Beatty Secondary School
Science Department (Chemistry Unit)
Chemistry 6092

Name: ________________________________ (   )   Date: ______________

Class: 4E___

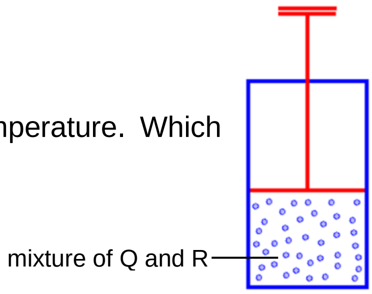
TOPIC: SPEED / RATE OF REACTION REVIEW (WORKSHEET 3)

Multiple-Choice Questions

1 Two gases Q and R react according to the equation.

$$Q\ (g) + R\ (g) \rightarrow T\ (g)$$

The reaction mixture is placed in a container at room temperature. Which action can speed up the formation of gas T?



A Adding an inert gas to the mixture.
B Using a bigger container.
C Lowering the piston in the container.
D Placing the container in water at 5 °C.

(C)

2 Catalase is an enzyme commonly found in living cells. It increases the rate of decomposition of hydrogen peroxide found in human bodies into water and oxygen gas.

Which statement about catalase is correct?

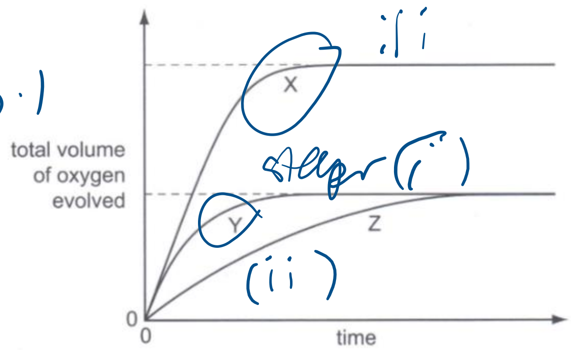
A It operates effectively at all temperatures.
B It operates most effectively over a certain range of pH.
C It increases the activation energy of the decomposition of hydrogen peroxide.
D It increases the yield of products from the decomposition of hydrogen peroxide. (B)

3 Hydrogen peroxide is catalytically decomposed by manganese(IV) oxide. To study the effect of the concentration of the solutions on the rate of reaction, the total volume of oxygen evolved was recorded against time. Three experiments were performed with a fixed mass of catalyst but with

0·05 x 2 = 0·1

(i) 50 cm^3 of 2.0 mol/dm^3 hydrogen peroxide
(ii) 100 cm^3 of 1.0 mol/dm^3 hydrogen peroxide 0·1
(iii) 100 cm^3 of 2.0 mol/dm^3 hydrogen peroxide
0·2

mol = C x V



On the graph above, which of the curves X, Y, Z relate to the solutions (i), (ii), (iii)?

	(i)	(ii)	(iii)
A	X	Y	Z
B	X	Z	Y
C	Z	X	Y
D	Y	Z	X

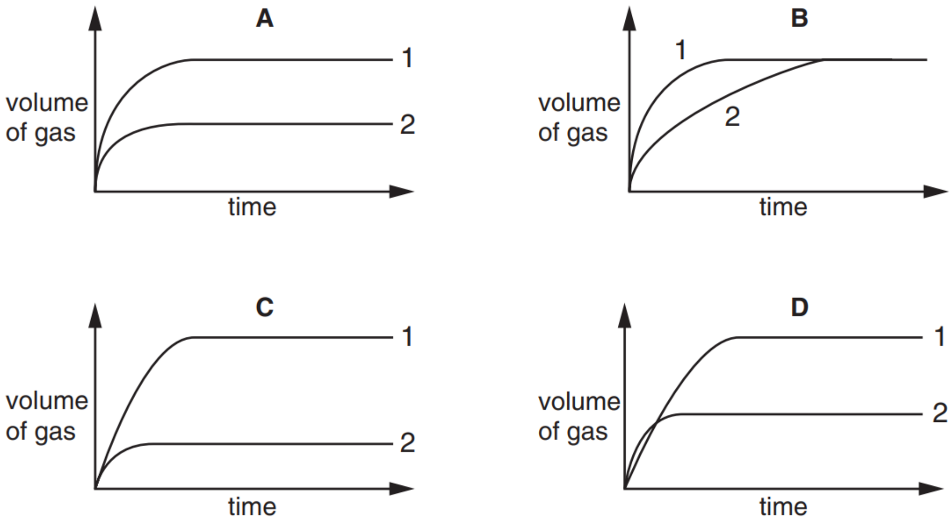
(D)

4 Calcium carbonate was reacted with an excess of dilute hydrochloric acid at room temperature.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

Two experiments were carried out.
Experiment 1: 100 g of calcium carbonate in large lumps.
Experiment 2: 50 g of calcium carbonate as a fine powder.

Which graph is correct?



(D)

5 A student investigated the rate of reaction of marble chips with excess dilute hydrochloric acid, by measuring the change in mass over a period of time. The experiment was then repeated using powdered calcium carbonate.

faster

How did the initial rate and the final mass change when powdered calcium carbonate was used?

yield vs rate

mol

	initial rate	final mass
A	faster	lesser
B	faster	unchanged
C	slower	unchanged
D	unchanged	unchanged

(B)

6 In the Contact Process, traces of vanadium(V) oxide is added in the oxidation of sulfur dioxide to sulfur trioxide in the equation as shown.

$$\overset{+4}{2SO_2} + O_2 \rightleftharpoons \overset{+6}{2\boxed{SO_3}}\ -6$$

oxidised O.A.

Which statement(s) correctly explain(s) why vanadium(V) oxide is added?

1 It acts as an oxidising agent to oxidise sulfur dioxide
2 It ensures that the yield would be 100%
3 It removes acidic impurities present in the reaction.
4 It speeds up the reaction by providing an alternative reaction pathway.

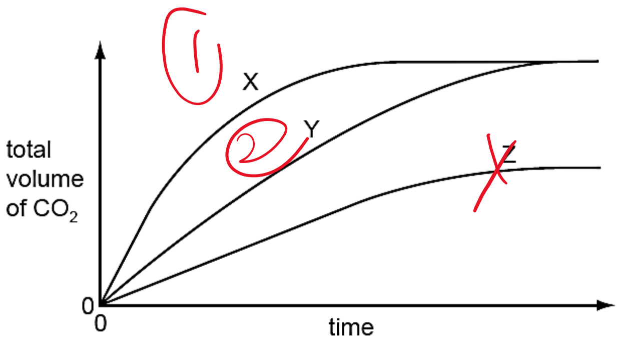
A 1 and 2 only **B** 1 and 4 only
C 3 and 4 only **D** 4 only (D)

7 Two experiments were carried out under the same conditions of temperature and pressure, reacting marble with dilute hydrochloric acid.

fast In experiment 1, an excess of powdered marble was added to 20 cm^3 of dilute hydrochloric acid.

In experiment 2, an excess of marble chips was added to 20 cm^3 of dilute hydrochloric acid of the same concentration.

The total volumes of carbon dioxide given off were measured at intervals and plotted against time.

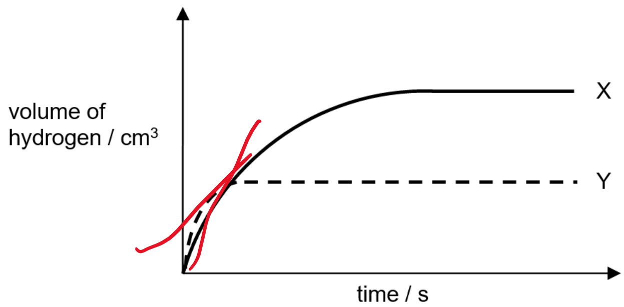


Which pair of curves would be obtained in both experiments?

	experiment 1	experiment 2
A	X	Y
B	X	Z
C	Y	X
D	Y	Z

(A)

8 Curve X shows the volume of hydrogen given off when 1.0 g of granulated zinc reacts completely with excess hydrochloric acid at 30 °C.

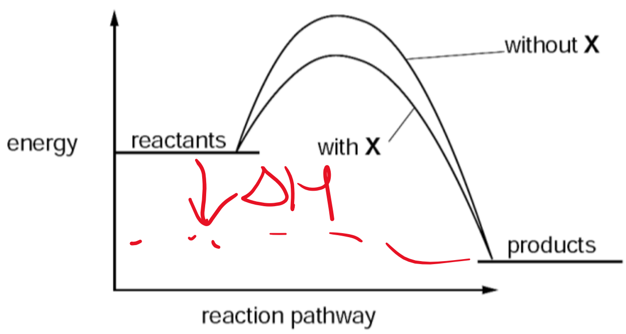


What change could produce curve Y?

A use 0.5 g of granulated zinc at 40 °C
B use 0.5 g of granulated zinc at 20 °C
C use 1.0 g of granulated zinc at 40 °C
D used 1.0 g of granulated zinc at 20 °C

(A)

9 The energy profile diagrams show how adding a substance **X** to a reaction mixture changes the reaction pathway.



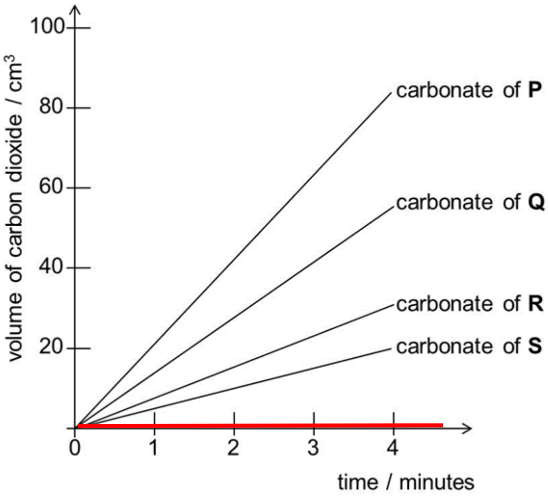
Which change is likely to be observed when **X** is added to the reaction mixture?

A The reaction becomes less endothermic.
B The reaction becomes less exothermic.
C The speed of reaction decreases.
D The speed of reaction increases.

(D)

Structured questions

10 When carbonates of some metals are heated, they decompose to produce carbon dioxide gas. The graphs below show the results of an investigation into the thermal stability of various metal carbonates. In each experiment, 0.0100 mol of carbonate was heated to the same temperature. The volume of carbon dioxide produced was measured every minute.



→ most CO_2 → easy to decompose
↓
least stable
↓
weaker f.o.a
↓
little amount of energy

(a) Suggest why, at the start of each experiment, the rate of reaction was very slow.

Thermal energy is absorbed to overcome the energy barrier of the activation energy.

(b) Explain how you would tell from the graphs that the decomposition of the carbonates was **not** complete?

The volume of carbon dioxide has not reached a constant value / is still increasing.

(c) A student suggested that the same mass rather than the same number of moles of carbonates should be used to ensure a fair comparison. Do you agree with him? Explain your answer.

No. Same number of moles of carbonate will yield the same volume of carbon dioxide gas for comparison. However, using the same mass of carbonate would yield different number of moles of carbonates and different volumes of gas produced due to the different molar mass of the metal carbonates. This will not be a fair comparison.

(d) **(i)** Based on the thermal stability of the various metal carbonates, arrange the metals in order of increasing reactivity.

P, Q, R, S

(ii) With reference to the graphs, briefly explain your answer in **(d)(i)**.

The higher the position of the metal in the reactivity series, the greater the thermal stability of the metal carbonate. As such, a larger amount of heat / thermal energy is needed to decompose the carbonate, releasing lower volume of gas per unit time.

(e) Sketch on the same graph, the graph expected when potassium carbonate is heated.

11 An experiment was carried out to measure the rate of reaction between excess powdered calcium carbonate and dilute acids.

(a) In **Experiment 1**, 20 cm^3 of 0.1 mol/dm^3 hydrochloric acid was used.

(i) Write a balanced chemical equation, with state symbols for the reaction.

$\underline{2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)}$

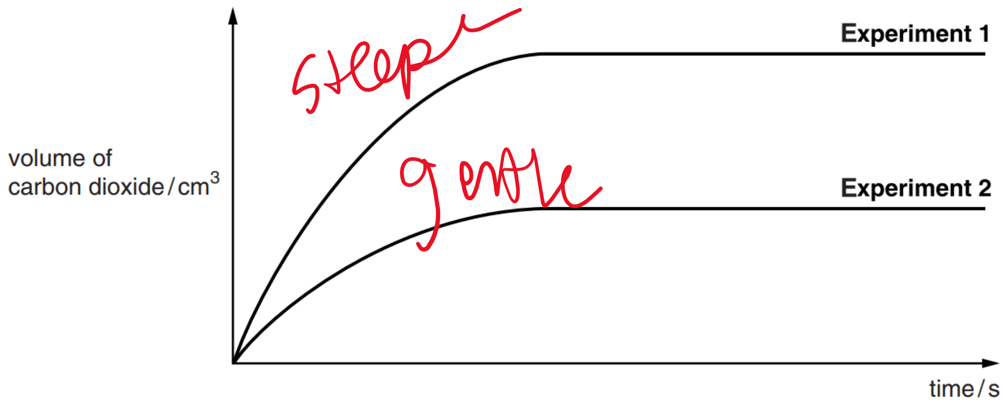
(ii) Calculate the volume of carbon dioxide made at room temperature and pressure.

No of mol of acid = **(20 ÷ 1000) × 0.1**
= **0.00200 mol**

No of mol of CO_2 = **0.02 ÷ 2**
= **0.00100 mol**

Volume of CO_2 = **0.01 × 24**
= **$\underline{0.0240\ dm^3}$**

(b) A further experiment using hydrochloric acid, **Experiment 2**, was carried out.
The results of **Experiments 1** and **2** are shown on the graph.



Suggest the concentration and volume of hydrochloric acid used for **Experiment 2**.

concentration: **<u>0.0500</u>** mol/dm^3 volume **<u>20.0</u>** cm^3

(c) A further experiment, **Experiment 3** was carried out using 20 cm^3 of 0.1 mol/dm^3 sulfuric acid. The initial rate of reaction for **Experiment 3** was faster than for the other experiments but the reaction stopped suddenly after only a small amount of gas had been given off.

(i) Name the salt formed in **Experiment 3**.

<u>calcium sulfate</u>

ABS

(ii) Explain why the reaction stops.

<u>Calcium sulfate is an insoluble salt and coats over calcium carbonate, preventing further reaction with the acid.</u>

(iii) Explain why the initial rate of reaction was faster than for the other reactions.

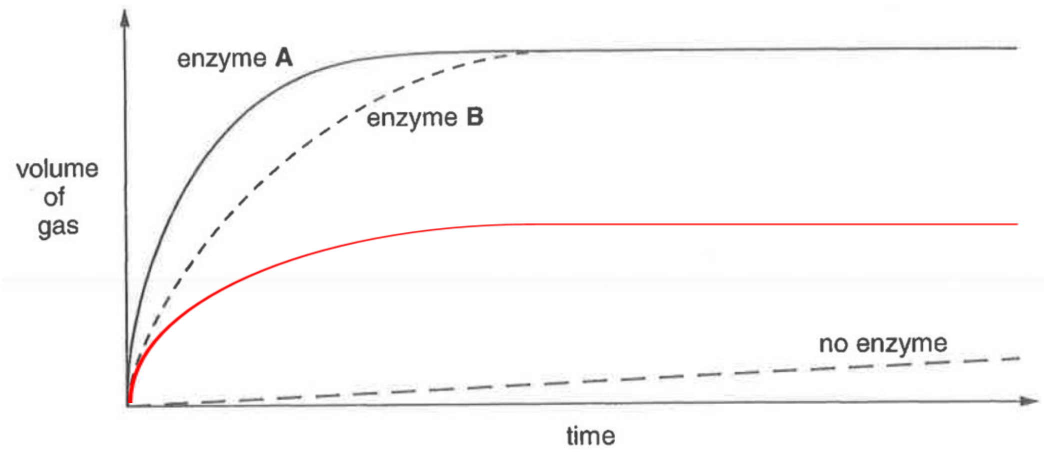
Sulfuric acid is a dibasic acid. For every 1 mole of sulfuric acid, 2 moles of H^+ ions are produced. Thus, there is a higher concentration of H^+ ions produced from sulfuric acid. Reaction becomes faster with a higher concentration of H^+ ions.

12 Hydrogen peroxide decomposes to give water and oxygen. A student carried out three experiments to investigate the rate of reaction. He used the same volume and concentration of hydrogen peroxide each time.

His first experiment only used hydrogen peroxide.

His second and third experiments investigated the effects of adding enzyme **A** and enzyme **B** to the reaction mixture.

The graph shows his results.



(a) Suggest a method the student could use to follow the rate of reaction.

He could measure the volume of gas produced at regular time intervals until the reaction is completed. /
He could measure the loss in mass of the flask and its contents at regular time intervals until the reaction is completed.

(b) **(i)** Describe the effect of each enzyme on the rate of reaction.

Both enzymes increase the rate of the reaction. However, enzyme A increases the rate faster than enzyme B.

(ii) Use ideas about activation energy to explain why the shape of the graph is so different when no enzyme was used.

When no enzyme is used, the reaction is very slow. This is due to the very high activation energy of the reaction.

(c) The student repeated the experiment for enzyme B. This time he used the same volume of hydrogen peroxide, but at half the original concentration.

He continued to measure the volume of gas until the reaction was complete. Sketch a curve on the graph to show the pattern of results for this experiment.

13 A series of experiments was carried out to investigate the effect of different catalysts on the rate of a reaction. The table shows the time taken for the reaction to finish when different metal compounds were used as catalysts.

The metal compounds contained Group I metals, Group II metals or transition metals.

experiment	catalyst	temperature at start / °C	time taken for reaction to finish / s
1	$NaCl$	19	45
2	$FeCl_2$	20	22
3	$CoCl_2$	19	26
4	$MgCl_2$	20	46
5	$NaNO_3$	19	45
6	$Fe(NO_3)_2$	20	22
7	$Fe(NO_3)_3$	19	15
8	$Co(NO_3)_2$	19	26
9	$Mg(NO_3)_2$	19	46

(a) Explain why it is important to take note of the temperature at the start of the reaction.

A higher temperature would lead to a faster rate of reaction. It is important to take note of the temperature to ensure that the initial temperature is constant. This is so that the results of the experiment will show the effects of different catalysts on the rate of a reaction.

(b) Group I and Group II metal compounds are less effective than transition metal compounds as catalysts.

Explain how the information in the table supports this statement.

The time taken for the reaction to complete using Group I and Group II metal compounds as catalysts are 45 seconds (based on experiments 1 and 5) and 46 seconds (based on experiments 4 and 9) respectively.

However, the time taken for the reaction to complete using transition metal compounds as catalysts are faster, ranging from 15 seconds in experiment 7 to 26 seconds in experiments 3 and 8. Hence, Group I and Group II metal compounds are less effective catalysts than transition metal compounds.

(c) Two different iron ions were used in the experiments.

(i) Give the formulae for the two ions.

Fe^{2+} and Fe^{3+} ions

(ii) Which iron ion appears to be the more effective catalyst? Explain your reasoning.

Fe^{3+} ion is a more effective catalyst. Comparing experiments 6 and 7, the time taken for the reaction to complete using Fe^{3+} ions (15 seconds) is shorter than using Fe^{2+} ions (22 seconds).

(d) A student wrote this conclusion from the results in the table.

The type of anion in the catalyst compound does not affect the rate of reaction.

(i) Do you agree with this conclusion? Explain your reasoning.

Yes, the type of anion does not affect the rate of reaction.

Comparing experiments 1 and 5, where the type of anion is different, the time taken for the reaction remains the same as 45 seconds. Similar results for the time taken for the reaction to complete are observed in experiments 3 and 8 (26 seconds), 4 and 9 (46 seconds), and 2 and 6 (22 seconds).

(ii) Predict the time taken for the reaction to finish if iron(III) chloride was used as a catalyst.

15 seconds

14 Some students investigated the effect of particle size of a reactant on the speed of reaction. Samples of quicklime (calcium oxide) of three particle sizes were dissolved in water and the temperature of each reaction mixture was measured at every half-minute for two minutes. The results of the experiments are shown in the table below.

time / min	**temperature / °C**		
	powdered quicklime	large lumps of quicklime	small lumps of quicklime
0	25.0	25.0	25.0
0.5	30.0	26.0	28.0
1.0	43.0	28.0	37.0
1.5	56.0	31.0	46.0
2.0	54.0	30.0	44.0

(a) Use the information in table to explain the results obtained.

At each 0.5 minute intervals, temperature rise increases in the order of large lumps, small lumps followed by powdered quicklime. This can be observed at the 1.0 minute mark, that the temperature rise increases from large lumps (3 °C), small lumps (12 °C), powdered quicklime (18 °C).

This effect is due to the increase in surface area in the same order. The greater the surface area, the faster the speed of reaction leading to a larger temperature change.

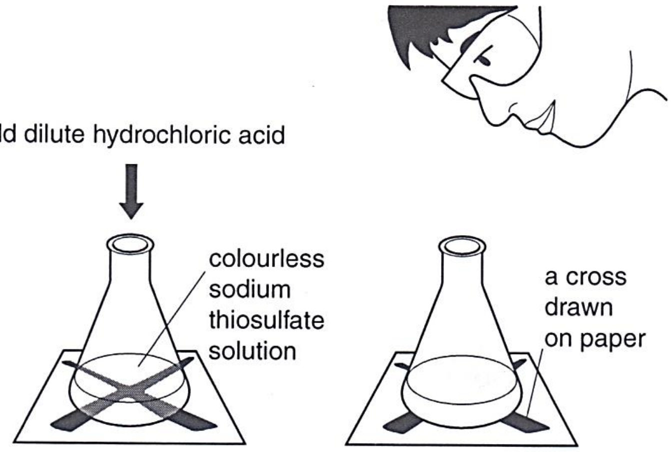
(b) State the type of energy change that has occurred in the reaction.

exothermic reaction

(c) Write a balanced chemical equation for the reaction between calcium oxide and water.

$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$

15 The rate of the reaction between dilute hydrochloric acid and sodium thiosulfate solution can be investigated using a cross drawn on a piece of paper.



As the reaction progresses, it becomes more difficult to see the cross through the solution.

(a) Look at the equation for the reaction.

$$2Na_2S_2O_3\,(aq) + 2HCl\,(aq) \rightarrow 2NaCl\,(aq) + SO_2\,(g) + H_2O\,(l) + S\,(s)$$

Explain why it becomes more difficult to see the cross as the reaction progresses.

As the reaction progresses, more sulfur is produced which appears as a pale yellow solid. Formation of the thick precipitate does not allow the cross to be seen / covers the cross.

(b) The table shows the results of some experiments to investigate the rate of reaction, using different concentrations of sodium thiosulfate.

A student measured the time from when the acid was added until the cross can no longer be seen. The same concentration and volume of dilute hydrochloric acid and the same volume of sodium thiosulfate were used each time.

concentration of $Na_2S_2O_3$ / mol/dm^3	time until cross cannot be seen / s
1.0	8
0.8	10
0.4	20
0.2	39

(i) Use ideas about collisions between particles to explain the trend in the results.

With a lower concentration (from 1.0 to 0.2 mol/dm^3), there are lesser reactant particles per unit volume of solution. There are less frequent collisions between reactant particles, resulting in the frequency of effective collision decreasing.
This slows down the speed of the reaction and a longer time is taken for enough sulfur to form so that the cross cannot be seen.

(ii) The experiment was repeated with another concentration of sodium thiosulfate. The cross could not be seen after 14 s. Estimate the concentration of sodium thiosulfate that was used.

0.6 mol/dm^3

(c) Some metal oxide can act as catalysts for the reaction.

(i) A student thinks that chromium(III) oxide acts as a catalyst for the reaction. Describe what he should do and what results he would obtain if he is right.

He could repeat the experiment using the same volume and concentration of dilute hydrochloric acid added to the same volume of sodium thiosulfate of concentration of 0.2 mol/dm^3 (or 0.4 mol/dm^3, 0.8 mol/dm^3, 1.0 mol/dm^3), but with a little chromium(III) oxide (0.1 g) added.
He could measure the time from when the acid was added until the cross could no longer be seen.
If he is right, the time measured should be less than 39 s (or 20 s, 10 s, 8 s).

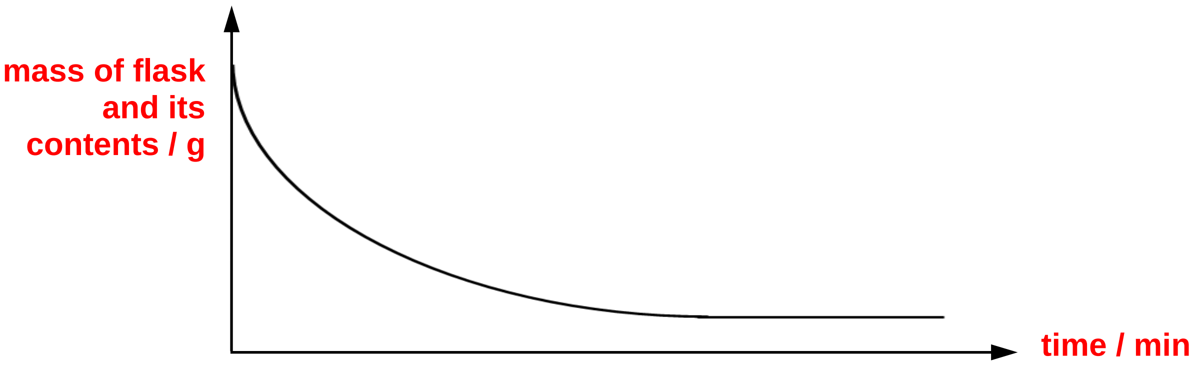
(ii) Catalysts lower the activation energy for the reaction. Explain how they do this.

They do so by providing an alternative pathway of lower activation energy for the reaction to occur.

16 A student investigates how the speed of a reaction changes over time.

Dilute hydrochloric acid is put in a flask. Excess powdered sodium carbonate is added to the acid in the flask. While the reacting mixture is fizzing, the flask and contents are carefully put on the pan of a balance. The mass of the flask and its contents are measured regularly until just after the fizzing has stopped.

(a) Sketch the graph that you would obtain from the data collected. Label both axes.



(b) Explain how your graph shows that the speed of the reaction changes over time.

The speed of reaction is the fastest at the start of the experiment as it has the largest / steepest gradient. The speed of reaction gradually decreases as the gradient becomes gentler. The speed of reaction becomes zero / stopped when the gradient becomes zero, indicating that the reaction is completed.

(c) Explain how you would calculate the **average** speed of the complete reaction.

The average speed can be calculated based on the difference in the mass of the flask and its contents divided by the total time taken when the reaction is just completed / fizzing has just stopped.

[PRSS PRELIM 2015]

17 This question is about the reaction between iodine and propanone. Aqueous iodine and propanone, CH_3COCH_3 react together in a small volume of acid according to the equation:

$$CH_3COCH_3\,(aq) + I_2\,(aq) \xrightarrow{H^+} CH_2ICOCH_3\,(aq) + HI\,(aq)$$

The concentration of propanone was measured when 0.0020 mol/dm^3 of iodine was used and the results are shown in Fig. 17.1.

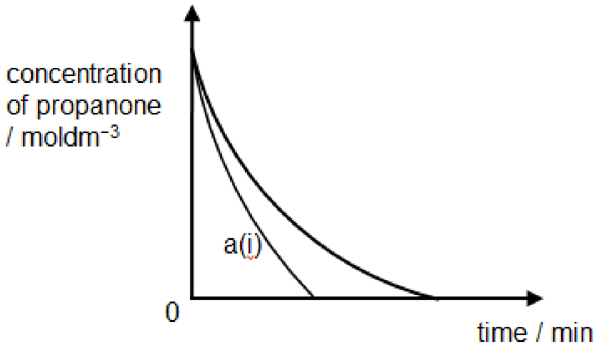


Fig. 17.1

Initial Rate of Reaction

To further investigate the reaction kinetics between aqueous iodine and propanone, the initial rate of reaction was determined through the changes in the concentration of the reactants. The initial rate of reaction is the rate of the reaction when the reagents are **first** brought together, as shown in Fig. 17.2.

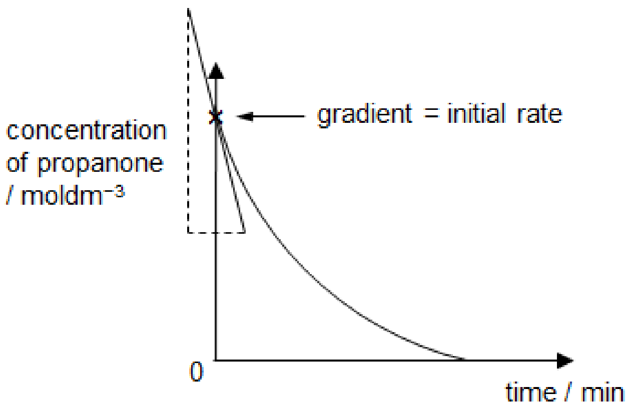
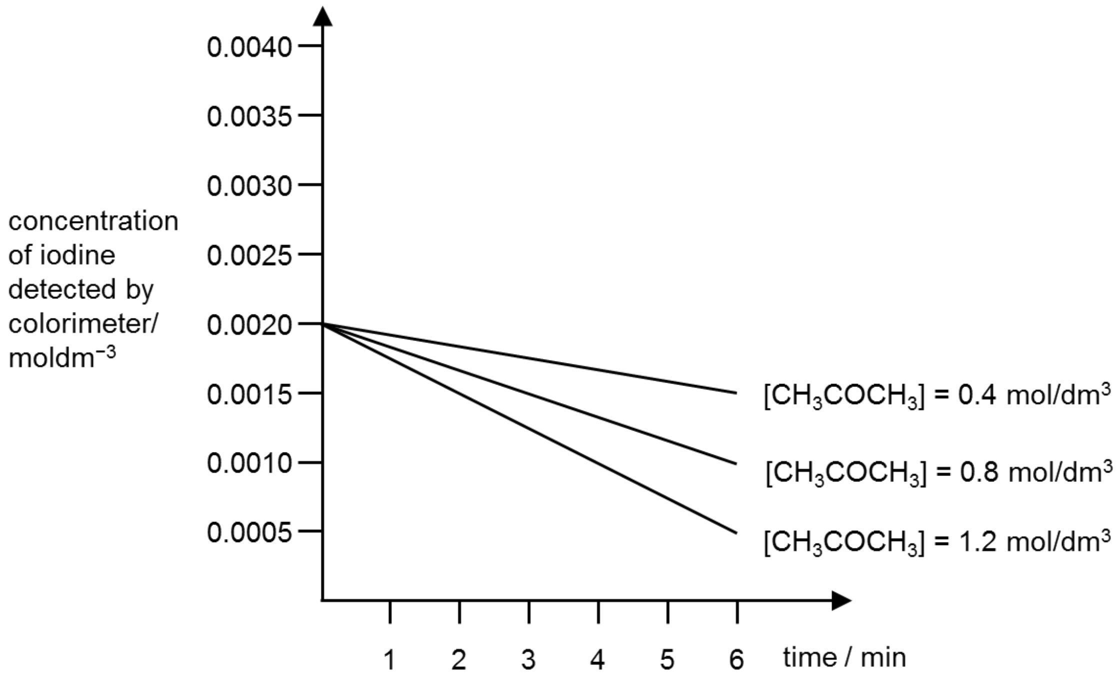


Fig. 17.2

Two sets of experiments were conducted in order to determine how the concentration of reactants affects the initial rate of reaction.

In Experiment 1, different concentrations of propanone were reacted with a fixed concentration of iodine. During the reaction, the concentration of iodine decreases and this causes the colour intensity of the reaction mixture to decrease. The changes in colour intensity are then measured by the use of a colorimeter. Thus, the concentration of iodine can be measured throughout the reaction. The initial rate of the reaction can then be determined by observing the change in the concentration of iodine over time. The graph in Fig. 17.3 shows how the concentration of iodine changes with time.

Experiment 1:

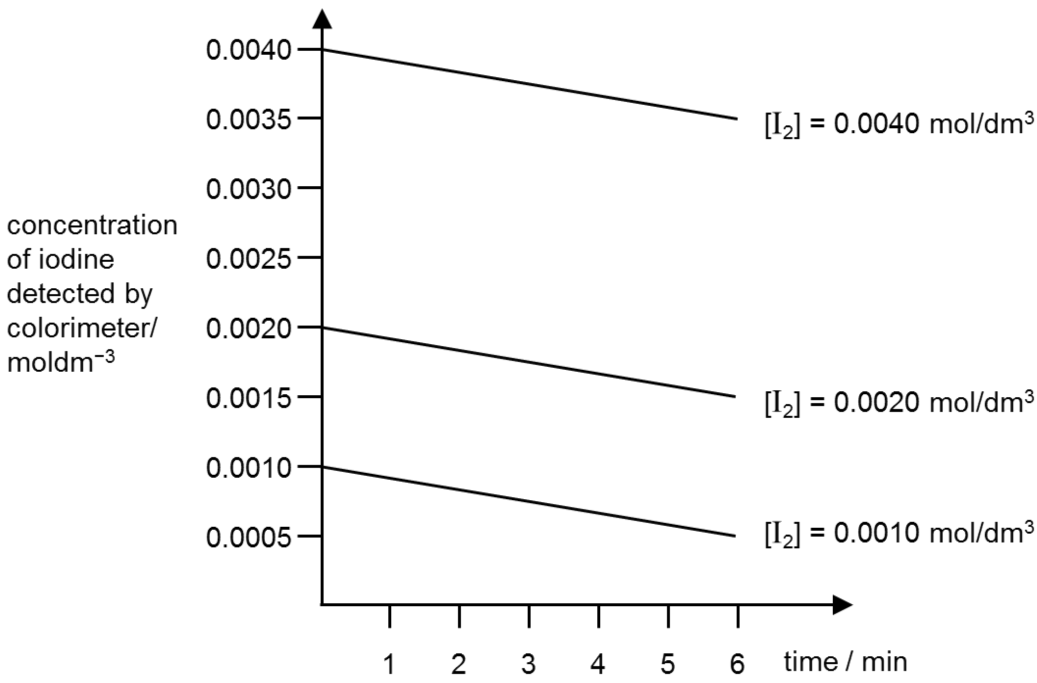


$[CH_3COCH_3]$ denotes the concentration of propanone

Fig. 17.3

In Experiment 2, different concentrations of iodine were reacted with a fixed concentration of propanone. The changes in the concentration of iodine were also measured in a similar manner. The graph in Fig. 17.4 shows how the concentration of iodine changes with time.

Experiment 2:



$[I_2]$ denotes the concentration of iodine

Fig. 17.4

(a) **(i)** Sketch on **Fig. 17.1**, the graph when 0.0040 mol/dm^3 of iodine was used.

(ii) With reference to **Fig. 17.1**, which reagent, propanone or iodine is in excess in this reaction? Explain your answer.

Iodine is in excess. Based on Fig. 17.1, all the propanone is completely used up in the reaction, indicating that propanone is the limiting reagent.

(b) What is the concentration of iodine detected by the colorimeter at **4 minutes** when 1.2 mol/dm^3 of propanone is used in **Experiment 1**?

0.00100 mol/dm^3

(c) The gradient of the graphs in Fig. 17.3 and Fig. 17.4 show the initial rate of reaction. How does the concentration of propanone and iodine affect the initial rate of reaction? Explain your answer.

Concentration of propanone:

As the concentration of propanone increases, the initial rate of reaction increases. This is illustrated by Fig. 17.3, where a steeper gradient was observed when higher concentrations of propanone were used.

Concentration of iodine:

As the concentration of iodine increases, there is no effect on initial rate of reaction. This is illustrated by Fig. 17.4, where the gradient remains constant when higher concentrations of iodine were used.

(d) **(i)** The acid added to the reaction between iodine and propanone serves as a catalyst. Explain, using ideas about collisions between particles, how a catalyst affects the rate of reaction.

A catalyst provides an alternative reaction pathway with a lower activation energy than the uncatalysed reaction. As such, there is an increase in the number of reacting particles colliding with energy equal to or greater than the activation energy and the frequency of effective collisions increases, increasing the rate of reaction.

(ii) Suggest why only a small volume of acid is required for the reaction between iodine and propanone.

The product formed is HI which can ionise to form hydrogen / H^+ ions. The hydrogen ions regenerated from the products can be used to catalyse the reaction, hence only a small volume of acid is needed.

(e) Predict the colour change observed at the end of Experiment 1.

Brown solution turns colourless / decolourises.