Mastering Chemistry

- Book 4A
- Topic 10 Rate of Reaction

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Content

37.1 Changing the rate of a reaction (p.32)

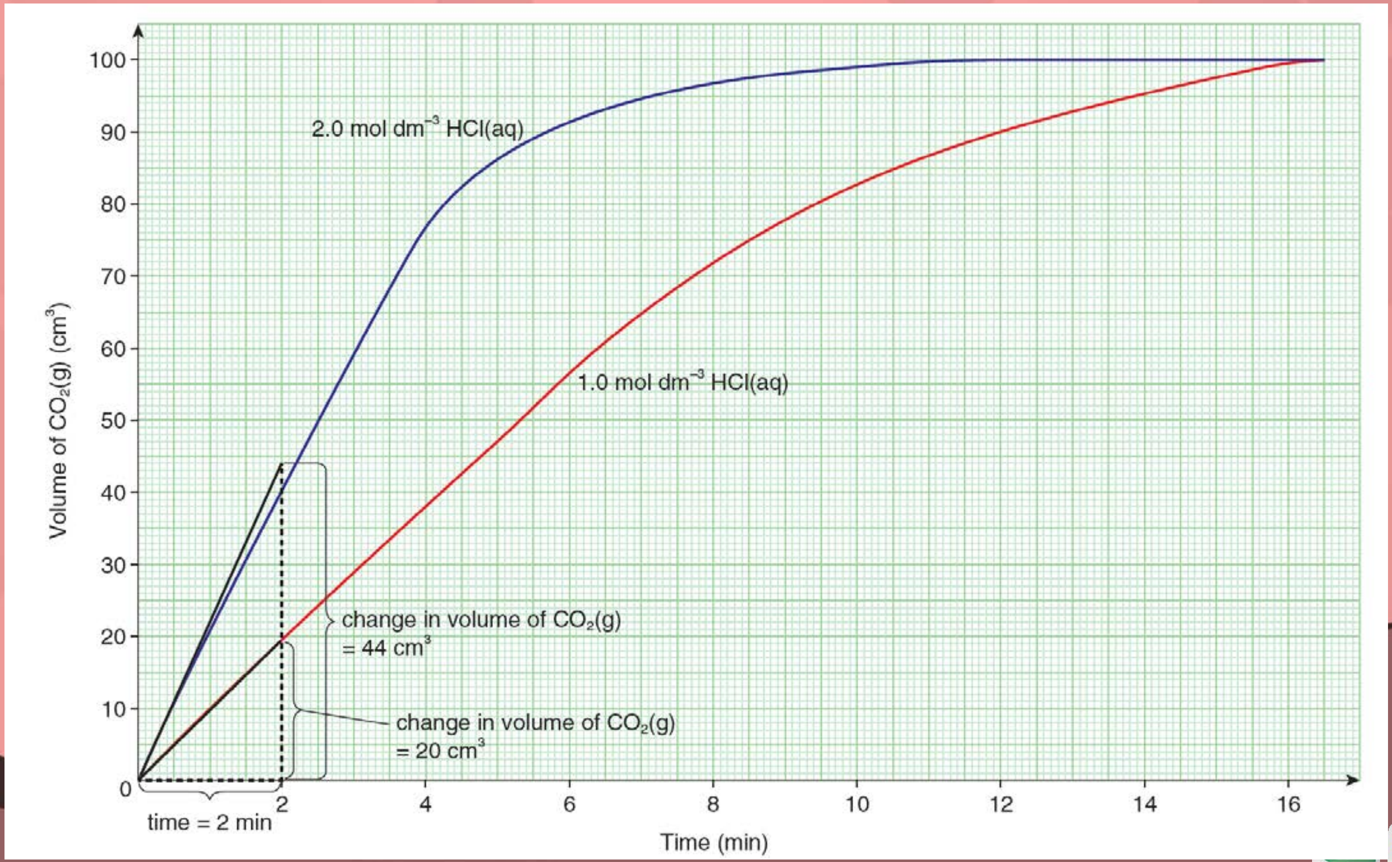
- Have you ever cooked? Do you know why fried potatoes cook faster than boiled potatoes?
- Cooking food by frying is quicker than by boiling. Temperature is one of the reasons. Hot oil has a much higher temperature than boiling water. Can you think of another reason?

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

- Consider the reaction between calcium carbonate and hydrochloric acid of different concentrations. The volume of carbon dioxide gas released is monitored over 16 minutes by using a gas syringe.
- One trial is carried out using a known mass of calcium carbonate with an excess of 2.00 mol dm^{-3} hydrochloric acid.
- This is repeated using the same mass of calcium carbonate and the same volume of 1.00 mol dm^{-3} hydrochloric acid, also in excess.

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

- The figure below shows the variations of volume of carbon dioxide released with time for the two different concentrations of hydrochloric acid.



37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

- Initial rate of reaction for 2.00 mol dm^{-3} HCl(aq) $= \dfrac{44\ cm^3}{2.0\ min}$ = 22 cm^3 min^{-1}
- Initial rate of reaction for 1.00 mol dm^{-3} HCl(aq) $= \dfrac{20\ cm^3}{2.0\ min}$ = 10 cm^3 min^{-1}
- The initial rate for 2.00 mol dm^{-3} hydrochloric acid is approximately twice the rate for the 1.00 mol dm^{-3} hydrochloric acid, so increasing the concentration of the acid increases the rate of the reaction.
- You must NOT assume that if you double the concentration of a reactant, the rate of a reaction will always be twice as high. The actual extent of the increase in rate is impossible to be predicted and can only be found by experiment..

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

The rate of a reaction increases when the concentration of a reactant is increased.

Investigating the effect of varying the concentration of hydrochloric acid on the rate of its reaction with magnesium

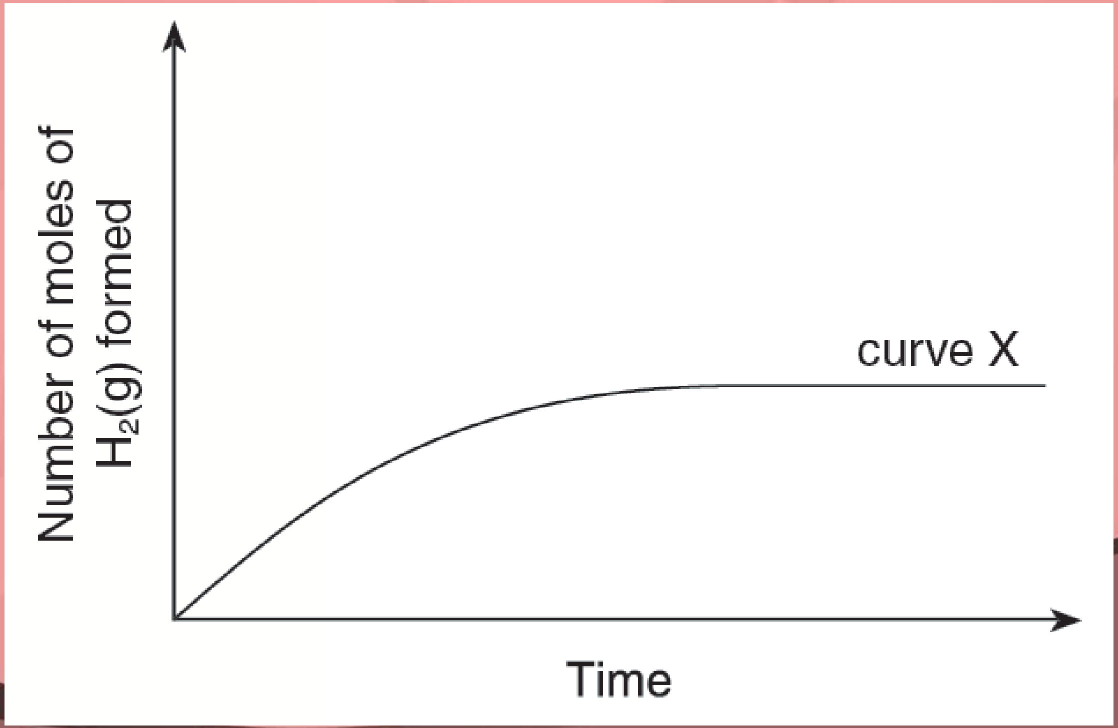
Investigating the effect of varying the concentration of hydroxide ion on the rate of its reaction with phenolphthalein

Effect of concentration on rate *Ref.*

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

Q (Example 37.1)

In an experiment, 0.0500 mole of zinc granules (excess) was allowed to react with 20.0 cm^3 of 1.00 mol dm^{-3} HCl(aq). Curve X shows how the variation of the number of moles of $H_2(g)$ formed with time.



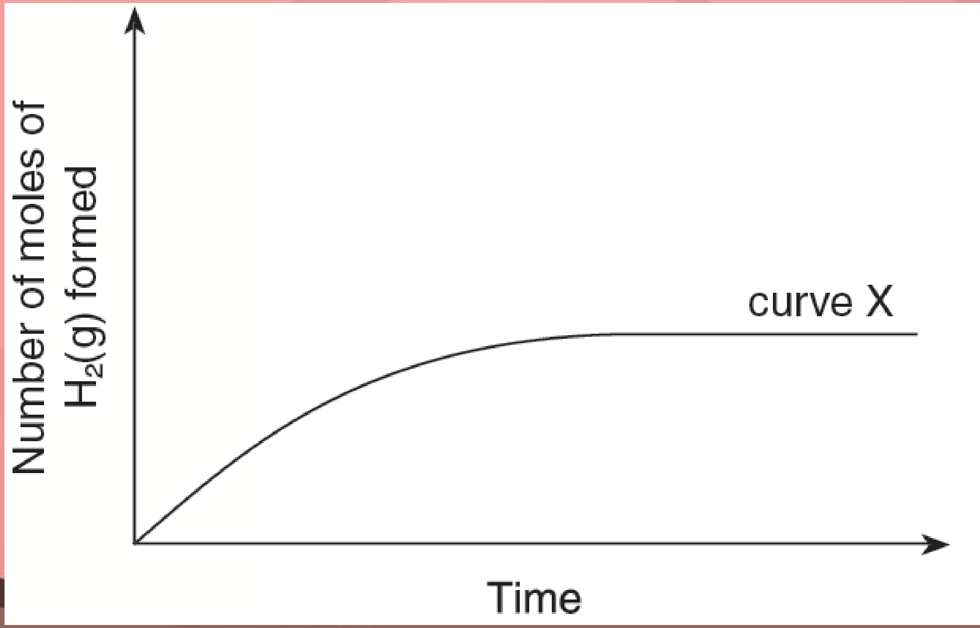
37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

a) What was the total number of moles of $H_2(g)$ formed?

b) The experiment was repeated using the same volume of 1.00 mol dm^{-3} $H_2SO^4(aq)$ instead of 1.00 mol dm^{-3} HCl(aq) while the other conditions remained constant. Zinc was also in excess.

i) Explain the effect of this change on the initial rate of this reaction.

ii) On the diagram below, draw a curve to show the results obtained.



37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

A

a) Zn(s) reacts with HCl(aq) according to the equation below:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Number of moles of HCl = $1.00\ mol\ dm^{-3} \times \frac{20.0}{1\,000}\ dm^3$ = 0.0200 mol

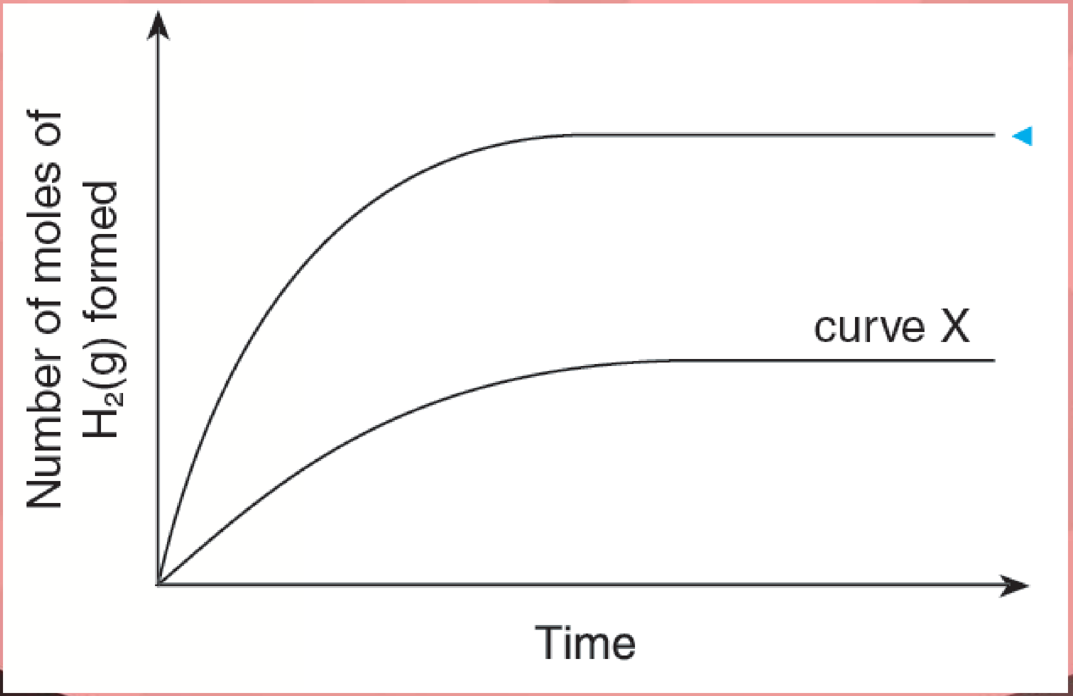
According to the equation, 1 mole of Zn reacts with 2 moles of HCl to form 1 mole of H_2.

i.e. number of moles of $H_2(g)$ formed = $\frac{0.0200}{2}\ mol$ = 0.0100 mol

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

b) i) $HCl(aq)$ is a monobasic acid while $H_2SO_4(aq)$ is a dibasic acid. The initial rate increases as the concentration of $H^+(aq)$ ions increases when using 1.00 mol dm^{-3} $H_2SO_4(aq)$.

ii)



Zn(s) reacts with $H_2SO_4(aq)$ according to the equation below:

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Number of moles of H_2SO_4

$$= 1.00 \text{ mol dm}^{-3} \times \frac{20.0}{1\ 000} \text{ dm}^3$$

$$= 0.0200 \text{ mol}$$

= number of moles of H_2

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

Q (Example 37.2)

Under certain conditions, a pink compound X reacts with NaOH(aq) to give a colourless product.

In an experiment, three NaOH(aq) solutions were prepared by mixing different volumes of 1 mol dm^{-3} NaOH(aq) with water. Then, two drops of X were added to each solution. The time taken for the pink colour to disappear was recorded. The table below shows the results obtained.

Trial	Volume of NaOH(aq) (cm^3)	Volume of water (cm^3)	Time taken for the pink colour to disappear (s)
1	6.0	4.0	203
2	8.0	2.0	153
3	10.0	0	122

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

a) Why is the total volume of the reaction mixtures kept the same for the trials?
b) Describe the effect of different concentrations of sodium hydroxide on the rate of this reaction.

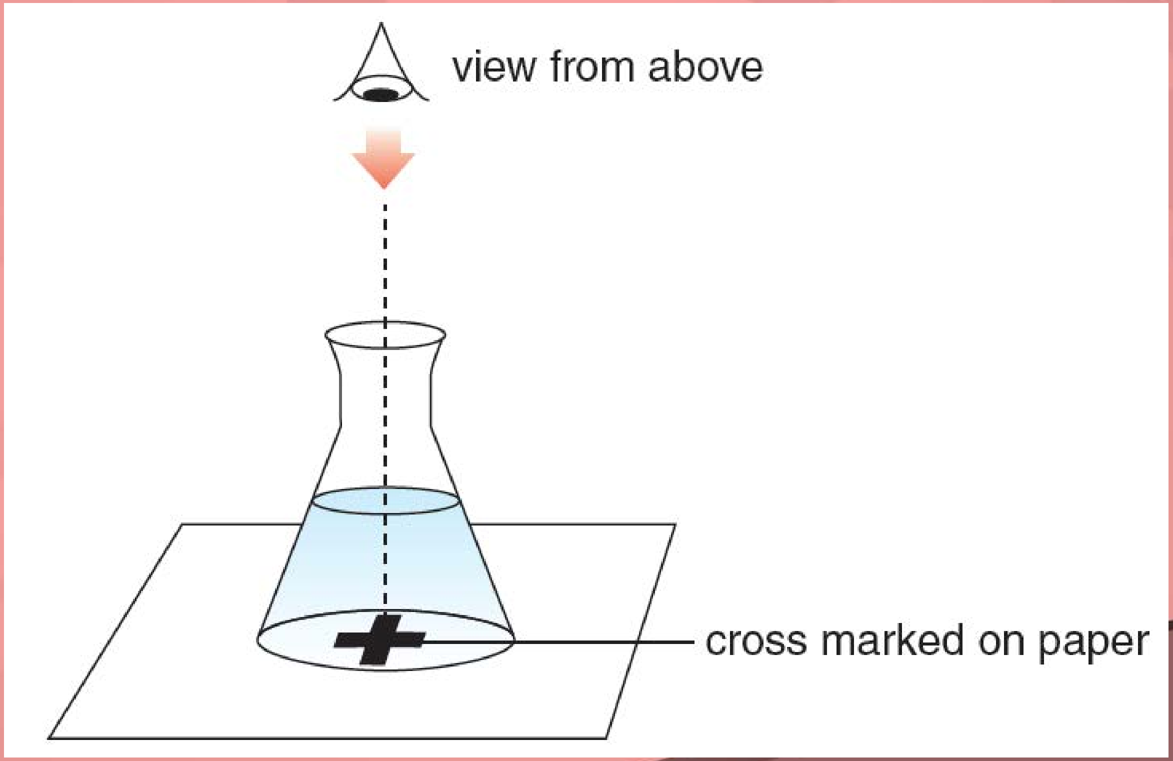
A

a) The volume of NaOH(aq) used can represent the concentration of sodium hydroxide in the reaction mixtures.
b) If the reaction is fast, the time for the pink colour to disappear will be short. If the reaction is slow, the time will be long.
The rate of the reaction increases when the concentration of sodium hydroxide is increased.

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

Practice 37.1

1 The experimental set-up shown below is used to study the rate of the reaction between X(aq) and Y(aq).



37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

Three trials of an experiment are carried out. In each trial, X(aq) is mixed with water in the flask. Then Y(aq) is added to the mixture. The time taken for the cross to become invisible when viewed from above is recorded. The table below shows the data obtained.

Trial	Volume used (cm^3)			Time (s)
	X(aq)	$H_2O(l)$	Y(aq)	
1	5.0	15.0	10.0	112
2	5.0	5.0	20.0	112
3	10.0	10.0	10.0	56

What is the effect, if any, of increasing the concentration of

a) X; and

b) Y

on the rate of this reaction?

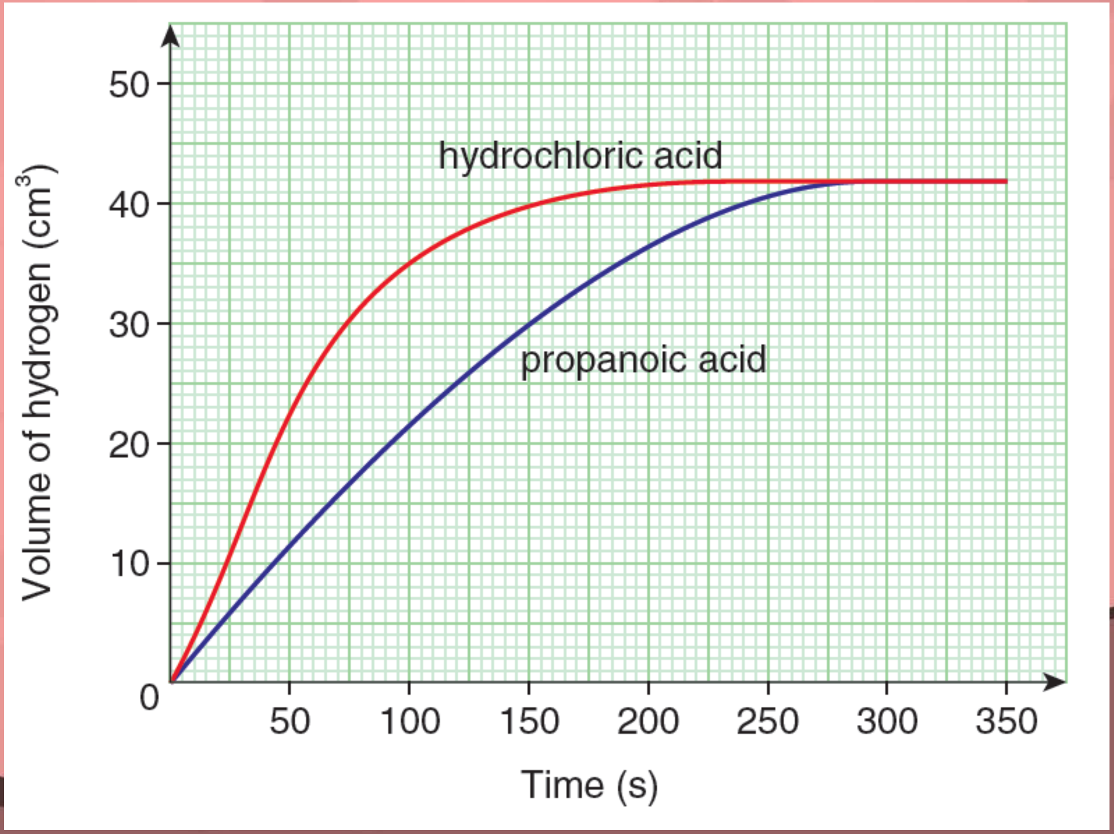
37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

a) Compare *Trials 1* and *3*.
The time becomes half when the concentration of X(g) increases.
Thus, increasing the concentration of X(g) increases the rate of the reaction.

b) Compare *Trials 1* and *2*.
The time remains unchanged when the concentration of Y(g) increases.
Thus, increasing the concentration of Y(g) has NO effect on the rate of the reaction.

37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

2 A student compared the rates of reactions of magnesium with hydrochloric acid and propanoic acid of the same concentration. The magnesium was in excess in both trials. The student measured the volume of hydrogen released at regular time intervals. The graph below shows the results.



37.2 Effect of change in concentration of a reactant on the rate of a reaction (p.32)

a) State TWO factors that should be kept constant in the two trials for a fair comparison.

Any two of the following:

- Temperature
- Mass of magnesium
- Particle size / surface area of magnesium
- Volume of acid used

b) The two curves have different shapes because the strength of each acid is different. Write about the difference between a strong acid and a weak acid, and explain why the two curves are different.

A strong acid is an acid that dissociates almost completely to give hydrogen ions in water.

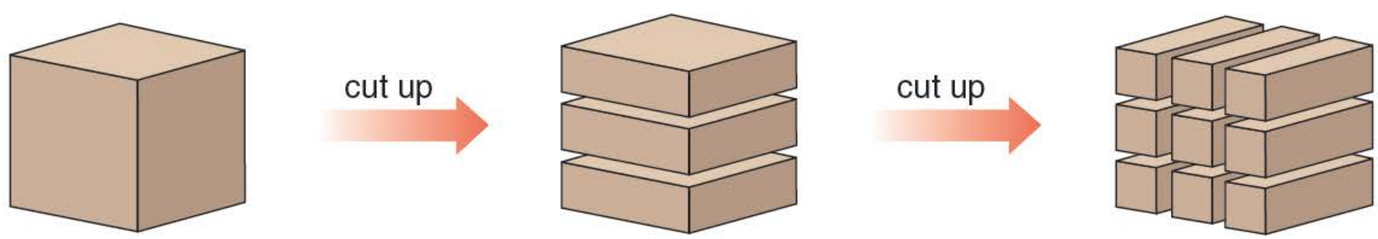
A weak acid is an acid that dissociates to give hydrogen ions partially in water.

Hydrochloric acid reacts faster than propanoic acid.

Hydrochloric acid has a greater concentration of hydrogen ions.

37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

- Have you ever lighted a camp fire? If you have, you will know that cutting a large wooden block into smaller sticks makes the combustion much faster.
- The surface area increases when the wooden block is cut into sticks. What effect will this have on the rate of the combustion?



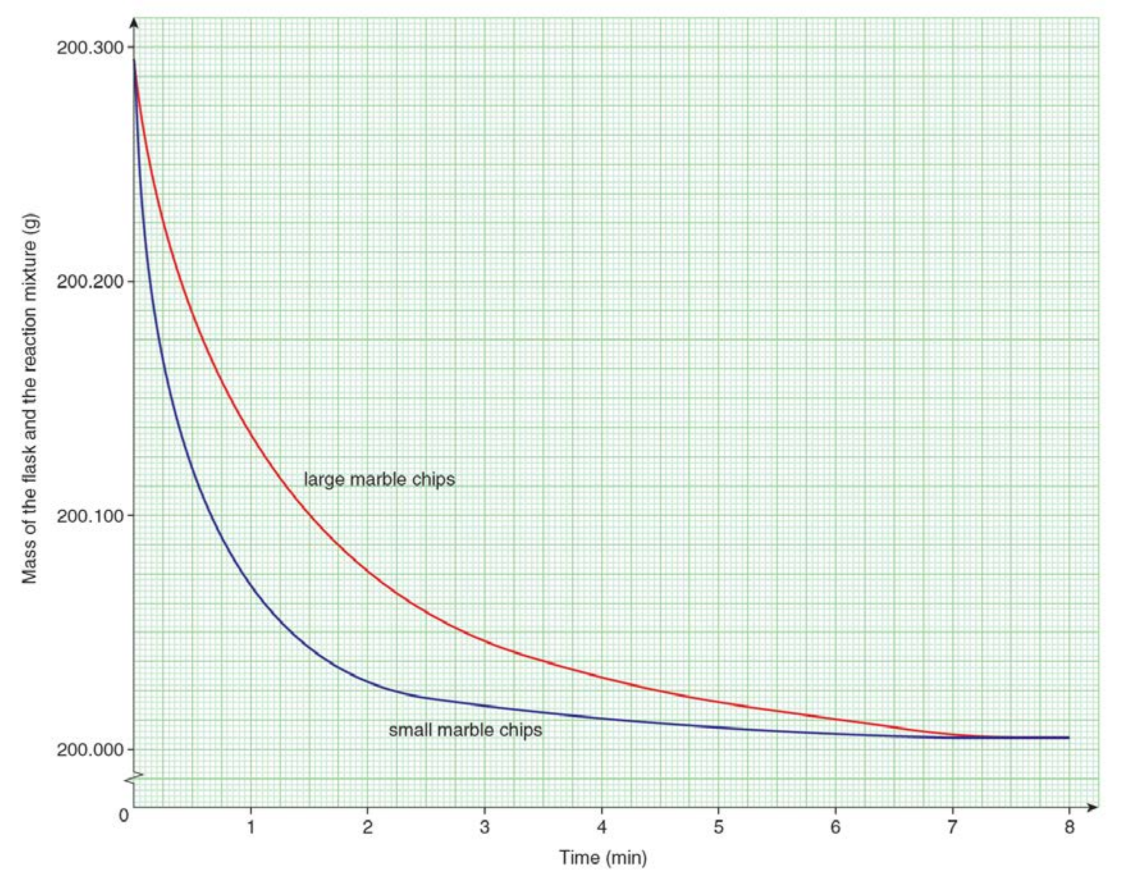
	1 large block	3 'slices'	9 'sticks'
Volume	= 3 cm x 3 cm x 3 cm = 27 cm^3	= 27 cm^3	= 27 cm^3
Surface area	= 6 x 3 cm x 3 cm = 54 cm^3	= 3(2 x 3 cm x 3 cm) + 3(4 x 3 cm x 1 cm) = 90 cm^3	= 9(2 x 1 cm x 1 cm) + 9(4 x 3 cm x 1 cm) = 126 cm^3

37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

- Consider again the reaction between dilute hydrochloric acid and marble (calcium carbonate) of different surface areas. The mass of the flask and the reaction mixture is monitored over 7 minutes using an electronic balance.
- One trial is carried out using a known mass of large marble chips with a known volume of 1.50 mol dm^{-3} hydrochloric acid.
- This is repeated using the same mass of small marble chips with the same volume of 1.50 mol dm^{-3} hydrochloric acid. The acid is excess in both cases.

37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

- The figure shows the variations of mass of the flask and the reaction mixture with time for the large and small marble chips.



37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

There are two important points about the curves:

- At the start, the slope of the curve for small marble chips is steeper than that for large marble chips. This means that the initial rate of reaction with small marble chips is higher. The small marble chips, with a greater surface area, react faster.
- The two curves flatten off at the same mass. This means that the same mass of carbon dioxide is being formed in both trials. This is because the mass of marble chips and the amount of acid are the same.

37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

The rate of a reaction increases when the surface area of a solid reactant is increased.

Investigating the effect of varying the surface area of marble chips on the rate of their reaction with dilute hydrochloric acid

37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

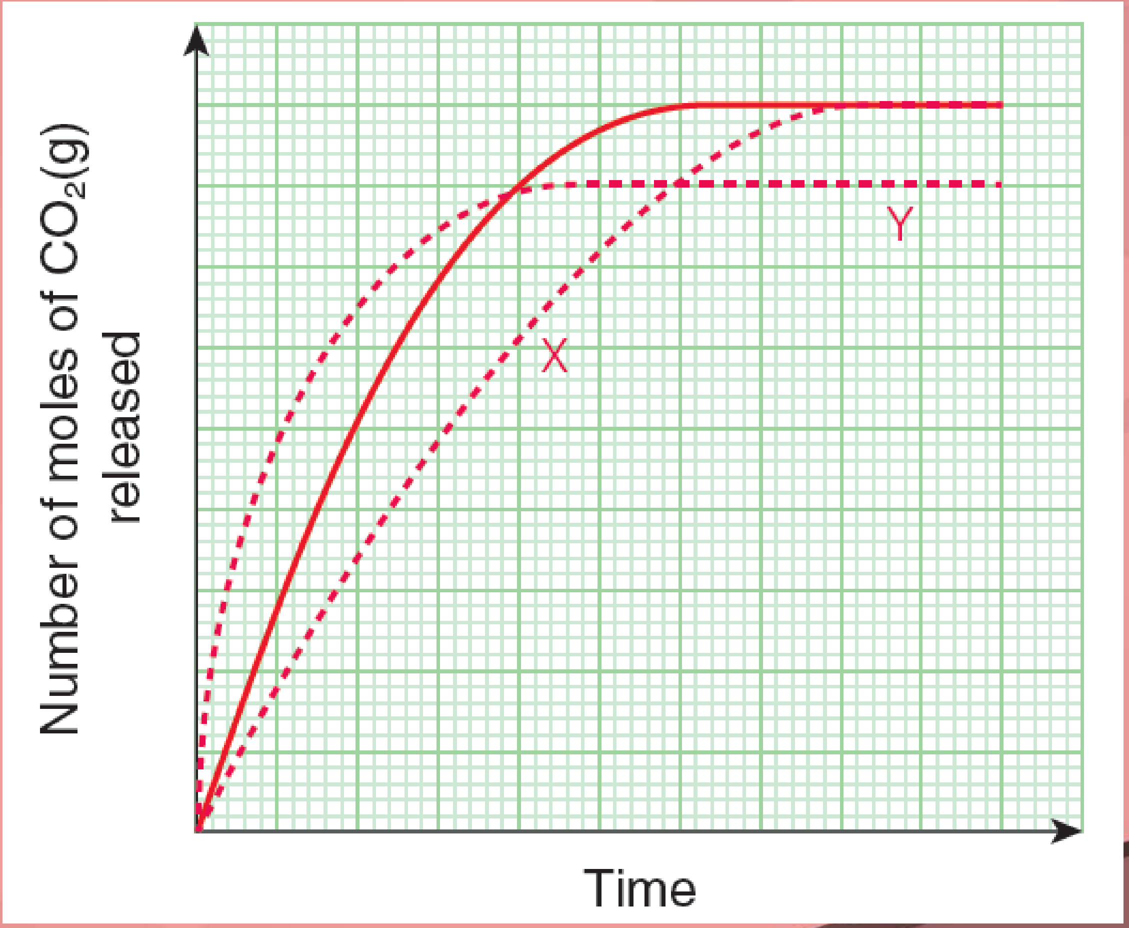
Practice 37.2

When ethanoic acid reacts with marble chips, the progress of the reaction can be followed by monitoring the number of moles of $CO_2(g)$ released with time. The curve below shows how the number of moles of $CO_2(g)$ released varies with time when 2.00 g of small marble chips (excess) are added to 30.0 cm^3 of 0.600 mol dm^{-3} ethanoic acid.

a) The reaction was repeated using 2.00 g of large marble chips and 30.0 cm^3 of 0.600 mol dm^{-3} ethanoic acid.
Sketch the curve to show how the number of moles of $CO_2(g)$ released varies with time in this second experiment. (Label this curve as X.)

b) The reaction was repeated using 2.00 g of small marble chips and 20.0 cm^3 of 0.800 mol dm^{-3} hydrochloric acid.
Sketch the curve to show how the number of moles of $CO_2(g)$ released varies with time in this third experiement. (Label this curve as Y.)

37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)



37.3 Effect of change in surface area of a solid reactant on the rate of a reaction (p.38)

Number of moles of CH_3COOH reacted

$= 0.600\ mol\ dm^{-3} \times \frac{30.0}{1\,000} dm^3$

= 0.0180 mol

Number of moles of HCl reacted

$= 0.800\ mol\ dm^{-3} \times \frac{20.0}{1\,000} dm^3$

= 0.0160 mol

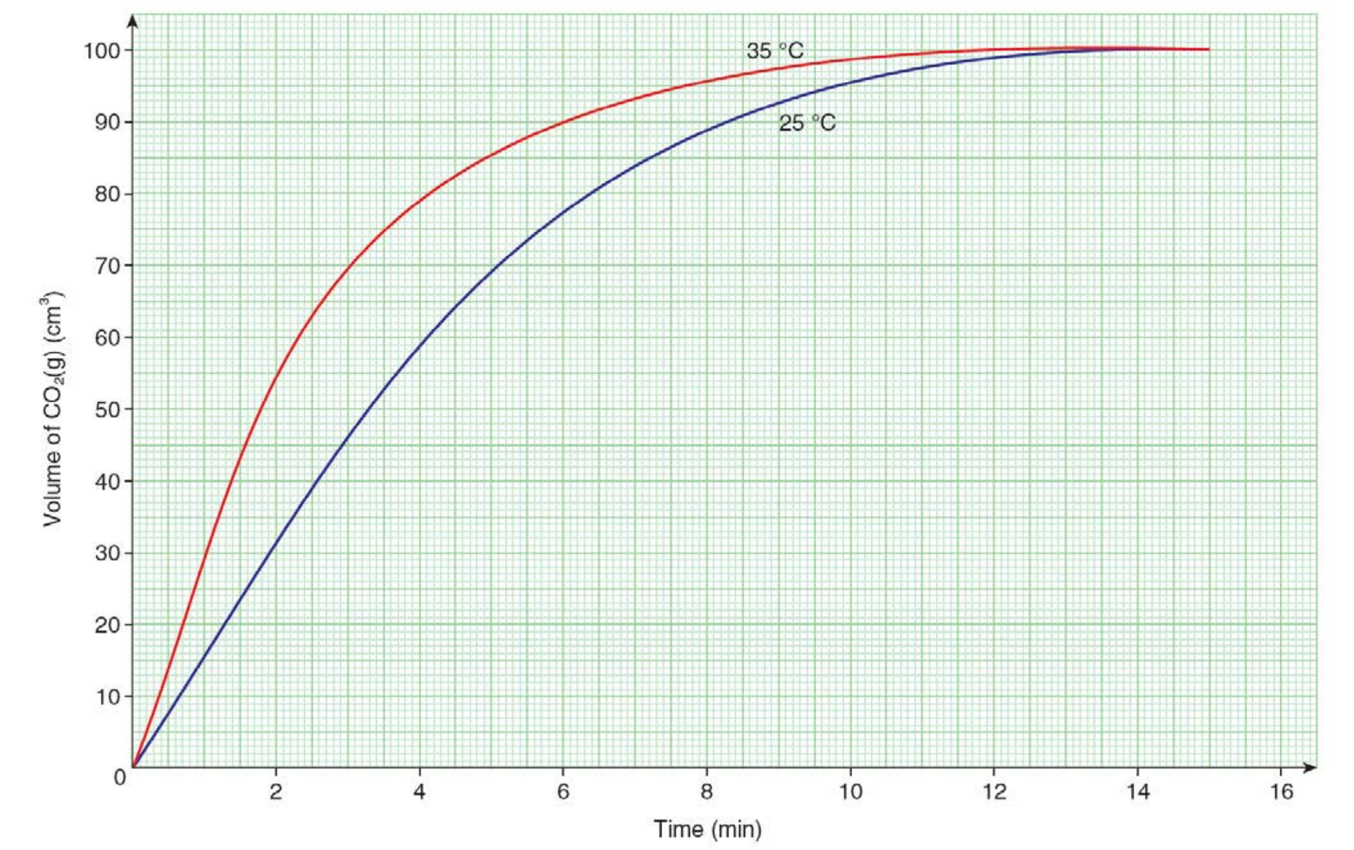
∴ less CO_2(g) was released in the third experiment.

37.4 Effect of change in temperature on the rate of a reaction (p.41)

- Consider the reaction between calcium carbonate and dilute hydrochloric acid of different temperatures. The volume of carbon dioxide gas released is monitored over 14 minutes by using a gas syringe.
- One trial is carried out using a known mass of calcium carbonate with an excess of 1.50 mol dm^{-3} hydrochloric acid at 25 °C.
- This is repeated using the same mass of calcium carbonate and the same volume of 1.50 mol dm^{-3} hydrochloric acid at 35 °C.

37.4 Effect of change in temperature on the rate of a reaction (p.41)

- The figure below shows the variations of volume of carbon dioxide released with time for the two different temperatures.



37.4 Effect of change in temperature on the rate of a reaction (p.41)

There are two important points about the curves:

- At the start, the tangent of the curve for 35 °C is steeper than that for 25 °C. This means that the initial rate of the reaction is higher at 35 °C.
- The two curves flatten off at the same volume of carbon dioxide released. This is because the same amounts of reactants are used in the two trials.

37.4 Effect of change in temperature on the rate of a reaction (p.41)

The rate of a reaction increases when the temperature is increased.

Effect of temperature on rate *Ref.*

Investigating the effect of varying the temperature on the rate of the reaction between sodium thiosulphate solution and dilute sulphuric acid *Ref.*

37.5 Effect of using a catalyst on the rate of a reaction (p.43)

- Hydrogen peroxide solution decomposes to form water and oxygen: $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
- You can follow the progress of this decomposition by monitoring the volume of oxygen gas released with time.
- In the first trial, a known volume of hydrogen peroxide solution is allowed to decompose.
- This is repeated by using the same volume of hydrogen peroxide solution of equal concentration <u>in the presence of a small amount of manganese(IV) oxide</u>.

37.5 Effect of using a catalyst on the rate of a reaction (p.43)

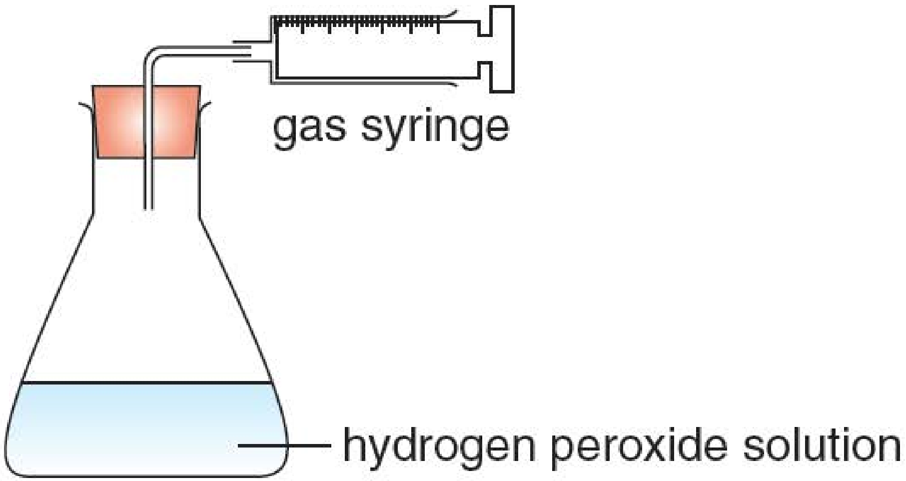


Fig. 37.7 *Experimental set-up for following the progress of the decomposition of hydrogen peroxide solution*

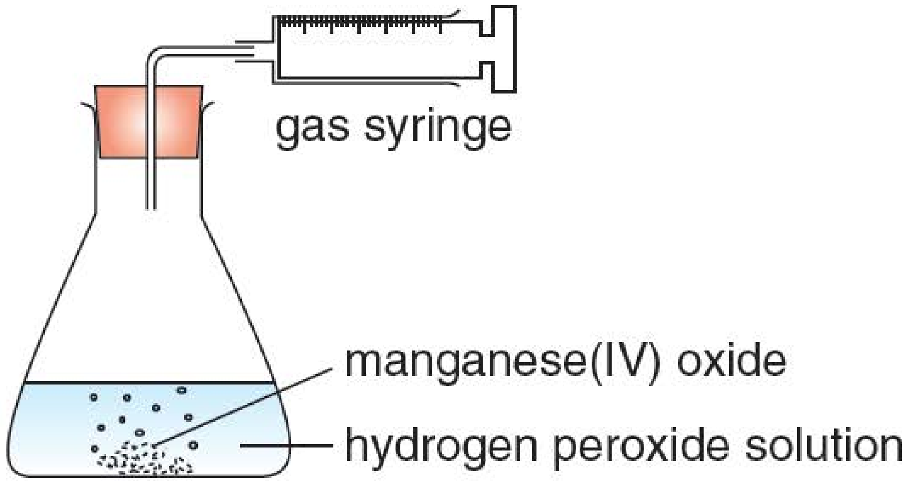
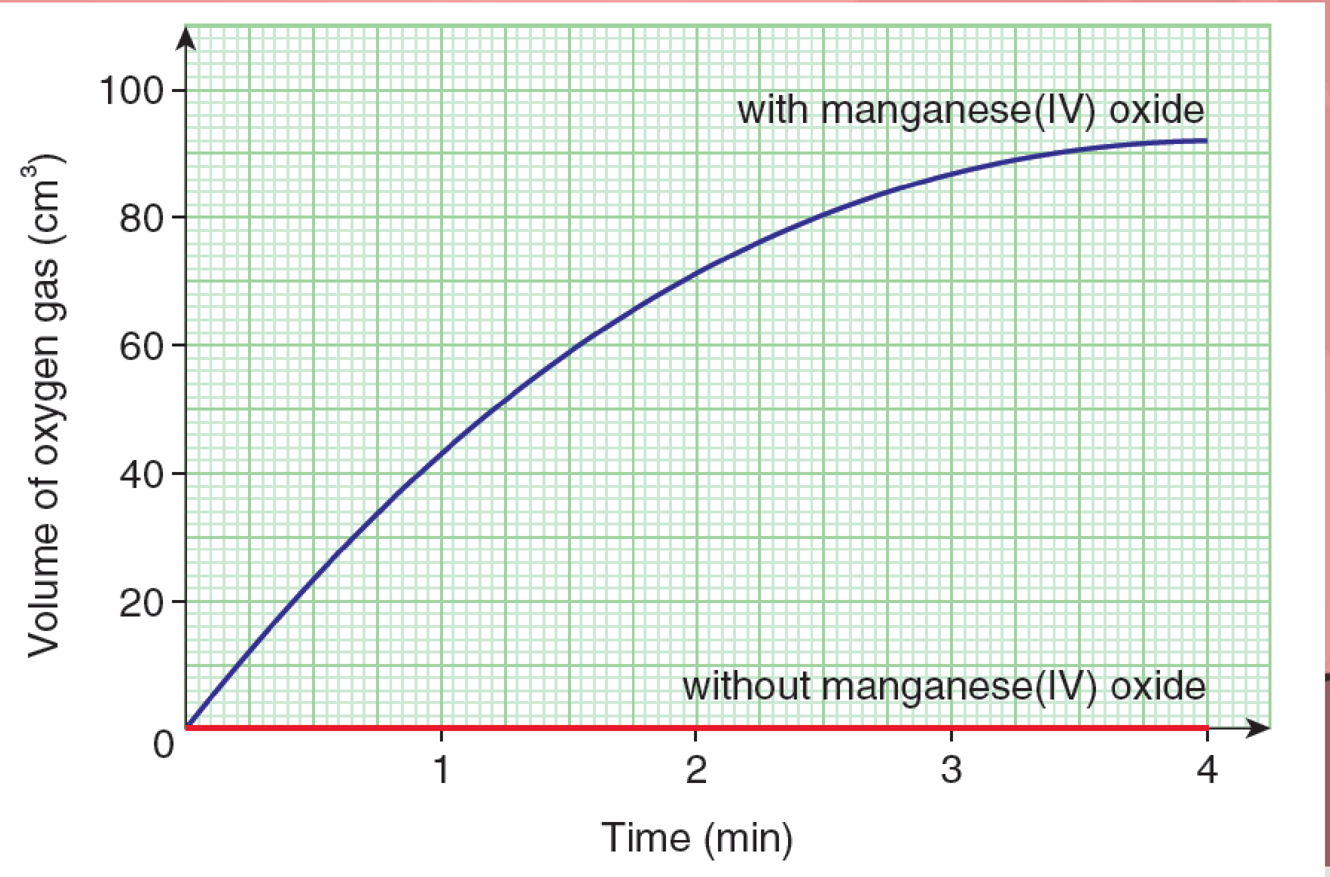


Fig. 37.8 *Experimental set-up for following the progress of the decomposition of hydrogen peroxide solution in the presence of manganese(IV) oxide*

37.5 Effect of using a catalyst on the rate of a reaction (p.43)

- The figure below shows the variations of volume of oxygen with time for the two trials.



37.5 Effect of using a catalyst on the rate of a reaction (p.43)

- At the start of the decomposition, the slope of tangent to the curve of the trial with manganese(IV) oxide is much steeper than that of the trial without manganese(IV) oxide. The initial rate of decomposition is much higher when manganese(IV) oxide is present.

- Manganese(IV) oxide is a **catalyst (催化劑)** for the decomposition.

A catalyst is a substance that can change (increase or decrease) the rate of a reaction but remains chemically unchanged at the end of the reaction.

37.5 Effect of using a catalyst on the rate of a reaction (p.43)

- There are two types of catalysts: *positive catalysts* and *negative catalysts*. A positive catalyst is one that speeds up a reaction while a negative catalyst is one that slows down a reaction.
- The black powder of manganese(IV) oxide does not disappear during the course of the decomposition. If the solid is filtered and dried after the end of the decomposition, the same mass of powder remains.

Catalysing the decomposition of hydrogen peroxide in solution

37.6 Factors affecting the rate of a reaction (p.44)

- Now, there are a number of factors which can affect the rate of a reaction:
  - concentration of a reactant;
  - surface area of a solid reactant;
  - temperature; and
  - use of a catalyst.
- To understand why the rate changes, you need to think about reactions in terms of the particles involved. The collision theory states that two reacting particles must collide for a reaction to occur. Usually, only a small percentage of collisions results in a reaction.

37.6 Factors affecting the rate of a reaction (p.44)

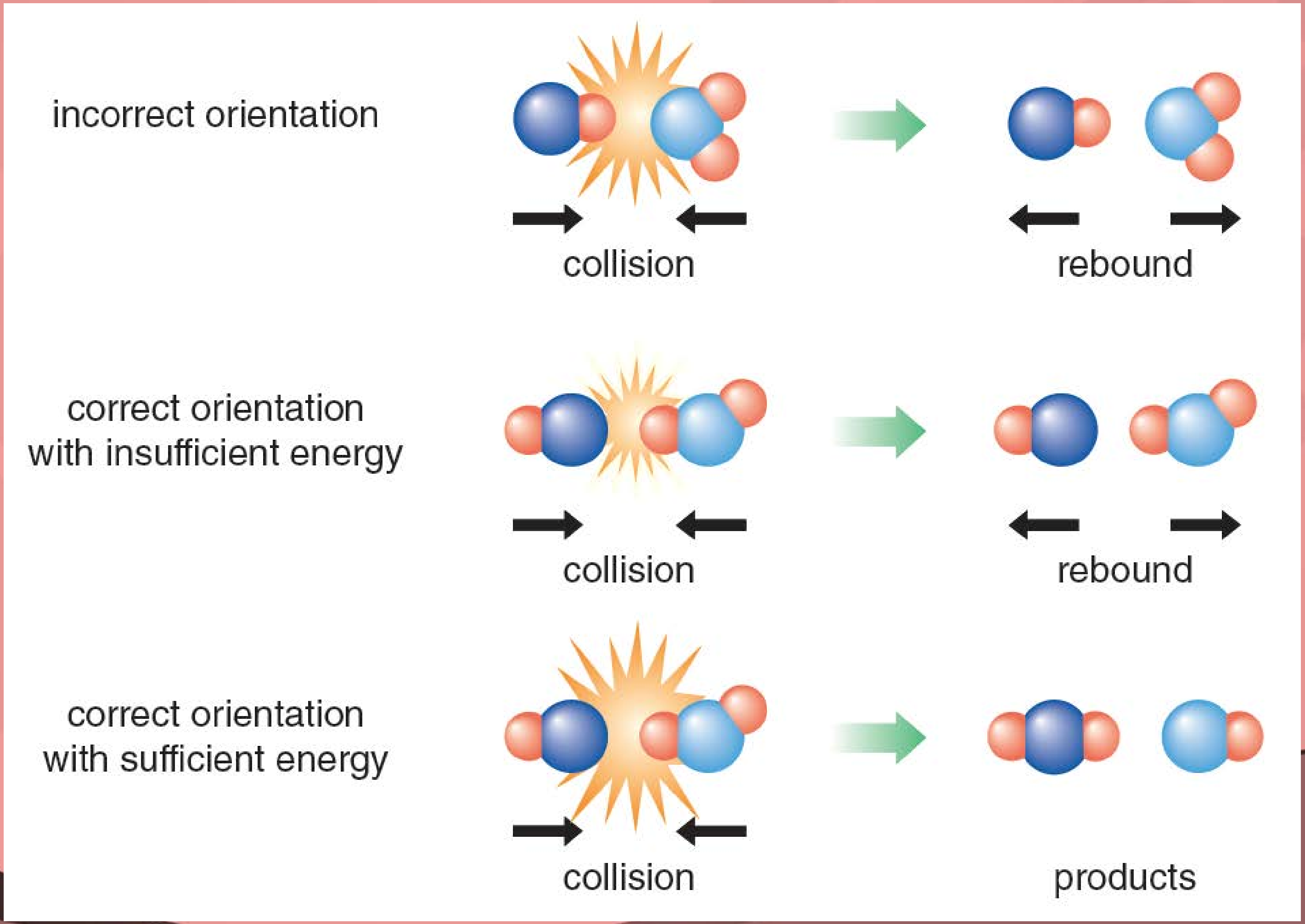
Why are some collisions effective but other ineffective?

- An **effective collision (有效碰撞)** is one that leads to a reaction. The rate of a reaction depends on how many effective collisions there are in a unit volume per unit time.
- A collision will be effective if two conditions have been met:
  - the particles collide with the correct orientation; and
  - the particles collide with sufficient energy for bonds to break. The minimum amount of energy required is called the **activation energy (活化能)**.

37.6 Factors affecting the rate of a reaction (p.44)

- For a reaction to happen, some bonds must be broken and this requires energy. The colliding particles must have an amount of kinetic energy equal to or greater than the activation energy.
- The energy is required to overcome the repulsion between electron clouds of the particles, and to start rearranging the bonds to form the product(s).
- Each reaction has its own different value of activation energy.

37.6 Factors affecting the rate of a reaction (p.44)



37.7 Why does increasing the concentration of reactant increase the rate of a reaction? (p.46)

- At the molecular level, the rate of a reaction depends on two factors:
  - how often the reactant particles hit one another; and
  - what percentage of the collisions have kinetic energy equal to or greater than the activation energy.

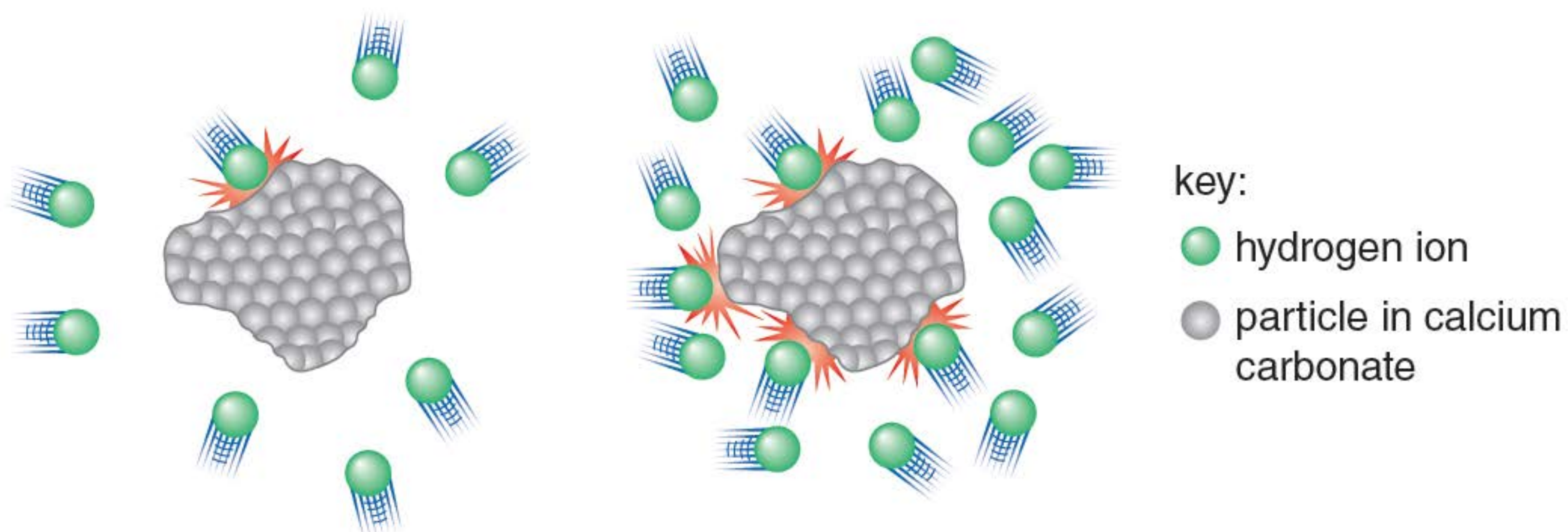
37.7 Why does increasing the concentration of reactant increase the rate of a reaction? (p.46)

- The reaction between calcium carbonate and dilute hydrochloric acid involves collisions between particles in calcium carbonate and hydrogen ions from the acid.
- The figure below shows the reaction of same size and same mass of calcium carbonate and same volume of hydrochloric acid of different concentrations.

37.7 Why does increasing the concentration of reactant increase the rate of a reaction? (p.46)

- Increasing the concentration of the acid means increasing the number of hydrogen ions per unit volume. The hydrogen ions are more crowded and collide more often with particles on the surface of the calcium carbonate.
- The chance of collision increases, so there will be more effective collisions in a unit volume per unit time. Hence the rate of the reaction increases.

37.7 Why does increasing the concentration of reactant increase the rate of a reaction? (p.46)



slow reaction of calcium carbonate with hydrochloric acid of low concentration

fast reaction of calcium carbonate with hydrochloric acid of high concentration

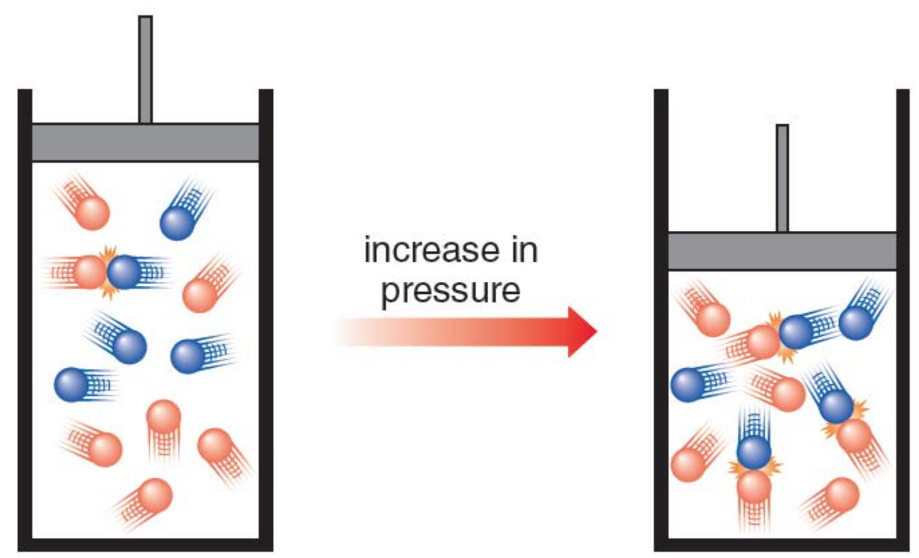
Effect of concentration on collisions *Ref.*

37.7 Why does increasing the concentration of reactant increase the rate of a reaction? (p.46)

- When one or more of the reactants are gases, an increase in pressure, at constant temperature, results in an increase in the rate of the reaction. Pressure can be increased at a given temperature by reducing the volume of the container. There is an increase in the number of gas particles per unit volume, so they collide more often.
- An increase in pressure can be regarded as an increase in 'concentration' since there are more gas particles per unit volume.

37.7 Why does increasing the concentration of reactant increase the rate of a reaction? (p.46)

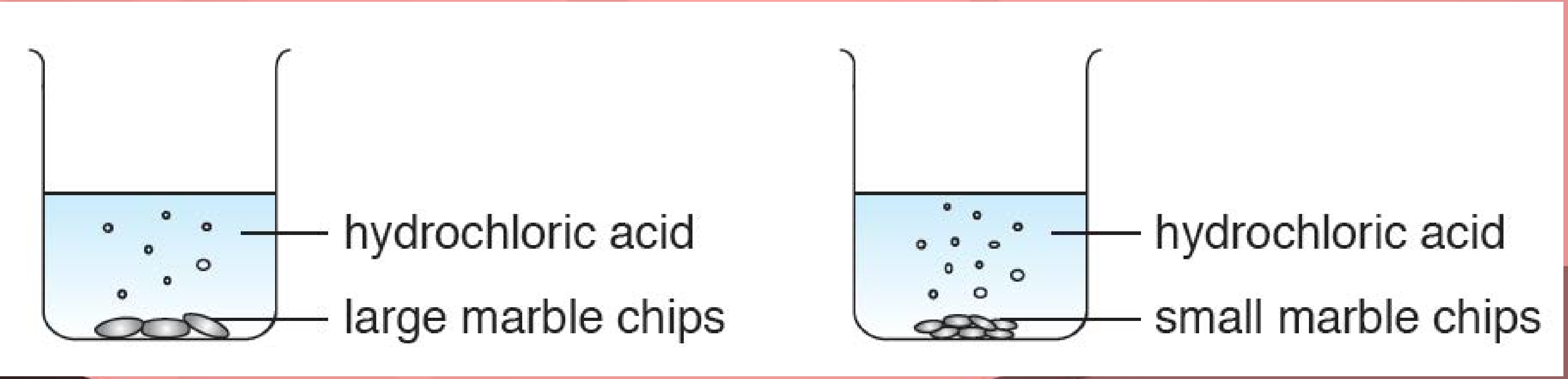
- The concentrations of the reactants decrease gradually during the course of a reaction. As a result, the rate also goes down with time. At the end of the reaction, at least one of the reactants has been used up completely, so the rate is zero.



Increasing the pressure results in more frequent collisions

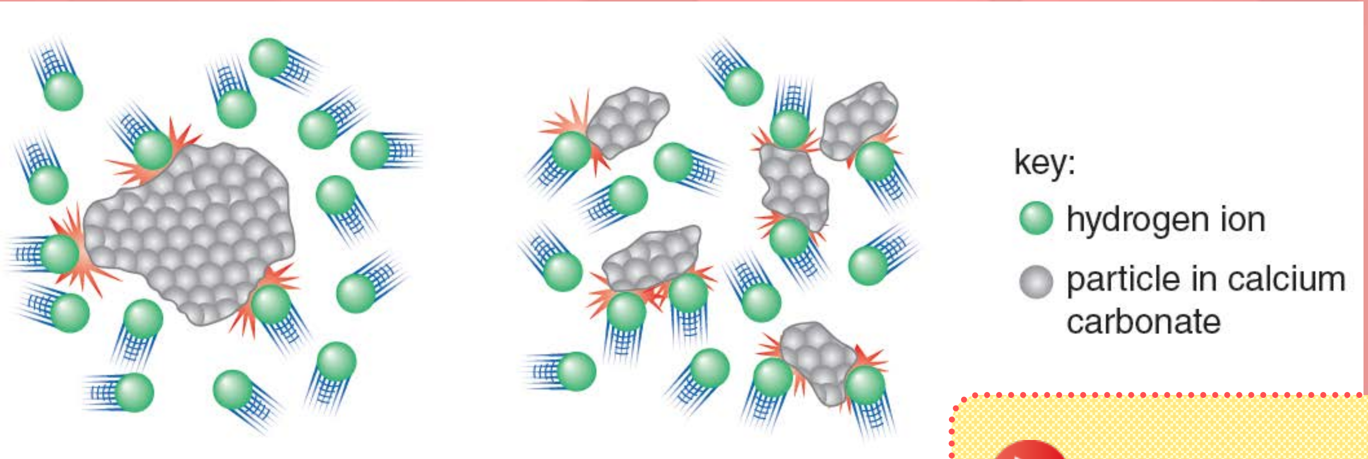
37.8 Why does increasing the surface area of a solid reactant increase the rate of a reaction? (p.47)

- The figure below shows the reactions of the same volume and concentration of hydrochloric acid with large marble chips and small marble chips of the same mass.
- Only collisions between hydrogen ions and particles on the surface of the marble chips can result a reaction.



37.8 Why does increasing the surface area of a solid reactant increase the rate of a reaction? (p.47)

- When a large marble chip is broken up into smaller pieces, the surface area is increased, giving a greater area over which collisions can occur, so there are more effective collisions in a unit volume per unit time. Hence the rate of the reaction increases.



large marble chip (slow reaction)

small marble chips (fast reaction)

Effect of surface area on collisions *Ref.*

37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

When particles in gases, liquids or solutions are heated, they gain more kinetic energy and move faster. This has two results:

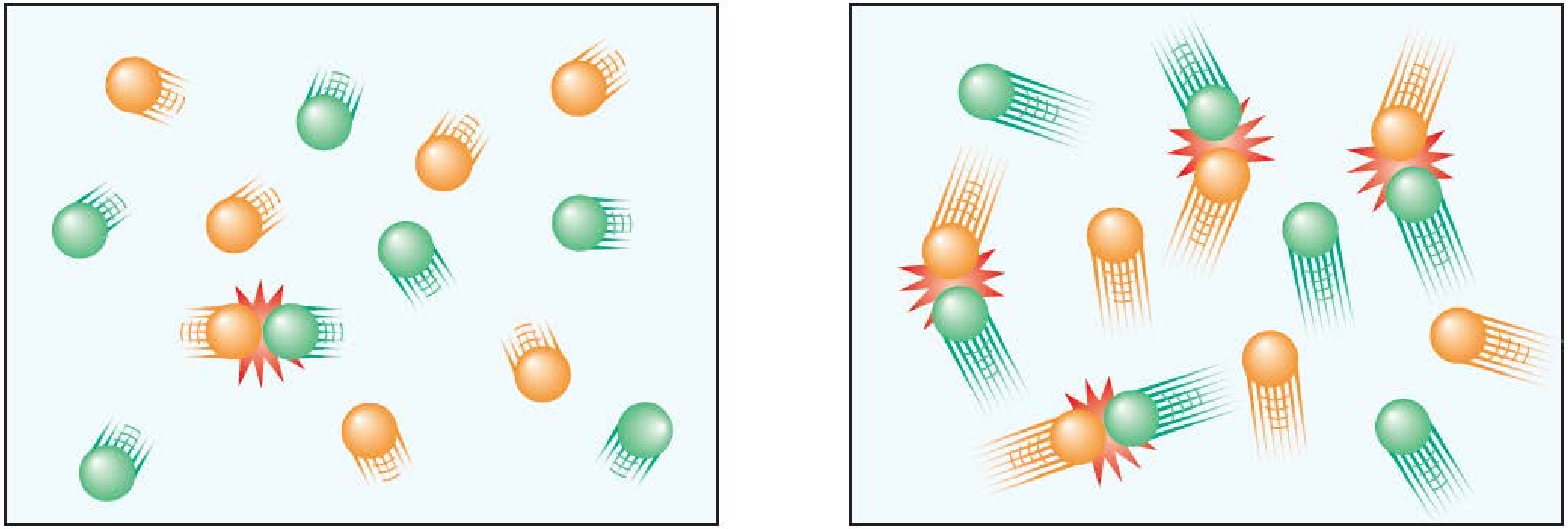
- the particles travel a greater distance per unit time and so will be involved in more collisions, resulting in more effective collisions and hence cause an increase in rate;
- more importantly, a larger fraction of the colliding particles will have kinetic energy equal to or greater than the activation energy, and hence a higher percentage of collisions can result a reaction.

37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

- So there will be more effective collisions in a unit volume per unit time with temperature increases. Hence the rate of the reaction increases.
- Generally, a rise of 10 °C approximately doubles the initial rate of a reaction.

The main reason that the rate of a reaction increases with temperature is an increase in the fraction of colliding particles with kinetic energy equal to or greater than the activation energy.

37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)



lower temperature — lower kinetic energy, slower movement, less collisions

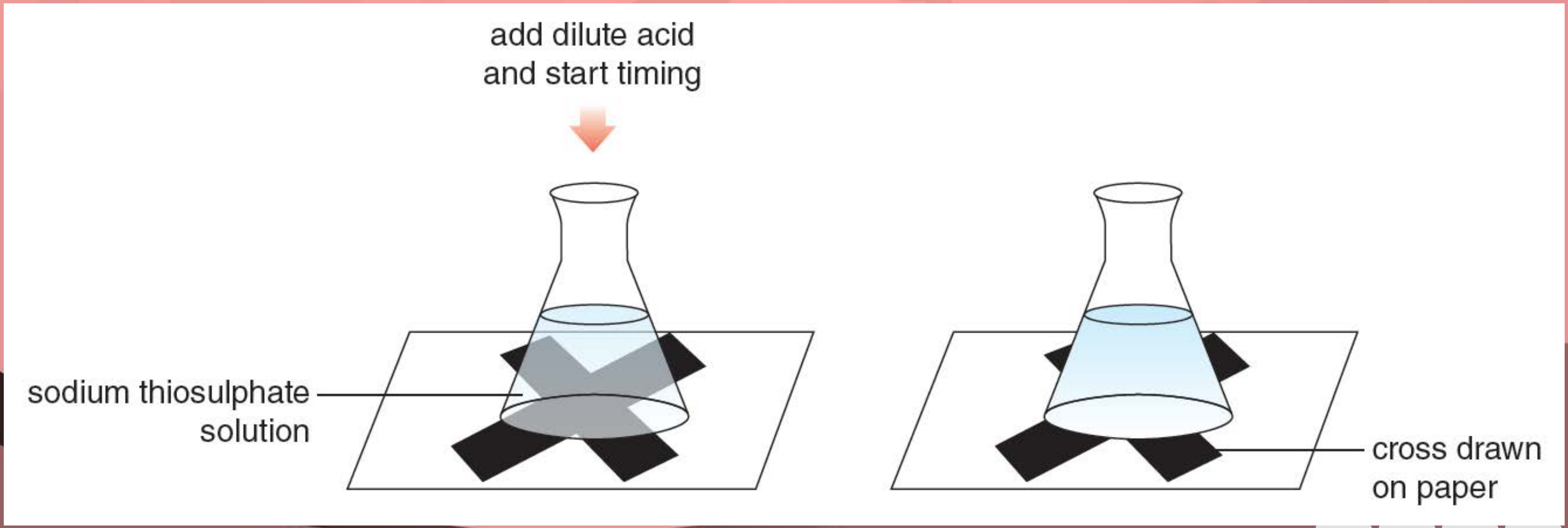
higher temperature — higher kinetic energy, faster movement, more collisions

Effect of temperature on rate *Ref.*

37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

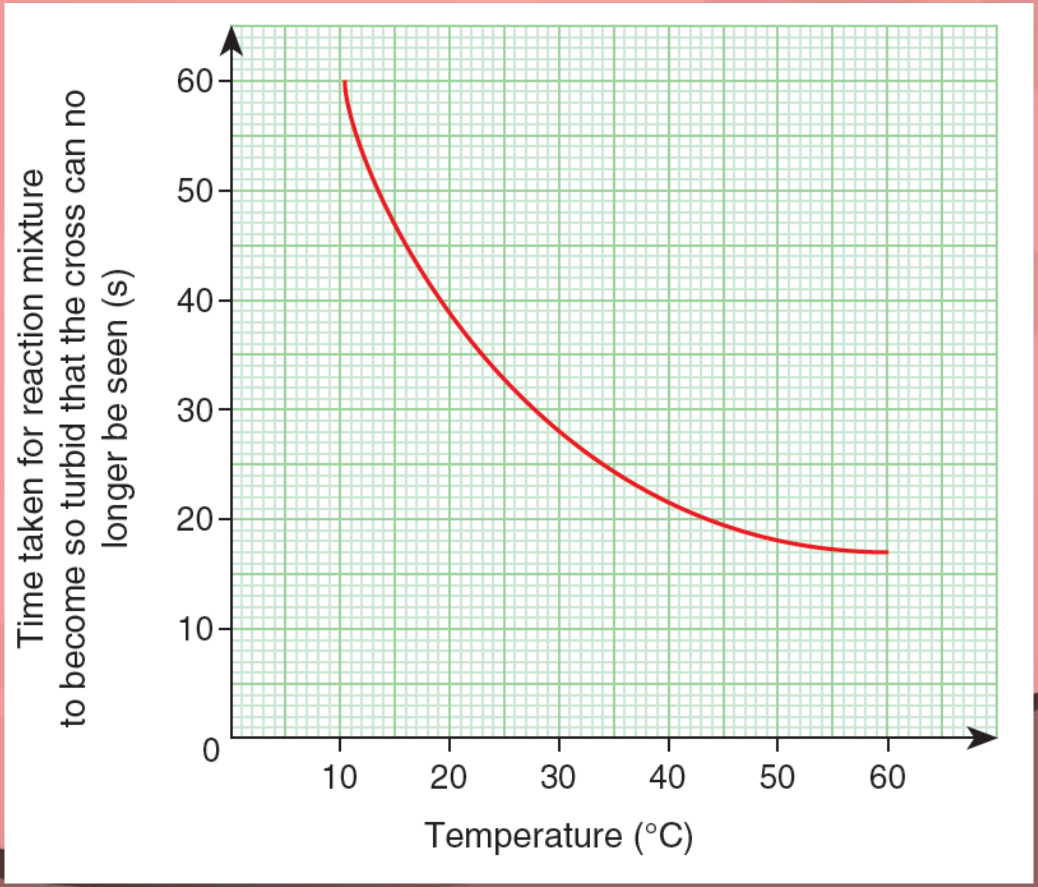
Practice 37.3

When sodium thiosulphate solution reacts with dilute acid, sulphur forms as a precipitate. The precipitate causes the reaction mixture to go cloudy. The rate of the reaction can be followed by timing how long it takes for the reaction mixture to become so turbid that the cross can no longer be seen from above.



37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

A student studied the effect of temperature on the reaction rate. The results obtained were shown in the graph below.



37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

a) What conclusion can you draw from the graph? Explain how temperature affects the rate of reaction in terms of the behaviour of particles.

As the temperature increases, the time taken for the cross no longer seen from above decreases. / As the temperature increases, the rate of reaction increases.

At a higher temperature, the reactant particles have more kinetic energy and move faster:

- the reactant particles involve in more collisions, resulting in more effective collisions;
- a larger fraction of the colliding particles will have kinetic energy equal to or greater than the activation energy and hence a higher percentage of collisions can result in a reaction.

37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

b) A second student carried out the same experiment using a higher concentration of acid.

i) Draw the curve you would expect him to obtain on the same grid.

The curve must be below the original one and steeper.

ii) Explain how the concentration of acid affects the rate of reaction in terms of the behaviour of particles.

Increasing the concentration of the acid means increasing the number of hydrogen ions per unit volume.

The hydrogen ions are more crowded and collide more often with the thiosulphate ions. So there are more effective collisions in a unit volume per unit time.

37.9 Why does increasing the temperature increase the rate of a reaction? (p.20)

c) Another student suggested using a light sensor and data-logger to study the rate of the above reaction.
Give ONE advantage of using a data-logger with a light sensor.

Continous readings / graph of light transmittance of reaction mixture against time plotted automatically.

37.10 Why does using a catalyst increase the rate of a reaction? (p.50)

- The left figure shows a hydrogen peroxide solution without adding any substance. The right figure shows the same solution with a little manganese(IV) oxide added at room temperature.
- The decomposition of hydrogen peroxide solution without any addition is very slow at room temperature. However, the addition of a positive catalyst, manganese(IV) oxide, makes the decomposition go much faster.

37.10 Why does using a catalyst increase the rate of a reaction? (p.50)

- A positive catalyst does not increase the number of collisions between reactant particles per unit time, nor does it make collisions more energetic.
- A positive catalyst works by providing an alternative reaction pathway of lower activation energy. Lowering the activation energy increases the fraction of molecules with enough energy to react, so there are more effective collisions in a unit volume per unit time. Hence the rate of reaction increases.

37.10 Why does using a catalyst increase the rate of a reaction? (p.50)

- A catalyst is usually specific in its action. A chemical that works as a catalyst for one reaction does not necessarily catalyse a different reaction.
- A catalyst does not get used up in the reaction, so a tiny amount of catalyst can be used to speed up a reaction over and over again.

37.11 Catalysts in industry (p.51)

- Catalysts increase the rate of many industrial chemical reactions by lowering the activation energy. As a result, the temperature needed for the processes and the energy requirements are reduced.

Table 37.1 The catalysts used in some chemical processes

Chemical process	Catalyst(s)
Haber process — for the manufacture of ammonia	iron
Contact process — for the manufacture of sulphuric acid	vanadium(V) oxide
Production of nitric acid	platinum and rhodium
Turning oils into fats in the production of margarine	nickel

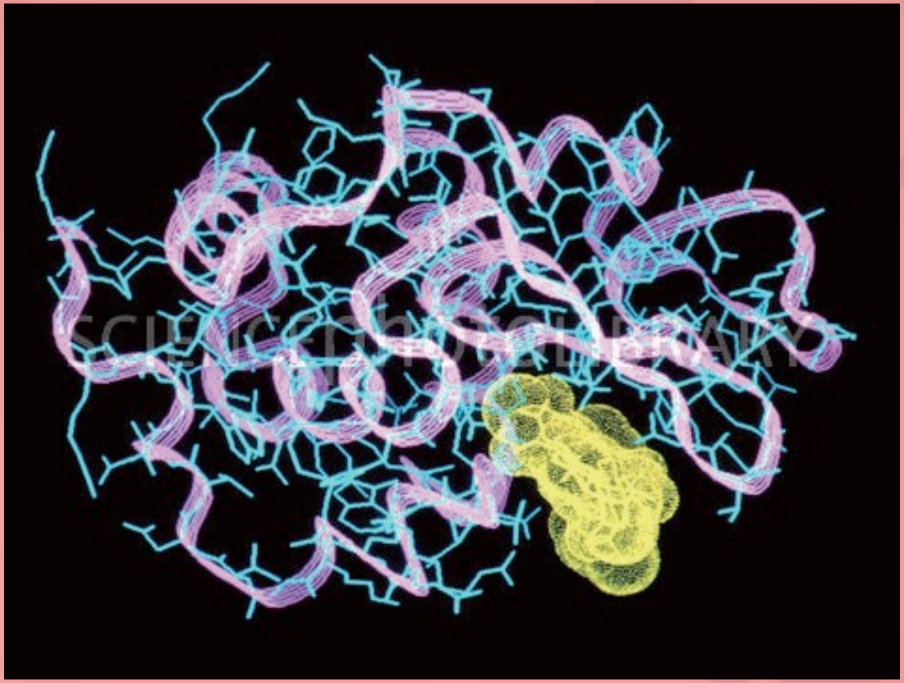
37.11 Catalysts in industry (p.51)

- Catalysts are normally used in the forms of powders, pellets or gauzes. These forms give them the biggest possible surface area, making them as effective as possible as the reactions they catalyse often involve gases reacting on their surfaces.

37.12 Enzymes (p.52)

- Chemical reactions that take place in living cells are very fast, even though they take place at just above room temperature and at low concentrations. This is because they are catalysed by efficient catalysts called **enzymes (酶)**.
- Enzymes are large protein molecules. Every living cell contains at least 1 000 different enzymes, each catalysing one of the many different chemical reactions that take place inside the cell. Some enzymes are so efficient that one molecule can catalyse the reaction of 10 000 reactant molecules every second.

37.12 Enzymes (p.52)



A computer image of an enzyme and a smaller reactant molecule

37.12 Enzymes (p.52)

Enzymes in industry

- Enzymes are being used increasingly as catalysts in industry. Biological laundry detergent uses enzymes to remove biological stains caused by sweat, blood and food. The enzymes in these detergents break down proteins and fats.
- Enzymes could transform the chemical industry of the future. Instead of using high temperatures and pressures to make reactions go quickly, chemists find that enzymes can do the job just as well at about room temperature.

37.12 Enzymes (p.52)

Biological laundry detergent

Key terms (p.54)

catalyst	催化劑	activation energy	活化能
collision theory	碰撞理論	enzyme	酶
effective collision	有效碰撞		

Summary (p.55)

1 An effective collision is one that leads to a reaction.

2 A collision will be effective if two conditions have been met:
- the particles collide with the correct orientation; and
- the colliding particles have an amount of kinetic energy equal to or greater than the activation energy.

3 A catalyst is a substance that can change (increase or decrease) the rate of a reaction but remains chemically unchanged at the end of the reaction.

Summary (p.55)

4 The table below summarises the factors that increase the rate of a reaction.

	Explanation
Increasing the concentration of a reactant	this increases the number of reactant particles per unit volume; the chance of collision increases, so there are more effective collisions in a unit volume per unit time
Increasing the surface area of a solid reactant	this gives a greater area over which collisions can occur, so there are more effective collisions in a unit volume per unit time
Increasing the temperature	When particles in gases, liquids or solutions are heated, • the particles travel a greater distance per unit time and so will be involved in more collisions, resulting in more effective collisions; • a larger fraction of the colliding particles will have kinetic energy equal to or greater than the activation energy, and hence a higher percentage of collisions will result in reaction.
Using a catalyst	a positive catalyst provides an alternative reaction pathway of lower activation energy; this increases the fraction of molecules with enough energy to react

Summary (p.55)

5 Enzymes are protein-based catalysts found in living things.

Unit Exercise (p.56)

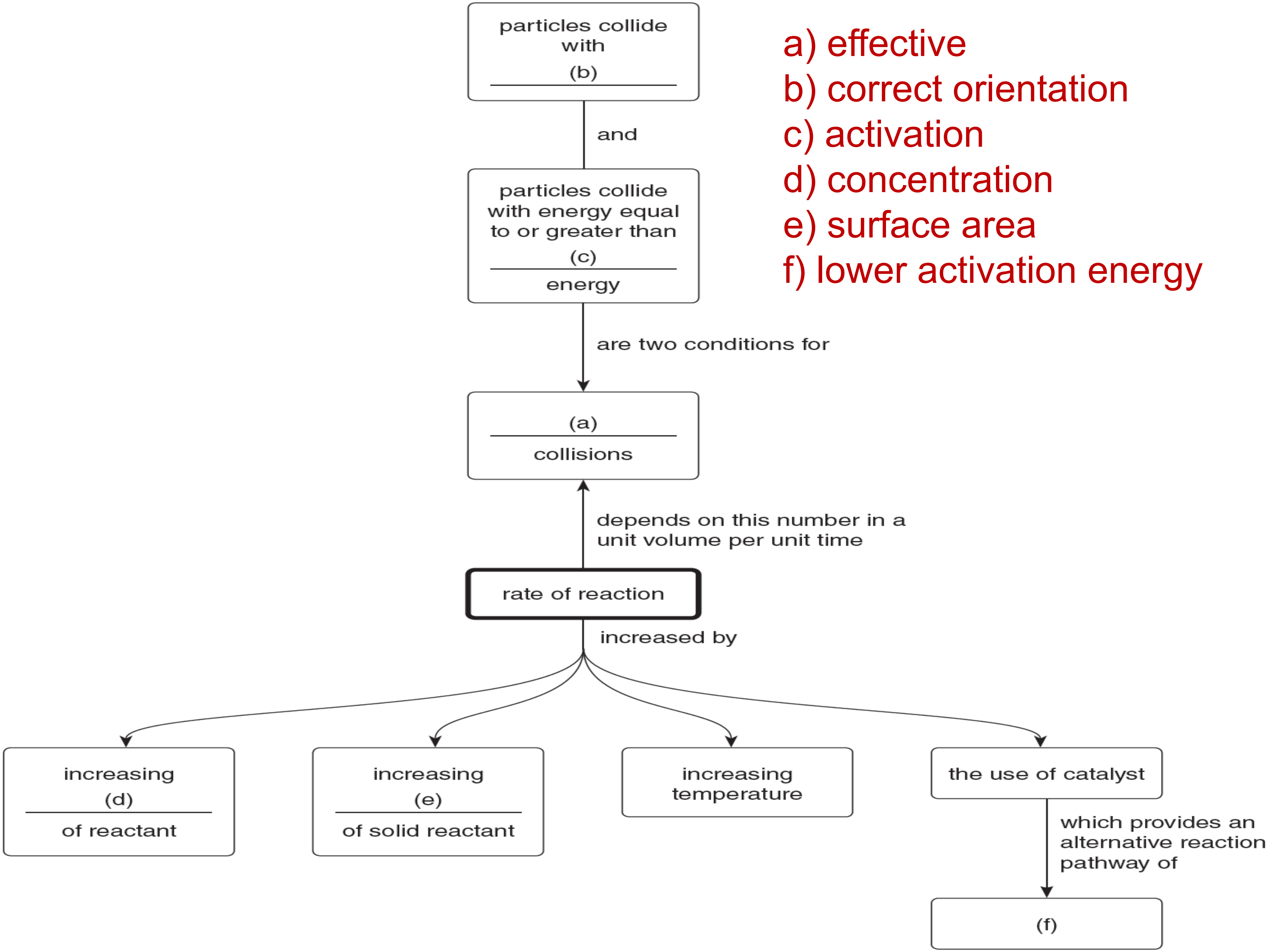
Note: Questions are rated according to ascending level of difficulty (from 1 to 5):

question targeted at level 3 and above;
question targeted at level 4 and above;
question targeted at level 5.
‘ * ’ indicates 1 mark is given for effective communication.

Unit Exercise (p.56)

PART I KNOWLEDGE AND UNDERSTANDING

1 Complete the following concept map.



a) effective
b) correct orientation
c) activation
d) concentration
e) surface area
f) lower activation energy

Unit Exercise (p.56)

PART II MULTIPLE CHOICE QUESTIONS

2 Consider the following reaction:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g) \qquad \Delta H < 0$$

How could the rate of this reaction be increased?

A Reduce the pressure.
B Increase the volume.
C Remove some NO_2(g).
D Increase the temperature.

Answer: D

Unit Exercise (p.56)

3 Which of the following pairs of chemicals, upon mixing under the same temperature, has the HIGHEST rate of gas formation?

A 0.10 g of Zn powder and 100 cm^3 of 1.0 M HCl(aq)
B 0.10 g of Zn granules and 200 cm^3 of 1.0 M HCl(aq)
C 0.10 g of Zn granules and 200 cm^3 of 1.0 M H_2SO_4(aq)
D 0.10 g of Zn powder and 100 cm^3 of 1.0 M H_2SO_4(aq)

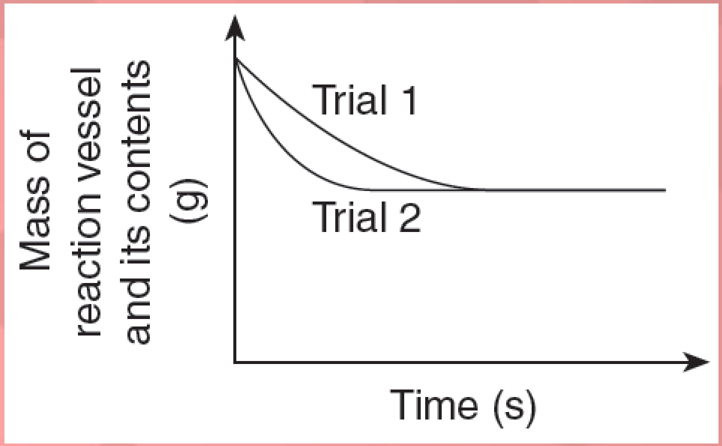
(HKDSE, Paper 1A, 2015, 28)

Answer: D

Unit Exercise (p.56)

4 The solution of chemical X decomposes to give a gas. A student carried out an experiment to test the effect of different factors on the rate of the decomposition.

In each trial, the mass of the flask and its contents was measured at regular time intervals. The results of two trials were shown below.



The student used 50 cm^3 of 1 mol dm^{-3} X(aq) in *Trial 1*. In *Trial 2*, the student must have

A added a catalyst.

B used a lower temperature.

C used a less concentrated solution.

D doubled the volume of the solution.

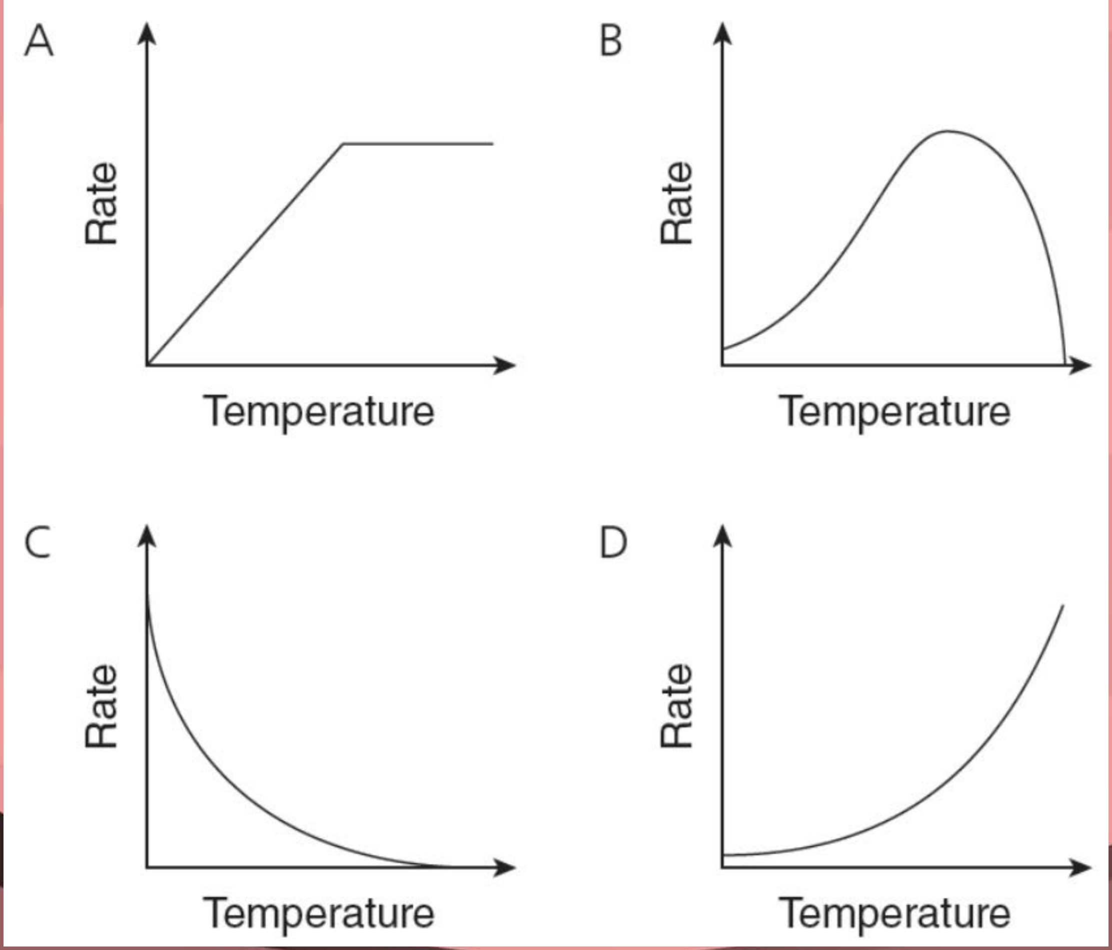
Answer: A

Explanation:

Adding a catalyst increases the rate of decomposition of X. The curve for *Trial 2* flattens off earlier.

Unit Exercise (p.56)

5 Which of the following graphs shows the effect of increasing temperature on the rate of reaction of magnesium with dilute sulphuric acid?

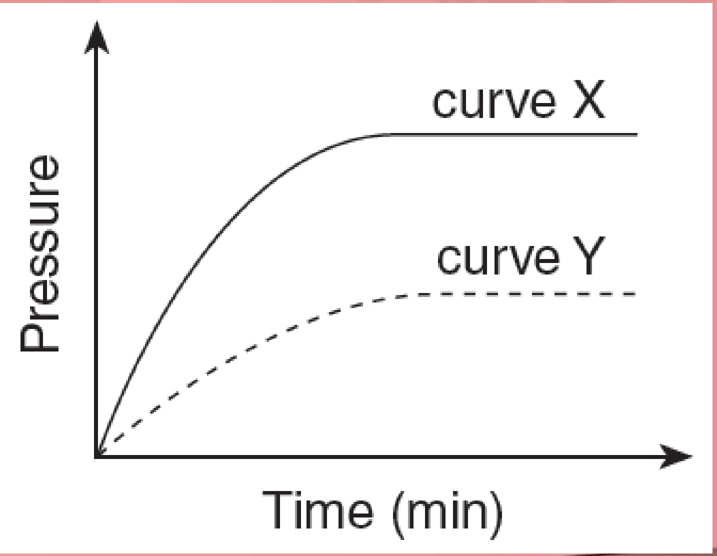


Answer: D

Unit Exercise (p.56)

6 A data-logger with pressure sensor was used to study the decomposition of 50.0 cm^3 of 1.00 mol dm^{-3} hydrogen peroxide solution in the presence of 0.5 g manganese(IV) oxide. Curve X shows the relation between the pressure and time.

The experiment was repeated. Curve Y shows the results obtained.



Which of the following changes would produce curve Y?

A Using 25.0 cm^3 of 1.00 mol dm^{-3} hydrogen peroxide solution

B Using 25.0 cm^3 of 2.00 mol dm^{-3} hydrogen peroxide solution

C Using 50.0 cm^3 of 0.500 mol dm^{-3} hydrogen peroxide solution

D Using 100.0 cm3 of 0.500 mol dm^{-3} hydrogen peroxide solution

Answer: C

Unit Exercise (p.56)

Explanation:

Hydrogen peroxide decomposes according to the equation below.
$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

When 50.0 cm^3 of 0.500 mol dm^{-3}, instead of 50.0 cm^3 of 1.00 mol dm^{-3}, $H_2O_2(aq)$ were used,

- the initial decomposition rate, and
- the total amount of oxygen gas formed would both decrease.

Unit Exercise (p.56)

7 Consider the following information of two reactions involving iron strips of the same shape and size.

Reaction	Mixture
1	1.0 g of Fe(s) + 50.0 cm^3 of 1.0 mol dm^{-3} H_2SO_4(aq)
2	1.0 g of Fe(s) + 25.0 cm^3 of 2.0 mol dm^{-3} H_2SO_4(aq)

Which of the following statements is correct?
(Relative atomic mass: Fe = 55.8)

A The initial rate of *Reaction 1* is higher than that of *Reaction 2*.
B Sulphuric acid is the limiting reactant in *Reaction 2*.
C The iron reacts completely in *Reaction 1*.
D A greater volume of gas is produced in *Reaction 2*.

Answer: C

Unit Exercise (p.56)

Explanation:

Fe(s) and H_2SO_4(aq) react according to the equation below.

$$Fe(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + H_2(g)$$

Number of moles

$$\frac{1.0\ g}{55.8\ g\ mol^{-1}} = 0.018\ mol \qquad 1.0\ mol\ dm^{-3} \times \frac{50.0}{1\ 000}\ dm^3 = 0.050\ mol$$

In *Reaction 1*, 0.018 mole of Fe reacts with 0.018 mole of H_2SO_4. Thus, Fe is the limiting reactant. The iron reacts completely in *Reaction 1*.

Unit Exercise (p.56)

8 The table lists the time taken for 1 g of zinc powder to react completely with the same volume of hydrochloric acid under different conditions.

Acid concentration (mol dm^{-3})	Temperature (^{o}C)	Reaction time (s)
0.1	20	100
0.1	25	80
0.2	30	40
0.2	40	20

What will be the reaction time for 1 g of zinc powder to react completely with 0.1 mol dm^{-3} hydrochloric acid at 30 $^{\circ}C$?

A Less than 20 s

B Between 40 s and 80 s

C Between 80 s and 100 s

D More than 100 s

Answer: B

Unit Exercise (p.56)

Explanation:

The reaction time between 1 g of zinc powder and 0.1 mol dm^{-3} HCl(aq) at 30 °C is shorter than that between the same reactants at 25 °C, i.e. less than 80 s.

The reaction time between 1 g of zinc powder and 0.1 mol dm^{-3} HCl(aq) at 30 °C is longer than that between 1 g of zinc powder and 0.2 mol dm^{-3} HCl(aq) at the same temperature, i.e. more than 40 s.

Unit Exercise (p.56)

9 Which of the following statements concerning a catalyst is correct?

A Its mass remains unchanged at the end of the reaction.
B It increases the energies of the reactant particles.
C It must have the same physical state as the reactants.
D It makes the reactant particles collide more often.

Answer: A

Unit Exercise (p.56)

10 In an experiment, 10 g of zinc carbonate are added to 100 cm^3 of 1 mol dm^{-3} HCl(aq) in a beaker. Which of the following changes can increase the initial rate of the reaction?

(1) Using 200 cm^3 of 1 mol dm^{-3} HCl(aq) to replace 100 cm^3 of 1 mol dm^{-3} HCl(aq)

(2) Using 50 cm^3 of 1 mol dm^{-3} H_2SO_4(aq) to replace 100 cm^3 of 1 mol dm^{-3} HCl(aq)

(3) Using 10 g of zinc carbonate of greater size instead

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

Answer: B

Explanation:
(2) HCl(aq) is a monobasic acid while H_2SO_4(aq) is a dibasic acid.
The initial rate increases as the concentration of H^+(aq) ions increases when 1 mol dm^{-3} H_2SO_4(aq) replaces 1 mol dm^{-3} HCl(aq).

Unit Exercise (p.56)

11 The collision between reactant particles is not necessarily successful. The collision may not lead to a reaction.

Which of the following are the reasons?

(1) The reactant particles are not energetic enough.

(2) The collision does not proceed in a proper orientation.

(3) The concentrations of the reactants are not high enough.

A (1) and (2) only

B (1) and (3) only

C (2) and (3) only

D (1), (2) and (3)

Answer: A

Unit Exercise (p.56)

12 Which of the following statements are correct?

(1) Magnesium oxide dissolves faster in 1 M HCl(aq) than 1 M CH_3CO_2H(aq).

(2) Powdered marble dissolves faster in 1 M HCl(aq) than granular marble does.

(3) H_2O_2(aq) decomposes faster in the presence of MnO_2(s) than without MnO_2(s).

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Answer: D

(HKDSE, Paper 1A, 2016, 33)

Unit Exercise (p.56)

13 Consider the following two reactions:

Reaction	Mixture
(I)	1.0 g of Na_2CO_3(s) + 100.0 cm^3 of 1.0 M HCl(aq)
(II)	1.0 g of Na_2CO_3(s) + 100.0 cm^3 of 1.0 M CH_3COOH(aq)

Which of the following statements are correct if the two reactions are performed under the same experimental conditions?

(Relative atomic masses: C = 12.0, O = 16.0, Na = 23.0)

(1) The decrease in mass for the two reaction mixtures is the same.
(2) The initial rate of *Reaction (I)* is higher than that of *Reaction (II)*.
(3) The heat given out for the two reactions is the same.

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Answer: A

(HKDSE, Paper 1A, 2018, 33)

Unit Exercise (p.56)

PART III STRUCTURED QUESTIONS

14 Carbon burns in air according to the equation below:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

List FOUR ways for increasing the rate of this reaction.

Any four of the following:

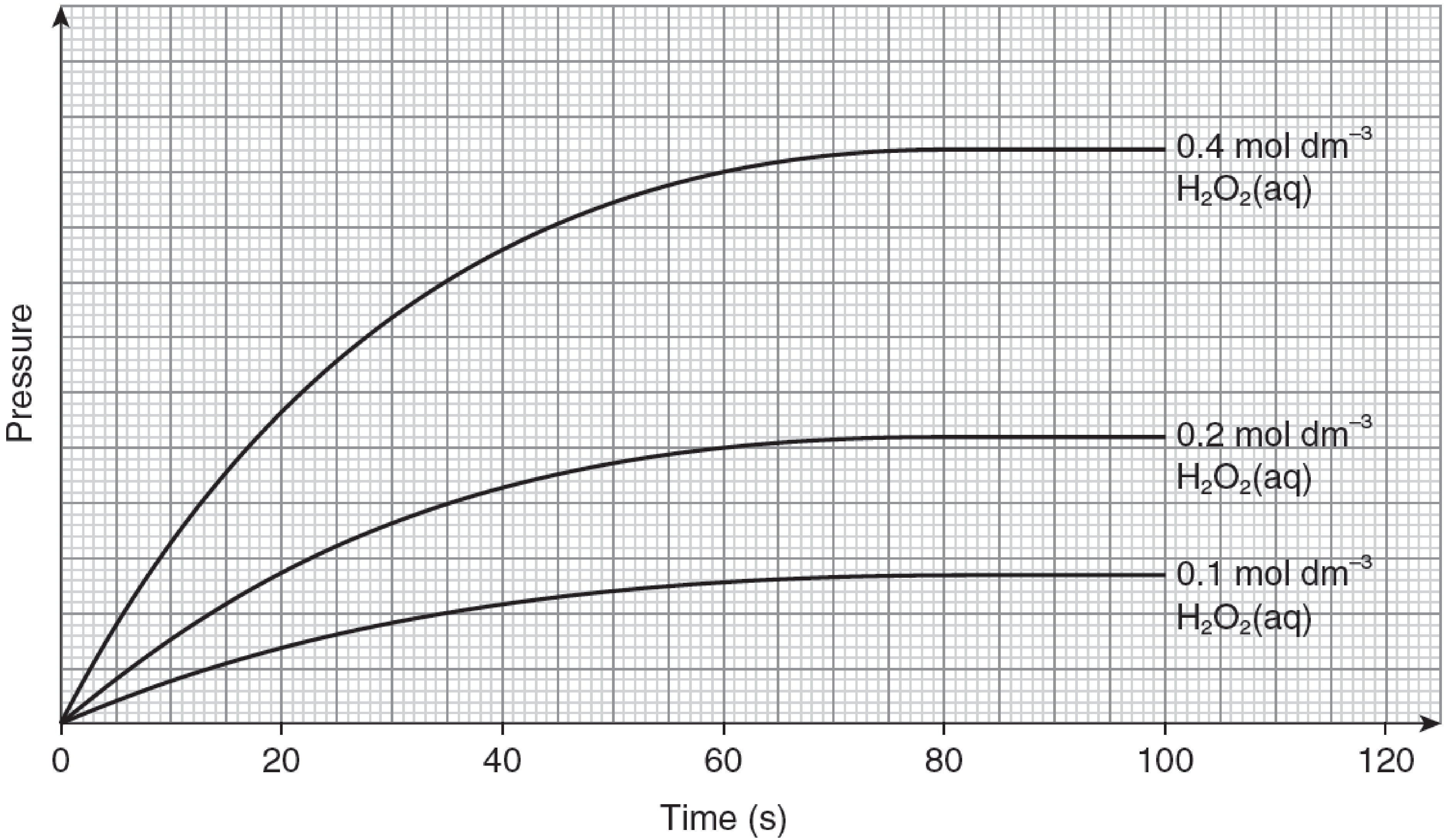
- Increase the temperature. (1)
- Increase the concentration of O2(g). (1)
- Increase the pressure. (1)
- Increase the surface area of carbon. (1)
- Add a catalyst. (1)

Unit Exercise (p.56)

15 Hydrogen peroxide solution decomposes in the presence of an enzyme called peroxidase. The products of this decomposition are water and oxygen.

A student followed the progress of this decomposition by using a data-logger with a pressure sensor. The graph below shows the relationship between the pressure and time measured with the use of hydrogen peroxide solutions of three different concentrations.

Unit Exercise (p.56)



Unit Exercise (p.56)

a) What is meant by the term ‘enzyme’?

A protein / catalyst that speeds up a reaction. (1)

b) Explain why a pressure sensor could be used in this experiment.

A gas evolves in the decomposition. (1)

c) How could you know that the decomposition had finished?

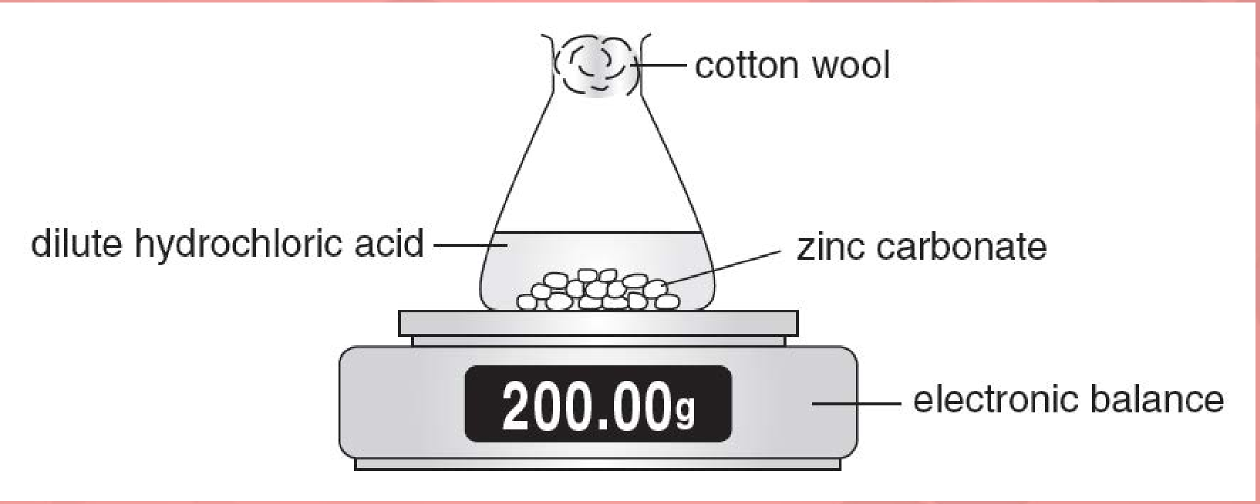
The pressure no longer changed. (1)

d) Describe how the concentration of hydrogen peroxide solution affects the rate of this decomposition.

Increasing the concentration increases the rate of decomposition. (1)

Unit Exercise (p.56)

16 A student investigated the reaction of zinc carbonate and dilute hydrochloric acid. The mass of the reaction flask and the reaction mixture was monitored during the course of the reaction using the experimental set-up shown below.



a) Explain why the mass of the reaction mixture decreased with time.
Carbon dioxide gas escaped. (1)

b) Why was a plug of cotton wool used in the experimental set-up?
To stop any acid from spraying out of the flask. (1)

Unit Exercise (p.56)

c) The student carried out the reaction at 25 °C using 10.0 g of small pieces of zinc carbonate and 50.0 cm^3 of 2.00 mol dm^{-3} HCl(aq).

i) Show, by calculation, that HCl(aq) was the limiting reactant.
(Relative atomic masses: C = 12.0, O = 16.0, Zn = 65.4)

$$ZnCO_3(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2O(l) + CO_2(g)$$

Number of moles

$ZnCO_3$: $\dfrac{10.0\ g}{125.4\ g\ mol^{-1}} = 0.0797$ mol

HCl: $2.00\ mol\ dm^{-3} \times \dfrac{50.0}{1\ 000}\ dm^3 = 0.100$ mol

In this reaction, 0.0500 mole of $ZnCO_3$ reacted with 0.100 mole of HCl. Thus, the acid was the limiting reactant. (1)

Unit Exercise (p.56)

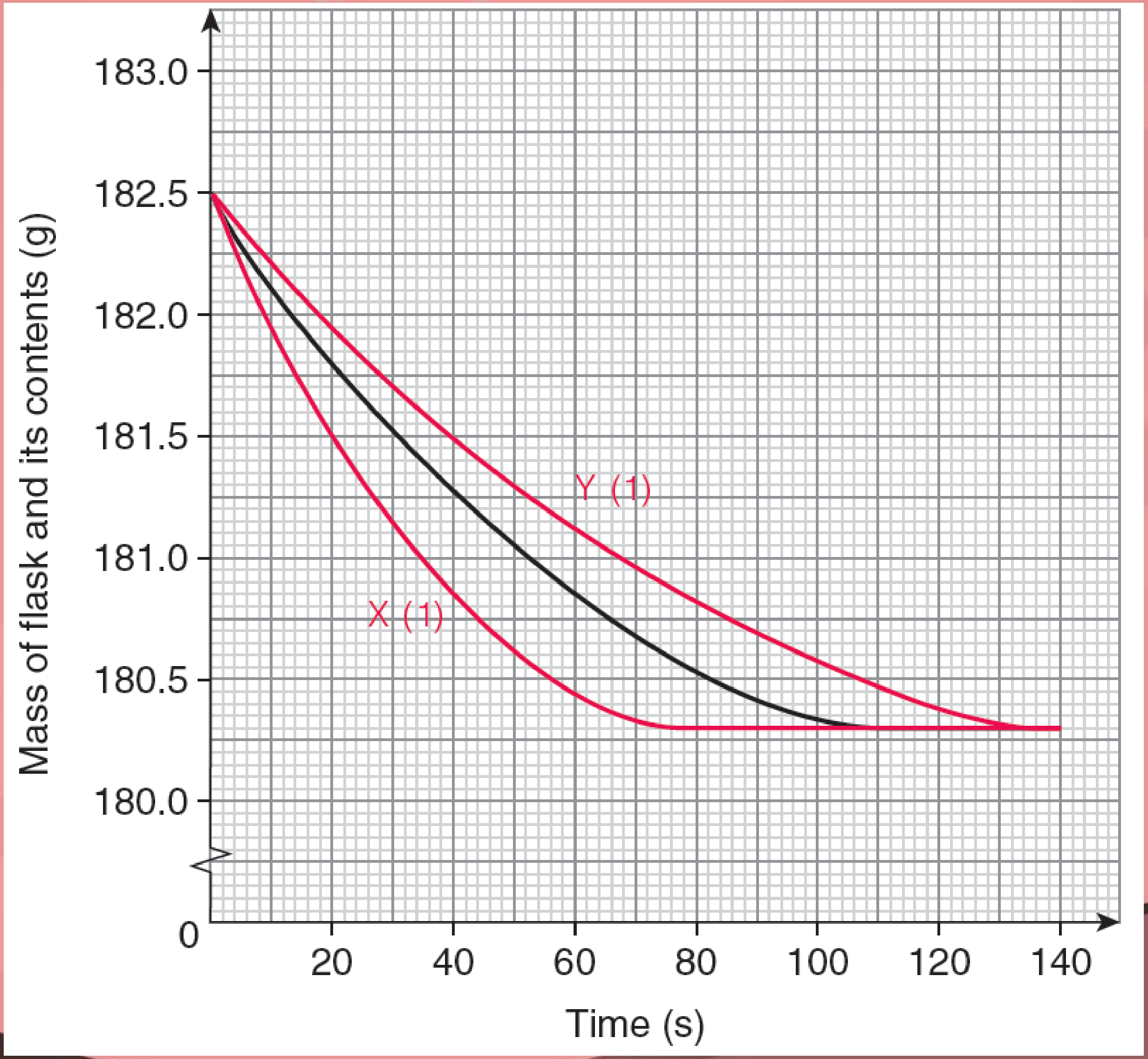
ii) The curve below shows the results obtained.

On the grid above, draw a curve to show how the mass of the flask and its contents would change with time under each of the following situations:

(1) When the experiment was carried out at 50 °C and all other conditions remained the same. (Label this curve as X.)

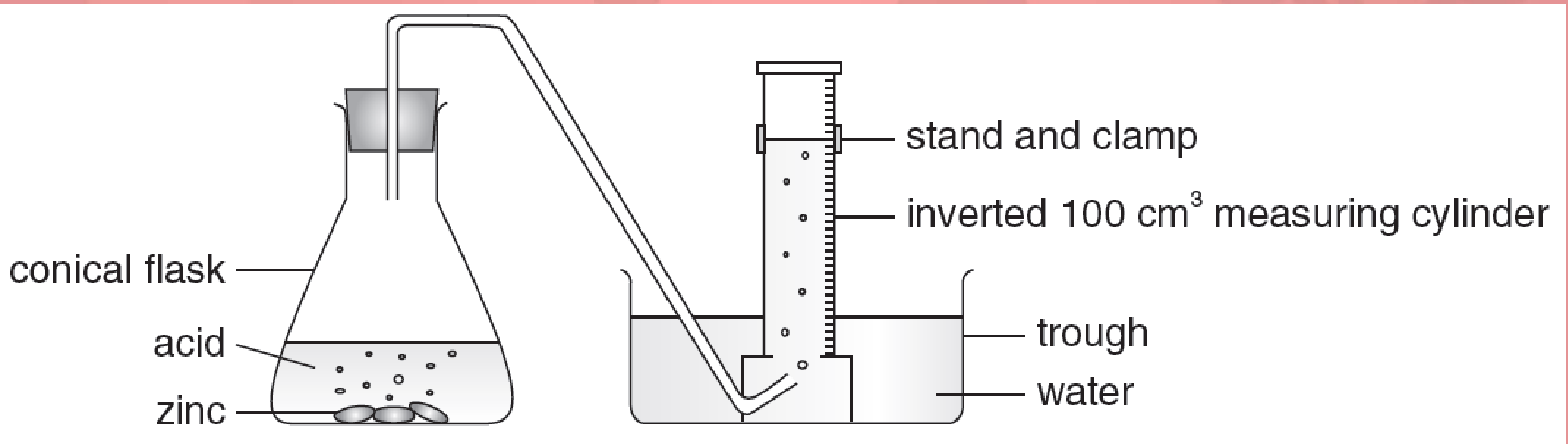
(2) When the experiment was carried out using 10.0 g of larger pieces of zinc carbonate and all other conditions remained the same. (Label this curve as Y.)

Unit Exercise (p.56)



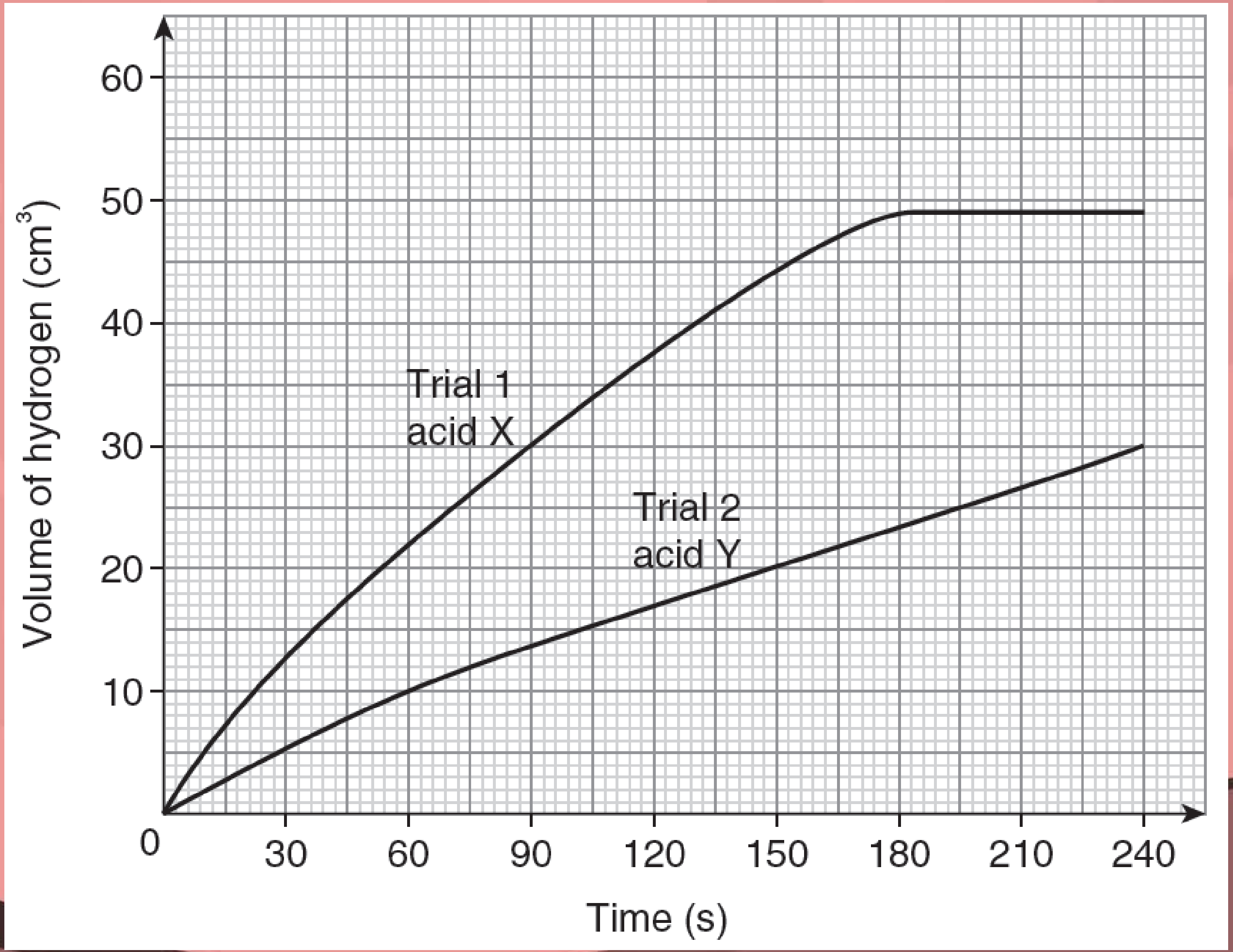
Unit Exercise (p.56)

17 A student investigated the rate of the reaction between zinc and two different monobasic acids X and Y, of the same concentration.
1.0 g of zinc granules (excess) was allowed to react with 50.0 cm^3 of each acid in the experimental set-up shown below.



Unit Exercise (p.56)

The curves below show the variations of the volume of hydrogen with time in the two trials.



Unit Exercise (p.56)

a) State which trial had the higher initial rate of reaction and suggest why the initial rate was higher in this trial.

Trial 1 had the higher initial rate of reaction. (1)

Acid X is stronger. (1)

b) The experiment was repeated using 50.0 cm^3 of acid X at a higher temperature but other conditions remained unchanged. Explain whether the total volume of gas obtained would be the same. (The volume of gas was measured at the same conditions.)

The same

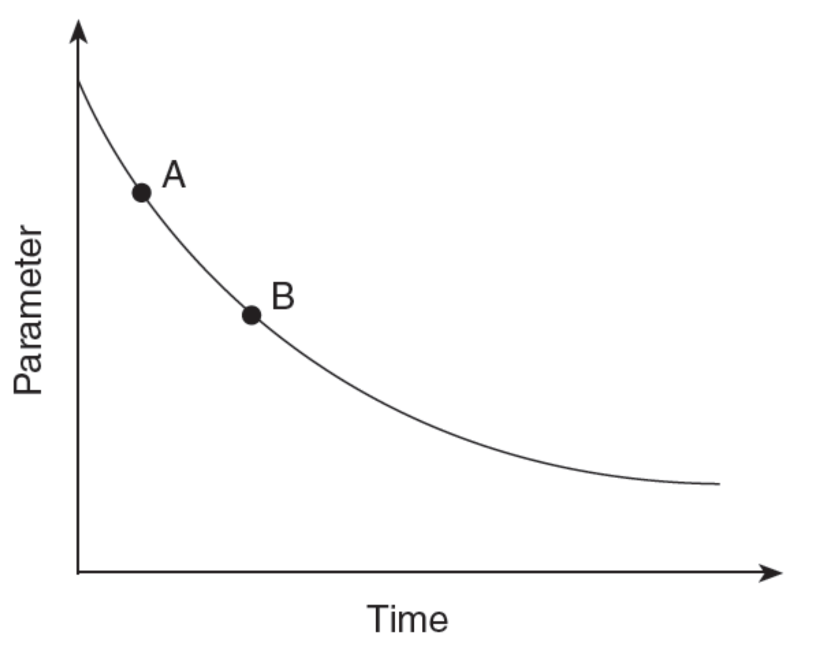
because the number of moles of acid (the limiting reactant) was the same for both cases. (1)

Unit Exercise (p.56)

18 Consider the following reaction:

$$Br_2(aq) + HCO_2H(aq) \rightarrow 2HBr(aq) + CO_2(g)$$

In an experiment to study the rate of consumption of $Br_2(aq)$, equal volume of 0.01 M $Br_2(aq)$ and 1.0 M $HCO_2H(aq)$ were mixed. The progress of the reaction was followed by measuring a certain parameter of the reaction system using a colorimeter. The graph below shows the results from the start of the reaction.



Unit Exercise (p.56)

a) Assume that the rate of change of the parameter with time can represent the rate of reaction.

i) According to the shape of the curve above, suggest what the parameter should be.

ii) The initial rate of the reaction can be determined by a suitable sketch on the above graph. Draw the suitable sketch on the above graph, and describe how the initial rate of the reaction can be obtained from the sketch.

iii) According to the graph above, the rate of reaction at A is higher than that at B. Explain this at molecular level.

b) Suggest another method that can follow the progress of the reaction.

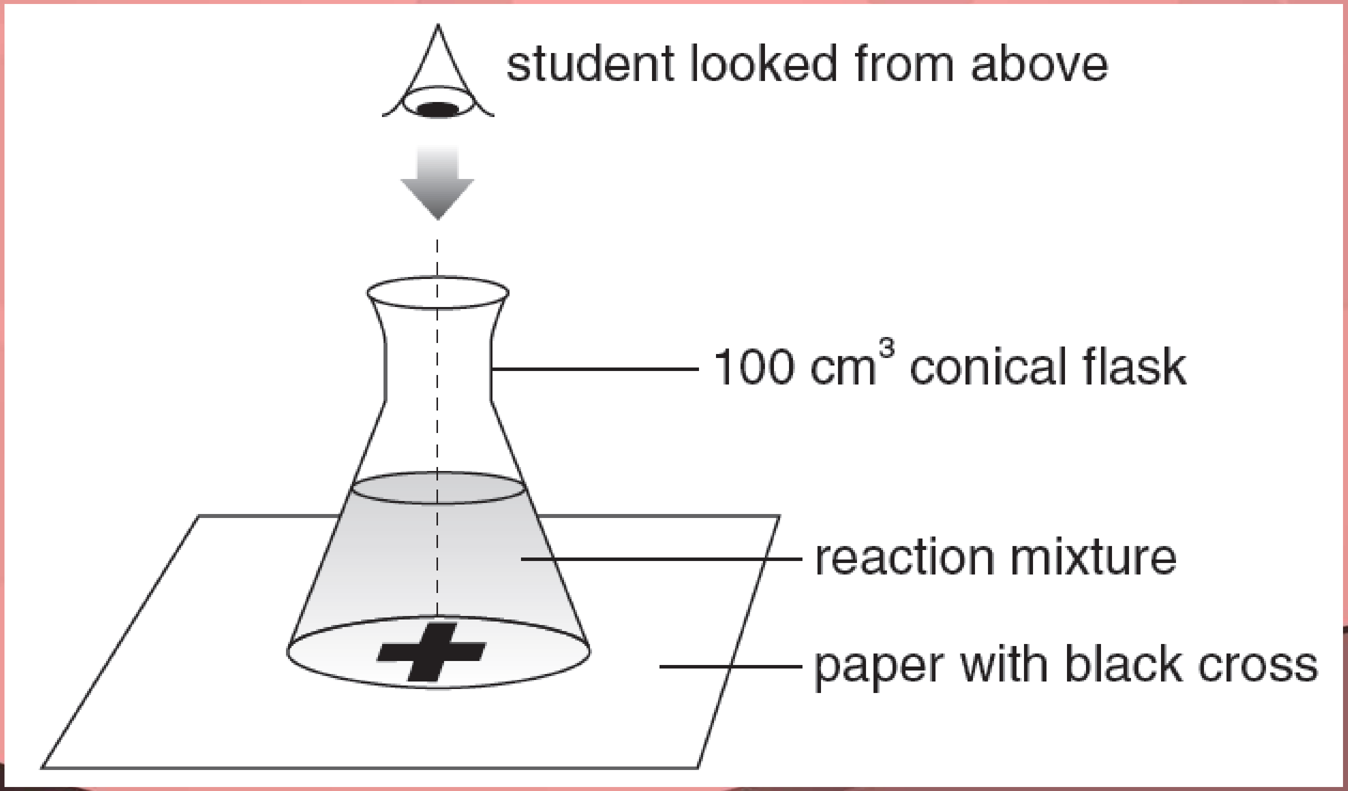
(HKDSE, Paper 1B, 2018, 11)

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.56)

19 The set-up shown below was used to investigate how the concentration of $S_2O_3^{2-}$(aq) ion affected the rate of the reaction below:

$$S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(aq) + H_2O(l)$$



Unit Exercise (p.56)

A student mixed 10.0 cm_3 of 1.0 mol dm^{-3} HCl(aq) and 20.0 cm^3 of $H_2O(l)$ in the flask. Then she added 5.0 cm^3 of 0.040 mol dm^{-3} $Na_2S_2O_3(aq)$ to the mixture. The time, *t*, required for the cross to become invisible when viewed from above was recorded.

The student repeated the procedure using the same volume of HCl(aq), but different volumes of $H_2O(l)$ and $Na_2S_2O_3(aq)$.

The table below lists the results obtained.

Trial	Volume used (cm^3)		*t* (s)
	$H_2O(l)$	0.040 mol dm^{-3} $Na_2S_2O_3(aq)$	
1	20.0	5.0	142
2	15.0	10.0	69
3	10.0	15.0	47
4	5.0	20.0	35
5	0.0	25.0	*x*

Unit Exercise (p.56)

a) Explain why different volumes of water were used in this investigation.

To keep the final volume of each reaction mixture the same. (1)

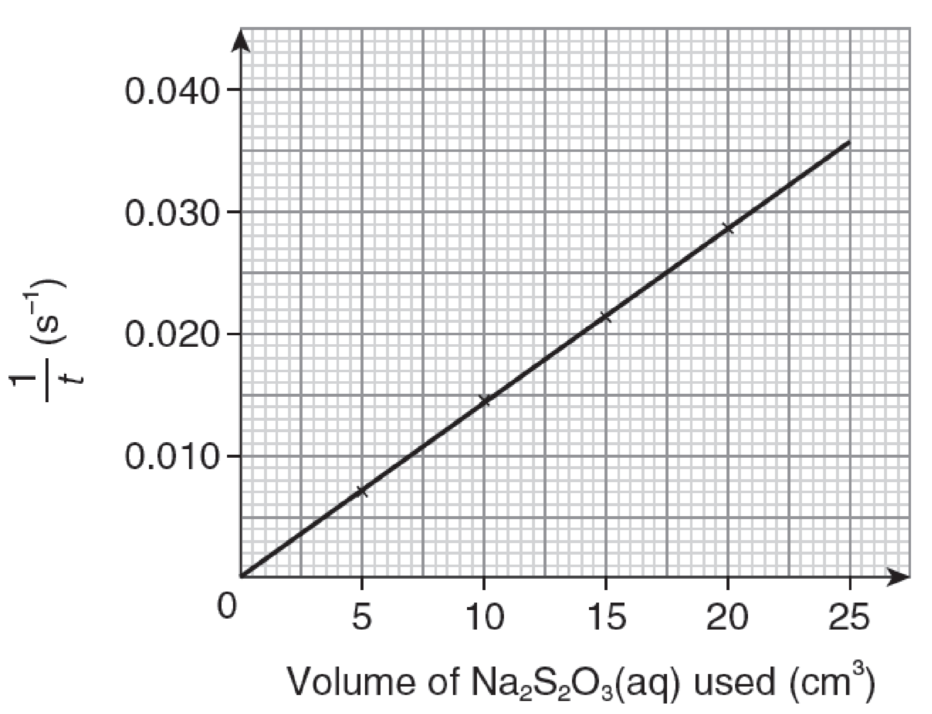
b) The student plotted a graph of 1/*t* against the volume of 0.040 mol dm^{-3} $Na_2S_2O_3(aq)$ used.

i) What conclusion can be drawn from this investigation?

The rate of reaction increases when the concentration of sodium thiosulphate is increased. (1)

ii) From the graph, estimate the value of x in the table.

$\frac{1}{x} = 0.036\ s^{-1}$, $x = 28$ s (1)



Unit Exercise (p.56)

c) Apart from the total volumes of the reaction mixtures and the concentration of the acid, name the most important factor which must be kept the same during each trial.

Temperature (1)

d) The student broke the conical flask carelessly after *Trial 4* and another student gave her a 250 cm^3 conical flask.

However, she said that she could not use the flask as it would affect the results.

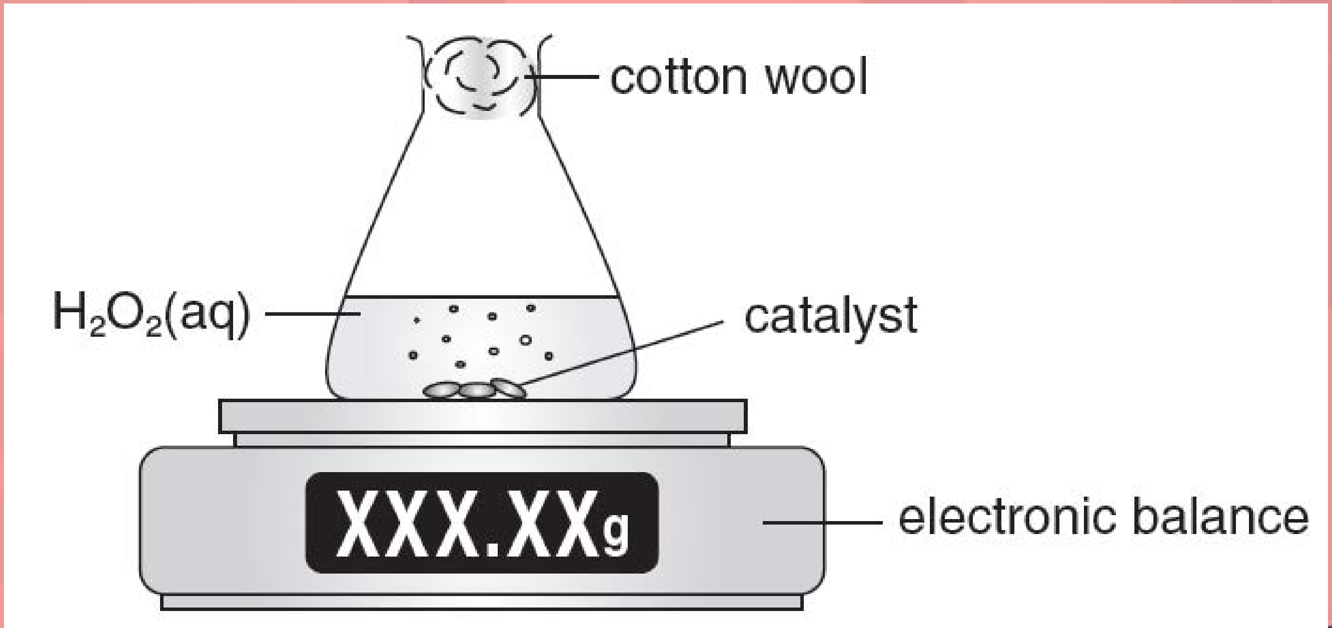
Suggest what effect, if any, this would have on the time taken for the cross to become invisible. Explain your answer.

The time would increase.

There would be a smaller depth of liquid and hence less sulphur precipitate through which the cross is observed. (1)

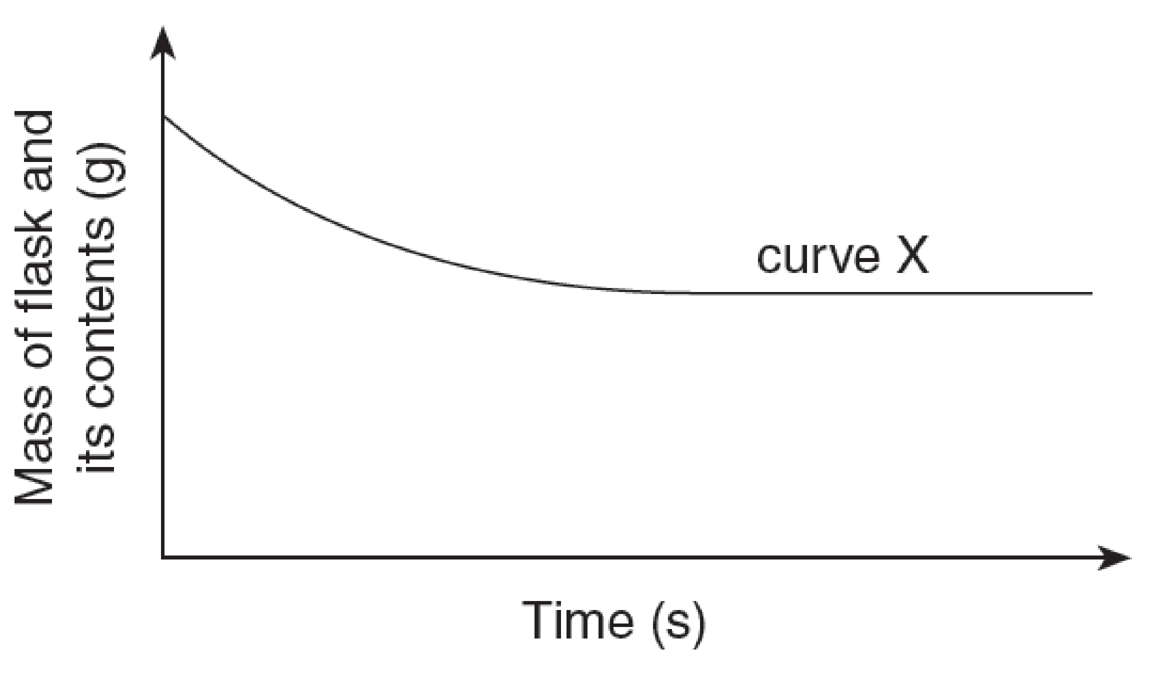
Unit Exercise (p.56)

20 At 298 K, 10.0 cm^3 of 1.00 mol dm^{-3} H_2O_2(aq) decomposed into O_2(g) and H_2O(l) in the presence of a catalyst. The mass of the flask and its contents was measured at regular time intervals using the experimental set-up shown below.



Unit Exercise (p.56)

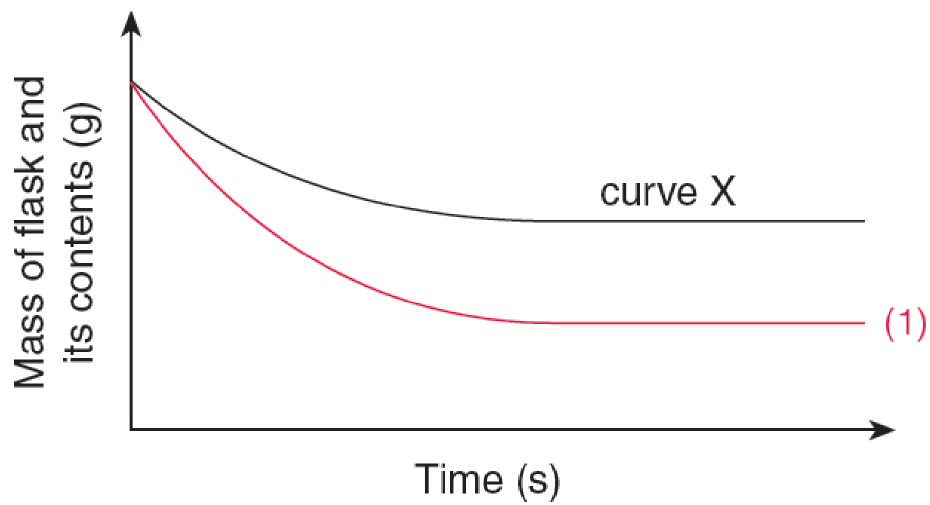
The results obtained are shown below.



Unit Exercise (p.56)

a) Draw a curve on Figure 1 to show how the mass of the flask and its contents would change with time if the experiment was repeated at 298 K using 10.0 cm^3 of 2.00 mol dm^{-3} $H_2O_2(aq)$ and all other conditions remained the same.

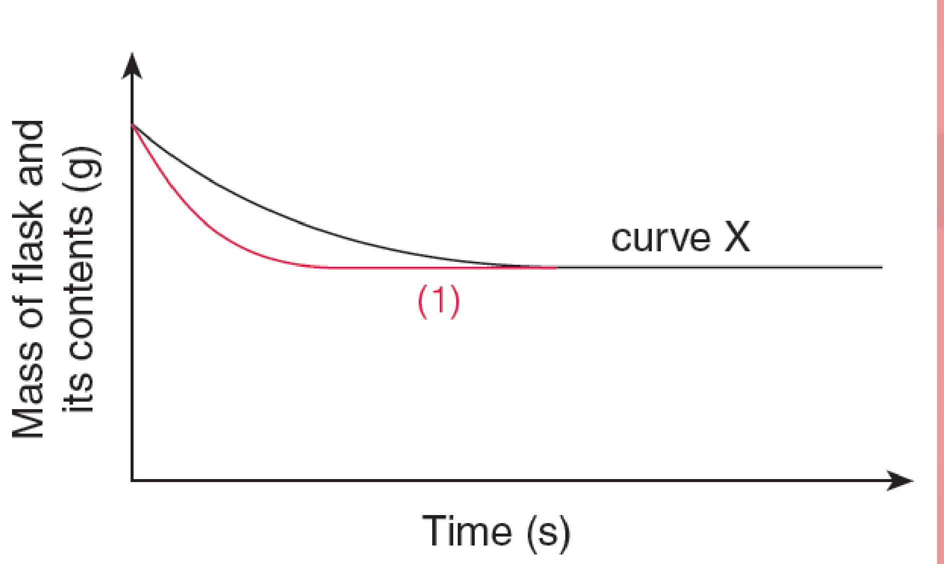
Figure 1



Unit Exercise (p.56)

b) Draw a curve on Figure 2 to show how the mass of the flask and its contents would change with time if the experiment was repeated at 308 K and all other conditions remained the same.

Figure 2



Unit Exercise (p.56)

21 Bromate ions oxidise bromide ions, in the presence of dilute acid, as shown in the equation below:

$$BrO_3^-(aq) + 5Br^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l)$$

Three trials were carried out using different initial concentrations of the reactants. The results are shown in the table below.

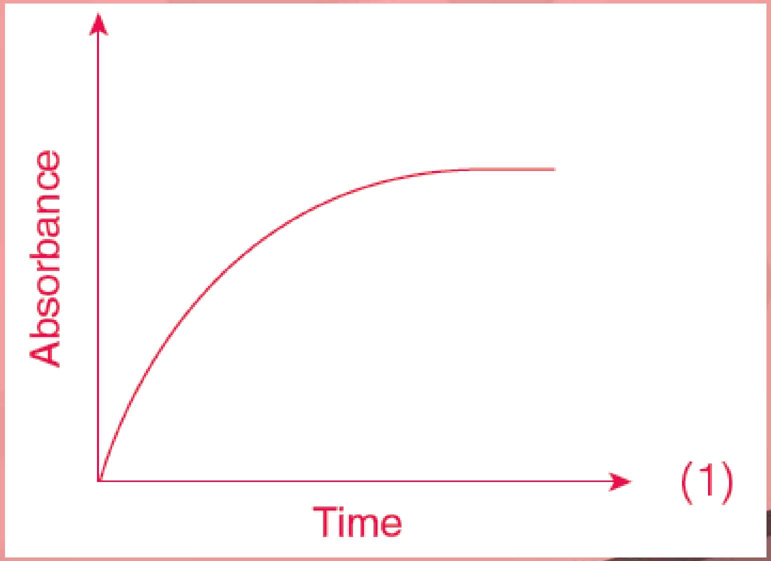
Trial	Initial concentrations of reactants (mol dm^{-3})			Initial rate of formation of Br_2(aq) (mol dm^{-3} s^{-1})
	BrO_3^-(aq)	Br^-(aq)	H^+(aq)	
1	0.050	0.25	0.30	1.68×10^{-5}
2	0.050	0.50	0.30	3.36×10^{-5}
3	0.15	0.50	0.30	1.01×10^{-4}

Unit Exercise (p.56)

a) i) Suggest a physical method, with justification, to follow the progress of the formation of Br_2(aq).

Follow the colour intensity of the reaction mixture / by colorimetry. (1)
The reaction mixture changes from colourless to yellow-brown. (1)

ii) Sketch a graph to show how the measured physical parameter would change with time.



Unit Exercise (p.56)

b) How is the initial rate of formation of $Br_2(aq)$ calculated from a graph of bromine concentration against time?

Determine the slope of the tangent to the curve at time = 0. (1)

c) What is the effect, if any, of increasing the initial concentration of

i) $BrO_3^-(aq)$; or Increase (1)

ii) $Br^-(aq)$ Increase (1)

on the initial rate of formation of $Br_2(aq)$?

d) Based on *Trial 1*, deduce the initial rate of consumption of $BrO_3^-(aq)$ ions under the experimental conditions.

Initial rate of consumption of $BrO_3^-(aq)$ ions = $\frac{1}{3}(1.68 \times 10^{-5})$ mol dm^{-3} s^{-1}

= 5.60×10^{-6} mol dm^{-3} s^{-1} (1)

Unit Exercise (p.56)

22 You are provided with common laboratory apparatus, calcium carbonate and 1 M hydrochloric acid. Outline how you would perform a fair comparison in studying the effect of different concentrations of acid on the rate of production of carbon dioxide from the following reaction:

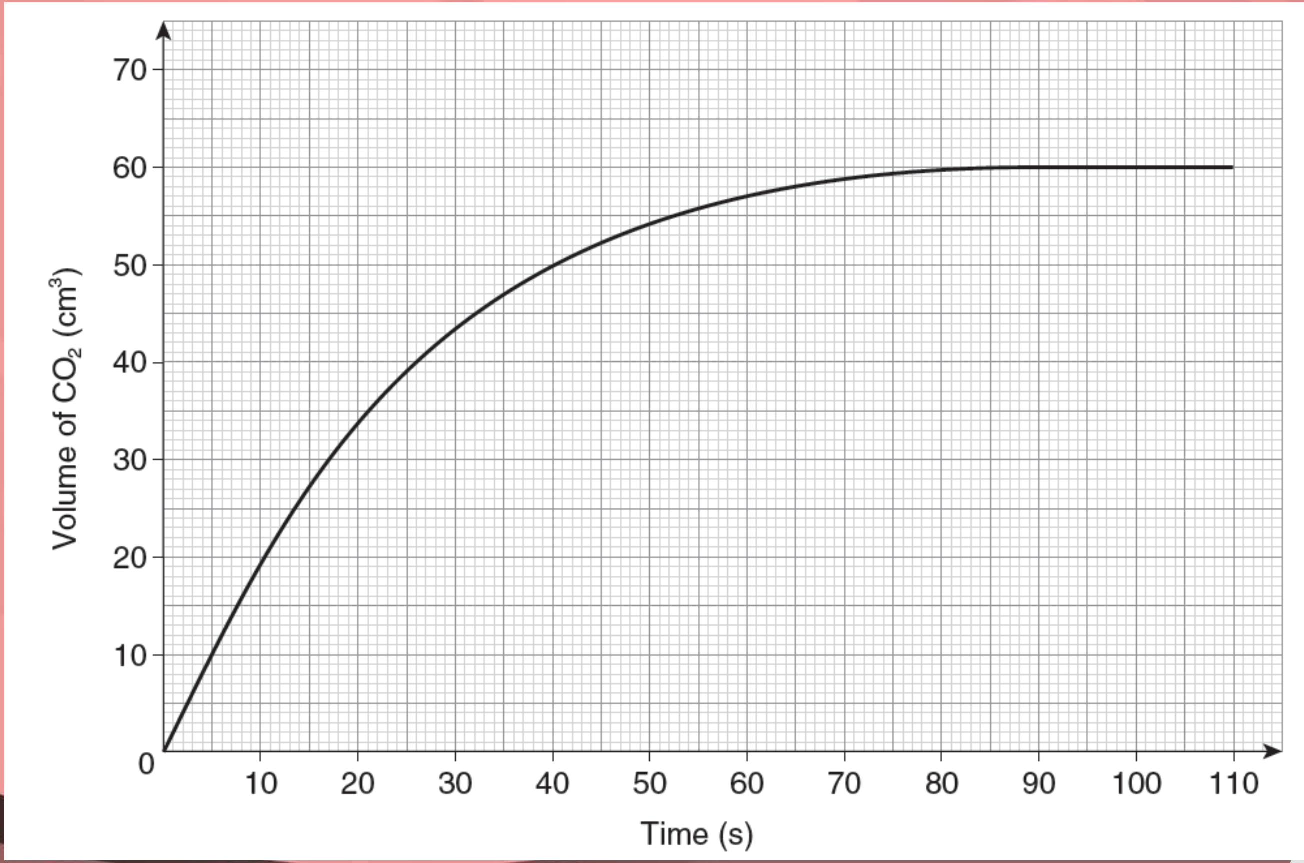
$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

(HKDSE, Paper 1B, 2014, 10)

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.56)

23 A student investigated the reaction between 0.25 g of small marble chips and 20.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid. The volume of carbon dioxide given off was measured every 10 seconds. The results are shown on the graph below.



Unit Exercise (p.56)

a) i) Describe how the rate of the reaction changes with time.

The rate of reaction is the highest at the start. (1)

The rate of reaction decreases with time. (1)

ii) Explain, in terms of collisions between particles, why the rate changes with time.

The concentration of the acid decreases gradually as the acid particles are consumed. (1)

The chance of collision decreases, so there would be less effective collisions in a unit volume per unit time. Hence the rate of reaction decreases.

b) The student repeated the experiment with 20.0 cm^3 of the same acid but using 0.25 g of large marble chips.

The rate of the reaction was found to be lower. Explain why in terms of the collisions between particles.

When large marble chips were used, the surface area decreased, with a smaller area over which collisions could occur. (1)

So there were fewer effective collisions in a unit volume per unit time.

Unit Exercise (p.56)

c) The student repeated the experiment with 50.0 cm^3 of the same acid and 0.25 g of small marble chips. Explain why the total volume of carbon dioxide was exactly the same as in the first two experiments.

The marble chips were the limiting reactant. (1)

Unit Exercise (p.56)

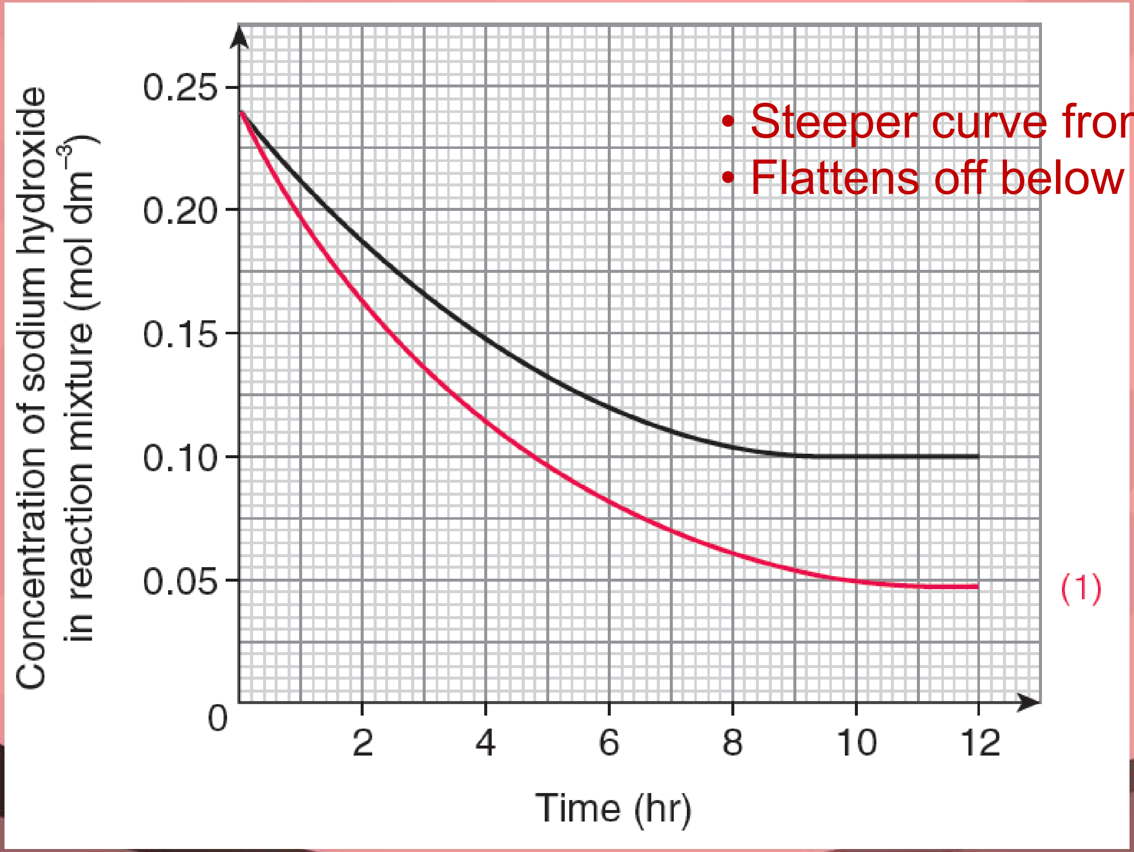
24 1-bromobutane reacts with excess sodium hydroxide solution to form butan-1-ol. A chemist followed the progress of this reaction by finding out how the concentration of sodium hydroxide in the reaction mixture changed with time.

a) Suggest how the progress of this reaction can be followed by titrimetric analysis.

Take samples of known volume at regular time intervals. (1)
Quench the reactions. (1)
Titrate each sample with a standard acid. (1)

Unit Exercise (p.56)

b) The results obtained are shown below. The experiment was repeated by increasing the concentration of 1-bromobutane in the reaction mixture and all other conditions remained the same. On the same grid, draw a curve to show how the concentration of sodium hydroxide in the reaction mixture would change with time.



- Steeper curve from the same starting point
- Flattens off below 0.10 mol dm^{-3}

Unit Exercise (p.56)

c) Increasing the concentration of 1-bromobutane increases the rate of this reaction. Explain in terms of collisions between particles.

Increasing the concentration of 1-bromobutane means increasing the number of 1-bromobutane particles per unit volume. The particles are more crowded and collide more often. (1)
So there are more effective collisions in a unit volume per unit time.

Unit Exercise (p.56)

25 Sulphur dioxide reacts with oxygen to form sulphur trioxide according to the equation below:

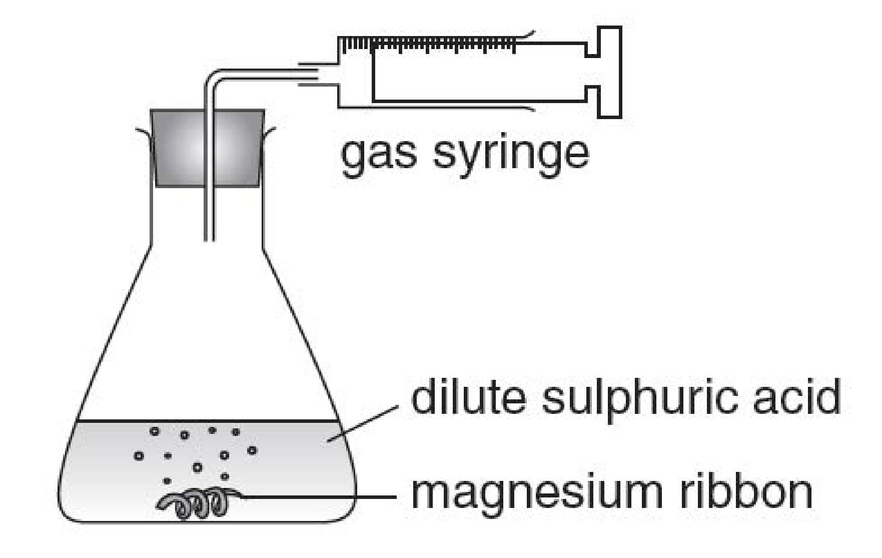
$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

How would increasing the pressure change the rate of the reaction? Explain your answer in terms of the behavior of particles.

Increasing the pressure increases the rate of the reaction. (1)
When the pressure is increased, there are more gas particles per unit volume, so they collide more often. (1)
There are more effective collisions in a unit volume per unit time.

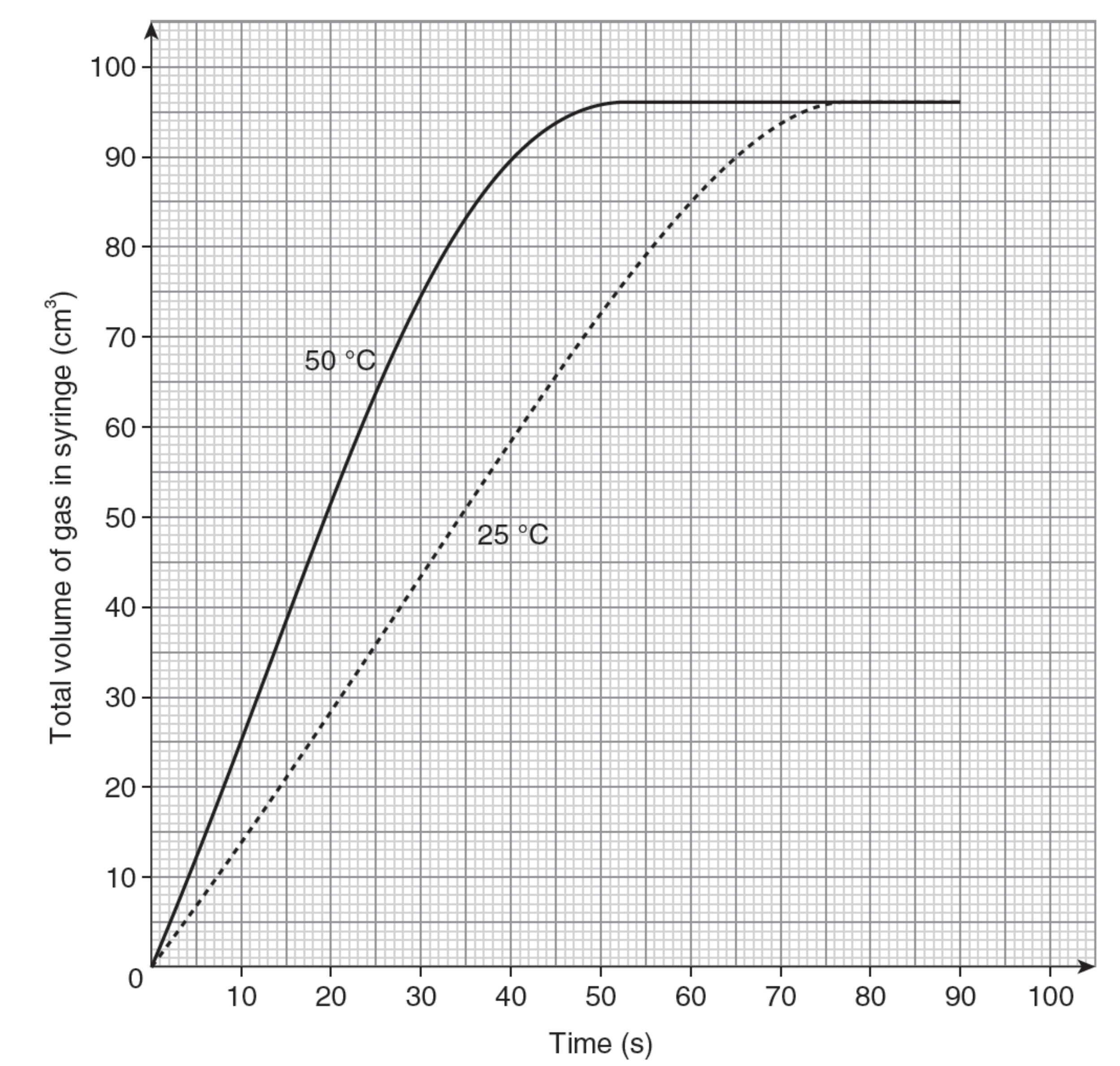
Unit Exercise (p.56)

26 A student investigated the rate of the reaction between magnesium ribbon and dilute sulphuric acid using the experimental set-up shown below.



The student recorded the volume of gas collected every 10 seconds and repeated the experiment using sulphuric acid at different temperatures. The student plotted results for the sulphuric acid at 25 °C and 50 °C on a graph.

Unit Exercise (p.56)



Unit Exercise (p.56)

State ONE conclusion the student could make about the effect of temperature on the rate of this reaction.
Explain this effect in terms of collisions between particles.

The rate at 50 °C is higher than the rate at 25 °C. / The higher the temperature, the greater the rate is. (1)
At a higher temperature, the reactant particles have more kinetic energy and move faster:

- the reactant particles involve in more collisions, resulting in more effective collisions; (1)
- a larger fraction of the colliding particles will have kinetic energy equal to or greater than the activation energy and hence a higher percentage of collisions can result in a reaction. (1)

Unit Exercise (p.24)

27 Harmful substances such as nitrogen monoxide and carbon monoxide are produced when petrol is burnt in car engines. Catalytic converters stop these harmful substances from being released into the air.

a) Write an equation to show how nitrogen monoxide and carbon monoxide react together in a catalytic converter.

$2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$ (1)

b) Platinum, palladium and rhodium are metals used inside catalytic converters. A very thin layer of the metals is used on a honeycomb ceramic support.

i) Why are catalysts used in chemical reactions?

To speed up the reactions. (1)

ii) Explain why a thin layer of catalysts is used.

It minimises the amount of catalysts used. (1)

It maximises the surface area on which reactions can take place. (1)

iii) Catalytic converters work for many years without replacing the catalysts. Explain why the catalyst replacement is NOT needed.

A catalyst remains chemically unchanged at the end of a reaction. (1)