare pairs of experiments, identify the change and consider the impact it will have
- Experiments **1** and **2**:
  - The temperature of the reaction has increased
  - This gives the particles more energy leading to more frequent successful collisions and a faster rate of reaction (steeper line)
  - The amount of reactants have not changed so the mass loss will be the same
- Experiments **1** and **3**:
  - The concentration of HCl has increased
  - This means that there are more acid particles in the same volume leading to more frequent successful collisions and a faster rate of reaction (steeper line)
  - HCl is the limiting reagent
    - 10 g of calcium carbonate = $\frac{10}{100}$ = 0.1 moles
    - 50cm^3 of 0.50 mol dm^{-3} HCl = 0.50 x $\frac{50}{1000}$ = 0.025 moles
    - 50cm^3 of 1.00 mol dm^{-3} HCl = 1.00 x $\frac{50}{1000}$ = 0.05 moles
  - Since the concentration of HCl has doubled, the mass loss will also double from 0.5 g to 1.0 g

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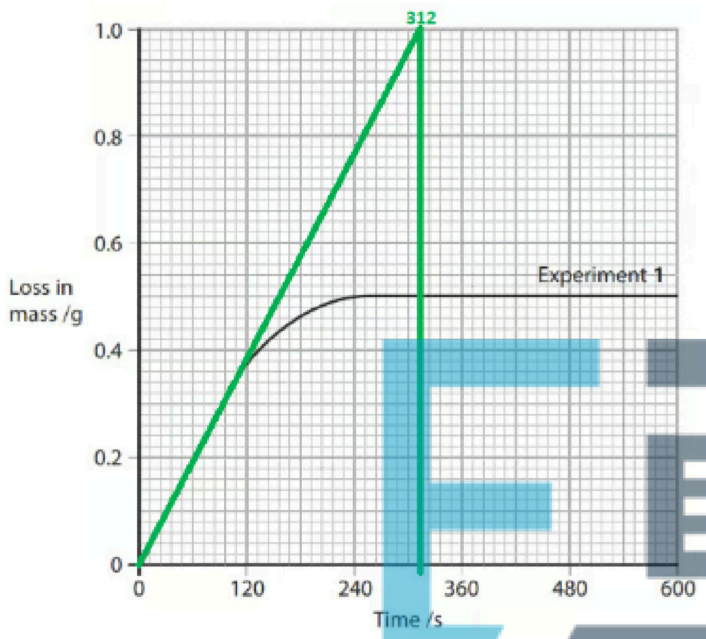
b) The initial rate of reaction for Experiment **1** is:



- Tangent drawn to the curve from time = 0
  AND
  Tangent extends to at least 120 s; [1 mark]
- Gradient calculation = $\frac{1.0}{312}$ = 3.21×10^{-3}
  OR
  Answers in the range of 3.13×10^{-3} to 3.53×10^{-3}; [1 mark]
- Gradient units = g s^{-1}
  OR
  g / s; [1 mark]

[Total: 3 marks]

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- The question says to show your working on the graph, which means that there is a mark for drawing the tangent
- When drawing tangents, the line should be extended as far as is convenient for you to perform the calculations
- Extending the tangent in this way decreases the amount of uncertainty
- **Remember:** Gradient = $\frac{\text{change in y} - \text{axis}}{\text{change in x} - \text{axis}}$
    - **Careful:** Each small square on the x-axis is worth 12 s
    - You can also deduce the units using the gradient equation by substituting the appropriate units for each axis
- If you didn't draw a tangent but used the straight line part of the graph from t = 0 sec to 108 sec where the change in mass is 0.35 g, you would score 2 marks
    - This would give a rate of 3.24×10^{-3} g / s

c) The average rate of reaction over the first 240 seconds of Experiment 1 is:

- 0.0021 / 0.00208 / 0.002083
  AND
- g s^{-1};

[Total: 1 mark]

- The average rate is calculated by the total mass lost in 240 s ÷ 240
- Average rate = 0.5 ÷ 240 = 0.0021 g s^{-1}
- You must give a unit in your answer, however, you would get error carried forward from part **(b)**

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d) The effect of using larger sized pieces would have on the rate of reaction is:

- The rate would decrease; [1 mark]
- The surface area is smaller, so less frequent (effective) collisions (between reactant particles); [1 mark]

[Total: 2 marks]

- Whilst the effect of surface area on rate of reaction is not specifically covered within this specification, you may have covered it in your previous studies
- The 'suggest' command word indicates that you need to apply your knowledge
- The rate of reaction depends on the frequency of effective collisions
- Larger particles will have a smaller surface area than the same mass of smaller particles, so there is less contact between reactant molecules
- This means fewer collisions will occur, resulting in less frequent effective collisions and the rate of reaction will decrease

Answer 5.

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a)

i) The mean rate of reaction over the first part of the experiment is:

- Rate = $\frac{change\ in\ volume}{change\ in\ time} = \frac{50\ cm^3}{15\ s} = 3.3\ (cm^3\ s^{-1})$; [1]

ii) The actual rate at 15 s is:

- Tangent drawn on the graph at 15 s; [1]
- Gradient of the tangent used to calculate rate; [1]
- Rate of 1.5 to 2.0 ($cm^3\ s^{-1}$); [1]

iii) The difference in the values is because:

- The fastest rate is at the start of the reaction/ the reaction slows down by 15 s; [1]
  OR
  The graph has started to curve by 15 s/ the gradient is shallower at 15 s; [1]

[Total: 5 marks]

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- The mean rate of a reaction is an appropriate calculation for the proportion of the initial reaction where the graph is essentially a straight line
- After this, the rate is best calculated by drawing a tangent to the curve at the point of the given time, and taking the gradient of that tangent. The larger the line drawn, the smaller the uncertainty of the measurement, so draw big tangents!
- Gradients are always calculated by the change in the value on the y-axis divided by the change in value on the x-axis
- A steeper line should give a larger value for the gradient, which represents a faster rate of reaction

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b) Using 25 cm^3 of 0.2 mol dm^{-3} nitric acid would mean that:

- The initial rate of reaction would increase/ the initial gradient of the graph would increase; [1]
- The end of the reaction would be reached sooner/ OWTTE; [1]
- The final/ total volume of gas produced would remain at 60 cm^3/ the same; [1]

[Total: 3 marks]

- The two solutions here would contain the same number of moles of acid, which is the limiting reagent since part a) states that the magnesium is in excess, so the final volume of gas produced would be the same in both cases
- The more concentrated acid would react more quickly due to more frequent collisions occurring, and the reaction would end faster due to the acid being used up more quickly

c) Changes to the reaction that would produce more hydrogen gas are:

- Increase the concentration of the acid without decreasing the volume;
  OR
  Increase the volume of the acid without decreasing the concentration;
  OR
  Increase the number of moles of acid reacting; [1]
- Increasing the moles/ particles of the limiting reagent increases the amount/ moles/ volume of the product; [1]

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[Total: 2 marks]

- The only way to increase the amount (volume/ mass/ moles) of a product is to increase the moles of the limiting reagent being used
- In this case the limiting reagent is an acid and therefore volume or concentration of the acid can be changed

d) It is often better to study a slower reaction as:

Any **two** from the following:

- Shorter reaction times give a larger (%) error in timing; [1]
- It is possible to compare reactions from the initial rate only (so the reaction does not need to always be complete); [1]
- More readings can be taken in the initial part of the reaction, reducing the change of anomalies being missed/ OWTTE; [1]

[Total: 2 marks]

- Reactions that are very fast may seem time efficient and more exciting, but the increased error associated with taking the recordings often makes the calculated rates less accurate
- The initial section of the reaction is all that is needed when comparing rates; an example of this is when clock reactions are conducted. The colour change occurs in the first 10–20% of the reaction and is still suitable to accurately compare rates across reactions

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