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VCE Chemistry ¾
Rates of Reaction [2.6]
Test Solutions

22 Marks. 1 Minute Reading. 17 Minutes Writing

Results:

Test Questions	_____ / 15
Extension Questions	_____ / 7

Section A: Test Questions (15 Marks)

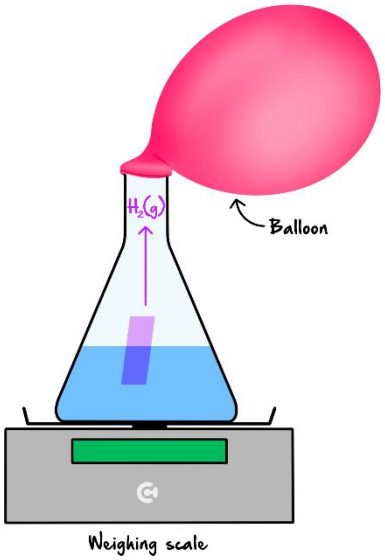
Question 1 (4 marks)

Tick whether the following statements are true or false.

Statement	True	False
a. The only two factors which need to be accounted for when discussing rates of reaction are: The fact that reactants must collide and the fact that they must collide with sufficient energy to react (greater than E_a).		✓
b. In general, the more collisions that occur between reactant particles, the greater the rate of reaction.		✓
c. The rate of reaction is linked to the gradient of a graph, whereas the yield is typically related to the height at which it plateaus.	✓	
d. If water is added to a reaction occurring in solution, the rate of reaction will decrease.	✓	
e. Using granular powder increases the rate of reaction as there is more of the reactant chemical present, thus leading to a greater frequency of total collisions and consequently, a greater frequency of successful collisions.		✓
f. Decreasing temperature has a two-fold effect on decreasing the rate of reaction: The particles collide less frequently, and when they do collide, they do so with less force on average.	✓	
g. Catalysts increase the rate of reaction by simultaneously decreasing the E_a and ΔH values.		✓

h. There will be a decrease in mass over time for the following set-up due to the production of hydrogen gas.

$Mg_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H_{2(g)}$



✓

Question 2 (8 marks)

Hydrogen peroxide, H_2O_2, in aqueous solution at room temperature decomposes slowly and irreversibly to form water, H_2O, and oxygen, O_2, according to the following equation:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g) \qquad \Delta H < 0$$

a. What effect will increasing the temperature have on the rate of O_2 production? Use collision theory to explain your answer. (3 marks)

Marks	0	1	2	3	Average
%	13	25	40	21	1.7

The following are the key points for which marks were awarded:

- increase in rate
- increase in temperature increases the average kinetic energy of all molecules, leading to increased frequency of collisions/increased number of collisions per second and increased energy of collisions
- increase in temperature leads to an increased proportion (percentage/ratio) of (total) collisions that are successful, i.e. collisions with energy greater than the activation energy.

This question was generally well answered. Many responses that were not awarded full marks did not make the 'time' link and simply referred to an increased number of collisions rather than increased frequency of collisions.

The increased proportion of collisions that are successful is a significant factor in the effect of temperature increase on reaction rate. Other factors such as increased concentration and surface area also increase the frequency of collisions that are successful, but only increased temperature and the introduction of a catalyst increase the proportion of collisions that are successful.

Some students attempted to apply equilibria and Le Chatelier's principle to answer this question, despite the reaction being irreversible and being directed to answer the question in terms of collision theory. This was linked to students trying to incorporate the exothermic nature of the reaction into their responses.

b. When a small lump of manganese(IV) dioxide, MnO_2, is added to the H_2O_2 solution, the rate of O_2 production increases, but when powdered MnO_2 is added instead, the rate of O_2 production is greatly increased. The MnO_2 is recovered at the end of the reaction. State the function of MnO_2 in this reaction.
(1 mark)

Marks	0	1	Average
%	8	92	0.9

MnO_2 acts a catalyst/lowers activation energy.

c. A solution of H_2O_2 is labelled '10 volume' because 1.00 *L* of this solution produces 10.0 *L* of O_2 measured at standard laboratory conditions (SLC) when the H_2O_2 in the solution is fully decomposed.
Calculate the concentration of H_2O_2 in the '10 volume' solution, in grams per litre, when this solution is first prepared. (2 marks)

Marks	0	1	2	Average
%	47	31	22	0.8

1.00 L $H_2O_2(aq) \rightarrow$ 10.0 L O_2 at standard laboratory conditions

$n(O_2) = 10.0/24.8$

$= 0.403$ mol

$n(H_2O_2) = 2 \times n(O_2) = 0.806$ mol

$m(H_2O_2) = 0.806 \times 34.0$

$= 27.4$ g

$c(H_2O_2) = 27.4$ g L^{-1}

This question proved challenging for some students, who did not take into account the stated fact that '1.00 L of this solution produces 10.0 L of O_2'.

Common calculation errors included overlooking the mole ratio $2H_2O_2(aq) \rightarrow O_2(g)$ indicated in the equation, assuming there was 1000 g H_2O_2 in 1.00 L of $H_2O_2(aq)$ and not using V_m at standard laboratory conditions, as well as incorrect molar mass for H_2O_2 and incorrect concentration units.

d. Propose a method to determine how quickly a solution of H_2O_2 decomposes when stored at a particular temperature. (2 marks)

Marks	0	1	2	3	Average
%	43	24	22	11	1

Effective responses to this question included:

- using a known amount of H_2O_2 solution
- an appropriate quantity measured at time intervals or over a set time
- how the quantity was measured at time intervals or over a set time.

For example:

- Take a known amount (mass or volume) of H_2O_2 solution.
- Measure the $V(O_2)$ produced over time.
- Use gas syringe or appropriate gas collection apparatus.

Such a method would provide a qualitative indication of the rate of decomposition, but the amount of O_2 produced could be used to calculate the $n(H_2O_2)$ reacted over a set time to determine a quantitative rate of decomposition.

Other methods might include:

- measure the change in mass of a known amount of $H_2O_2(aq)$ using electronic balance
- measure the change in pH of $H_2O_2(aq)$ using a pH meter.

The methods listed above were all based on the temperature being a controlled variable. Some students interpreted the question as 'determining how quickly the solution decomposes at different temperatures', which was incorrect.

High-scoring responses also described graphing the quantity change against time and using the gradient of the graph as an indication of the rate of decomposition.

Space for Persona

Question 3 (3 marks)

Suveer is studying the Haber process, which involves reacting nitrogen and hydrogen to produce ammonia.

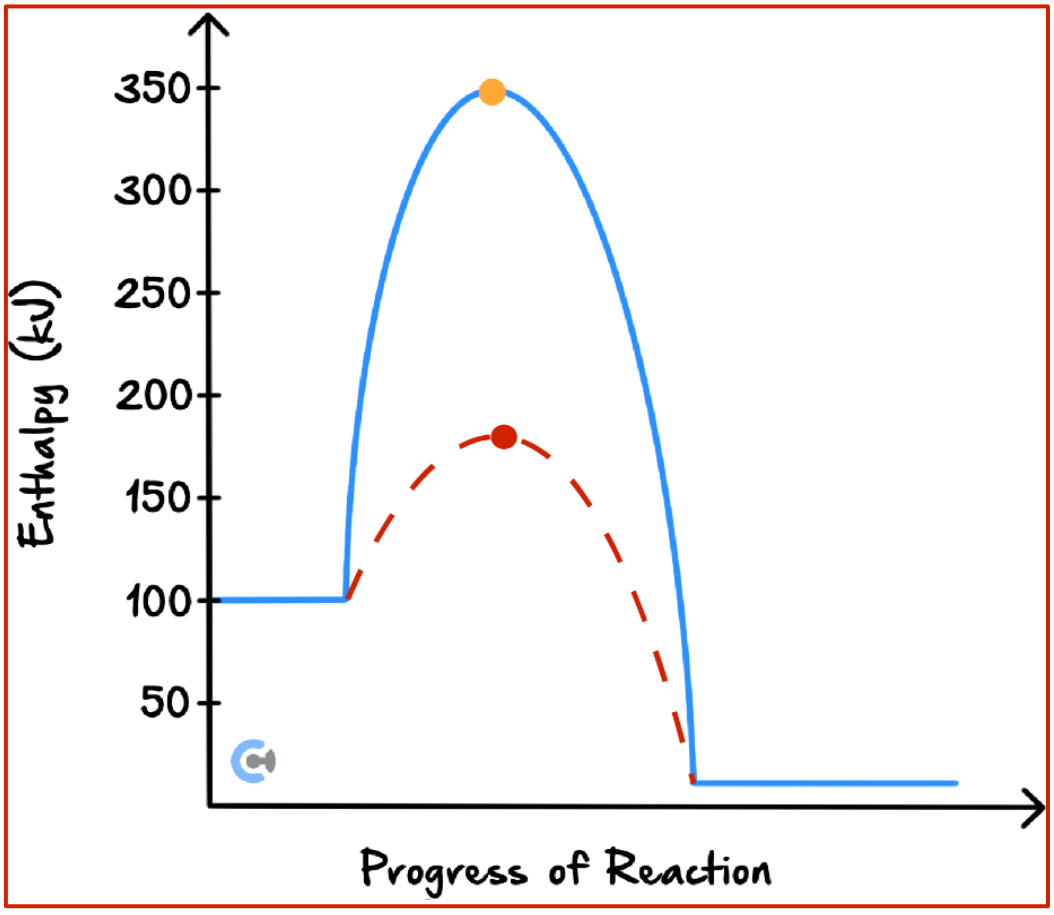
$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

He notices that the reaction is quite slow and as a result, decides to inject $10.0\ mL$ of argon gas into the vessel to increase the rate of reaction.

a. Explain what difference Suveer will observe in the reaction speed and why. (2 marks)

No difference. Inert gases do not affect the partial pressures/concentrations of reactant gases and therefore, the frequency of total collisions between **reactant** particles will be unaffected, and therefore, so will the frequency of fruitful collisions between them, and thus no impact on rate.

b. This process often makes use of an iron catalyst as it lowers the reaction's activation energy by $150\ kJ$. If the energy profile diagram below is for the **uncatalysed** reaction, draw what the catalysed reaction's energy profile diagram would look like on the same graph. (1 mark)



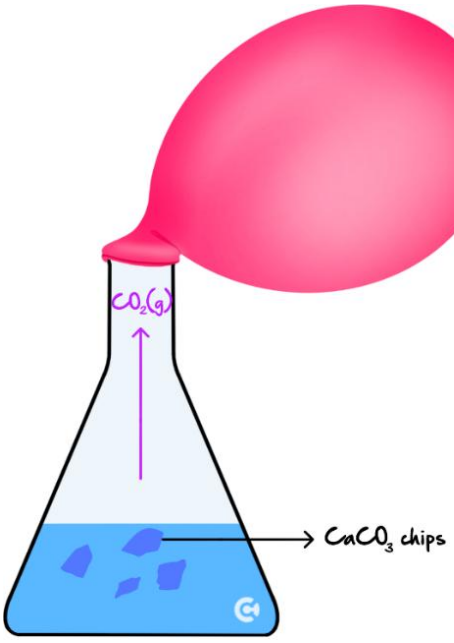
Section B: Extension Questions (7 Marks)

Question 4 (7 marks)

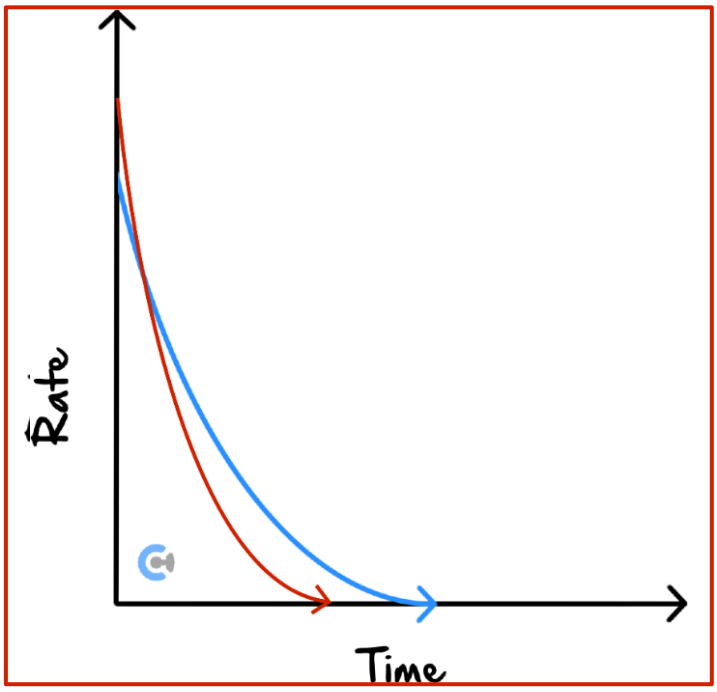
Emily, a Year 11 chemistry student, who has been studying acids and bases, is investigating the reaction between marble chips (calcium carbonate) and phosphoric acid as part of her extended investigation SAC. The overall reaction occurring is:

$$2H_3PO_4(aq) + 3CaCO_3(s) \rightarrow Ca_3(PO_4)_2(aq) + 3H_2O(l) + 3CO_2(g)$$

A diagram of her setup is shown below.



You are currently studying Unit 3 Chemistry. As her friend, you want her to go above and beyond to score 101% for this SAC. As a result, you decide to offer her some insight into how her reaction could be altered by various factors. To begin with, you provide her with the following graph to depict how the rate of the reaction occurring varies with time:



a. Had Emily used smaller, more granular marble chips instead, draw what effect this would have had on the rate of the reaction taking place, by drawing a curve on the same set of axes provided above. (1 mark)

b. In the original experiment, Emily used 10 g worth of marble chips and placed them into a 100 mL solution of 0.50 M phosphoric acid.

In a subsequent trial, if she decides to use identical chips, also weighing 10 g, but instead used 200 mL of 1.0 M phosphoric acid:

i. Determine the limiting regent in both the original trial and the subsequent trial. Show your working. (2 marks)

$n(H_3PO_4) = cV = 0.50 \times 0.1 \times 0.05$ mol

$n(CaCO_3) = \frac{m}{M} = \frac{10}{40.1+12+48} = 0.10$ mol

$n(CaCO_3)_{needed} = \frac{3}{2} \times n(H_3PO_4) = 0.075$ mol

$\therefore$ $CaCO_3$ is in excess in original trial.

subsequent trial: $n(CaCO_3) = 0.10$ mol (unchanged)

$n(H_3PO_4) = cV = 1.0 \times 0.2 = 0.20$ mol

$n(CaCO_3)_{needed} = \frac{3}{2} \times n(H_3PO_4) = 0.30$ mol $\rightarrow \therefore$

$\therefore$ Now H_3PO_4 is in excess.

ii. Hence or otherwise, explain what difference/s (if any) she would visually **observe** when conducting this subsequent trial, as opposed to the original experiment. (4 marks)

The rate of reaction would be **increased** as concentration of phosphoric acid is increased, so the balloon would inflate quicker (1). However, as the limiting reactant shifts, more $CO_2(g)$ can be produced in experiment 2, and therefore she would see a bigger/more inflated balloon in the subsequent trial. (2)

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