Speed of Reaction

Learning Objectives

- Suggest suitable methods for investigating the effect of a given variable on the speed of a reaction.
- Interpret data obtained from experiments on speed of reaction.
- Describe the effect of concentration, pressure, particle size and temperature on the speeds of reactions.
- Explain speeds of reactions in terms of collisions between reacting particles.
- Use the concept of basicity of acids to explain its effects on the speeds of reactions.
- Define the term catalyst and describe the effect of catalysts on the speeds of reactions.
- Explain how pathways with lower activation energies account for the increase in speeds of reactions.
- State that some compounds act as catalysts in a range of industrial processes and that enzymes are biological catalysts.

A. INTRODUCTION

- Some chemical reactions are fast while some are slow.
    - Examples of fast reactions: explosive reactions, precipitation reactions, etc.
    - Examples of slow reactions: decay of food, rusting of iron, etc.
- The speed of a chemical reaction means how quickly a reactant is used up or how quickly a product is formed.

$$\text{speed of reaction} = \frac{\text{amount of reactant used up}}{\text{time taken}} = \frac{\text{amount of product formed}}{\text{time taken}}$$

B. MEASURING SPEED OF REACTION

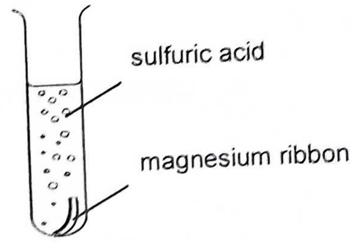
- The methods of measuring the speed of a reaction include:
    1) Measuring the time taken for a reaction to be completed

    2) Measuring the amount of product formed at regular time intervals
       Example: measure the volume of gas produced
       measure the amount of precipitate formed

    3) Measuring the amount of reactants used up at regular time intervals
       Example: measure the mass of reaction mixture that remains

- The method chosen for an investigation depends on the particular reaction.

B1. Measuring Time Taken for Completion of Reaction



- Consider the reaction between magnesium and dilute sulfuric acid.

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

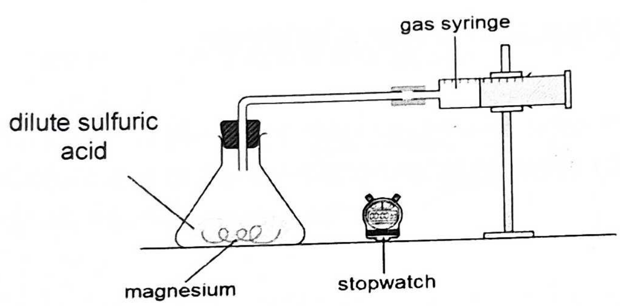
- We can measure the speed of the reaction by measuring the time for the magnesium ribbon to be completely 'dissolved' in the acid.
- The average speed of this reaction is inversely proportional to the time taken for the reaction to complete. **[speed $\propto$ 1/ time]**

B2. Measuring Volume of Gas evolved at Regular Time Intervals

- Consider the same reaction between magnesium and dilute sulfuric acid.

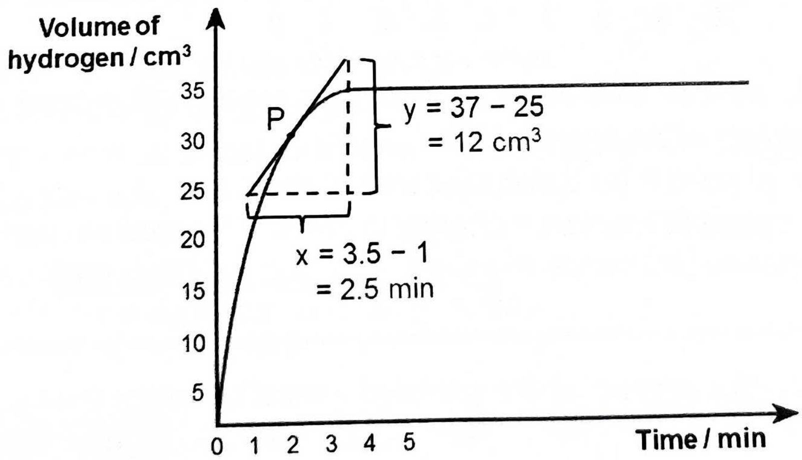
$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

- We can measure the speed of the reaction by collecting the hydrogen evolved in a gas syringe and measuring the volume of hydrogen collected at regular time intervals.



Time / min	Volume of hydrogen gas / cm^3
0	0.0
0.5	12.5
1.0	17.5
1.5	27.0
2.0	32.5
2.5	33.5
3.0	34.5
3.5	34.5
4.0	34.5
4.5	34.5

- As the reaction proceeds, the total volume of hydrogen collected increases with time as hydrogen is collected in a gas syringe.
- A typical graph of volume of gas collected against time is as shown.



- From the graph, we can determine the speed of reaction at various times by finding the gradient of the graph at the specific time.
  - Example: At point **P** (at 2 min after start of reaction),

speed of reaction = change in volume of gas / change in time

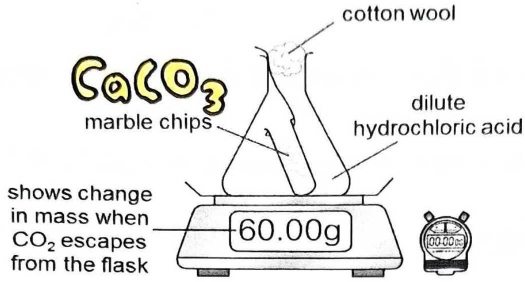
$= \underline{12 \div 2.5 = 4.80\ cm^3/min}$

B3. Measuring Mass of Reaction Mixture at Regular Time Intervals

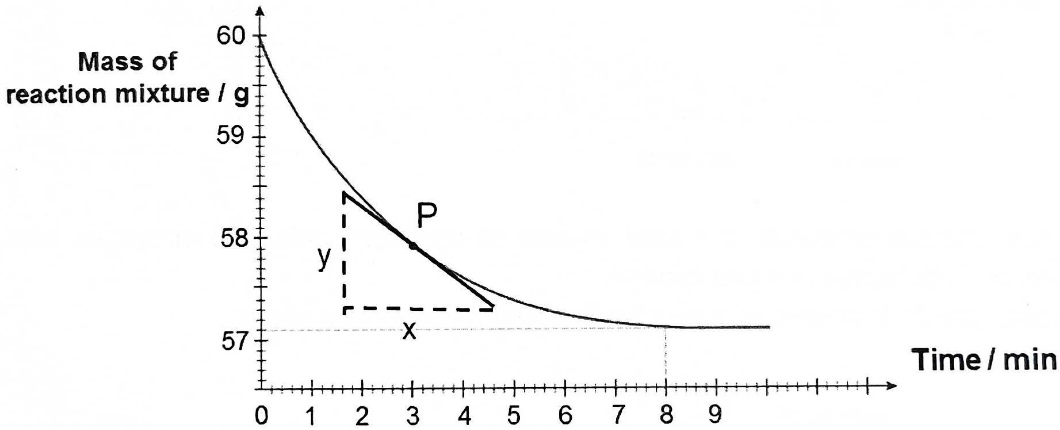
- Consider the reaction between marble (calcium carbonate) and dilute hydrochloric acid

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

- We can measure the speed of the reaction by standing the flask on an electronic balance and measuring the mass of reaction mixture at regular time intervals.



- As the reaction proceeds, the mass of the reaction mixture **decreases** with time as carbon dioxide escapes from the flask.
- A typical graph of mass of reaction mixture against time is as shown.



- From the graph, we can determine the speed of reaction at various times by finding the gradient of the graph at the specific time.
  - Example: At point P (at 3 min after start of reaction),

    speed of reaction = change in mass of reaction mixture / change in time

    = y ÷ x = 1.2 ÷ 3.0 = 0.400 g/min

Question:

Is this method suitable for measuring the speed of a reaction which produces hydrogen gas? Explain your answer.

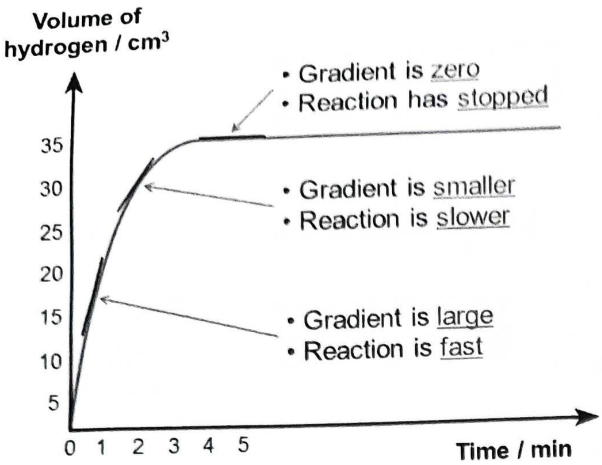
Answer:

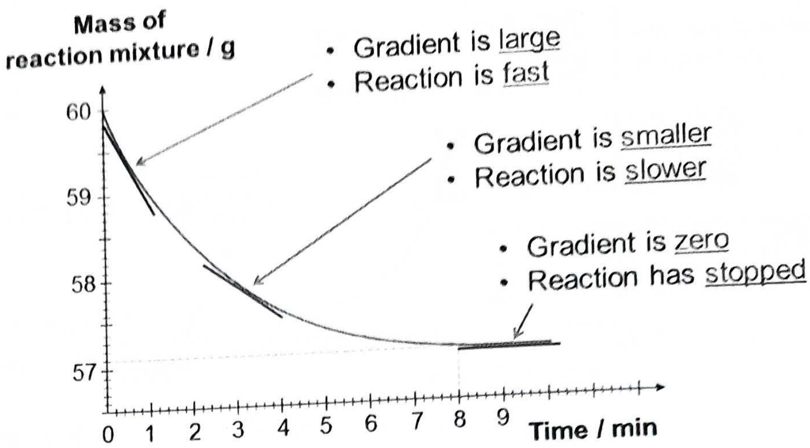
No, this method is not suitable.

This is because the **molar mass of hydrogen is very low**, thus the changes in mass of reaction mixture (due to loss of hydrogen gas) would be insignificant.

C. GRAPHS FOR SPEED OF REACTION

- Study the graphs below.



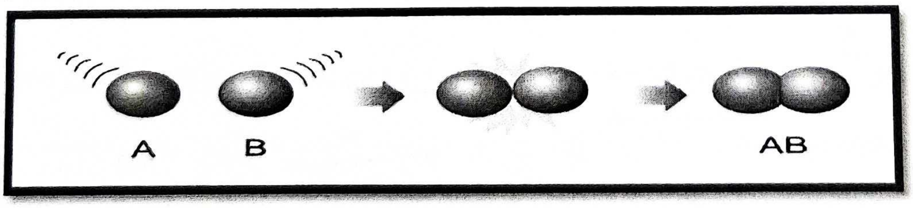


- The graphs show several important points:
  - Gradient of the graph is the **steepest** at the start of the reaction, thus speed of reaction is **greatest**.
  - Gradient of the graph **decreases** with time, thus speed of reaction **decreases with time**.
  - Gradient of the graph becomes **zero** when reaction stops, thus speed of reaction is **zero** when reaction is completed.

D. THE COLLISION THEORY

- According to the collision theory, a chemical reaction can occur only when the **reacting particles (atoms, molecules, ions)** satisfy all of the following:
  - reacting particles must **collide with one another**.
  - reacting particles must be **orientated properly**.
  - reacting particles must **possess a minimum energy**, known as **activation energy**, before reaction can take place.

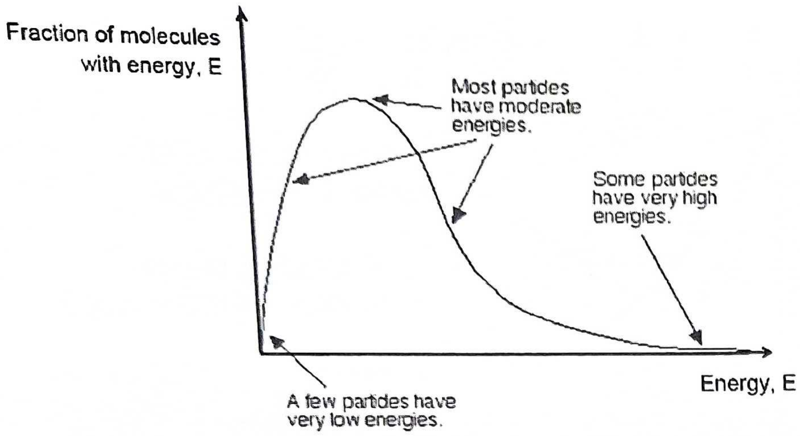
Recall: that the activation energy (E_a) is the minimum amount of energy that reacting particles must possess in order for a chemical reaction to occur.

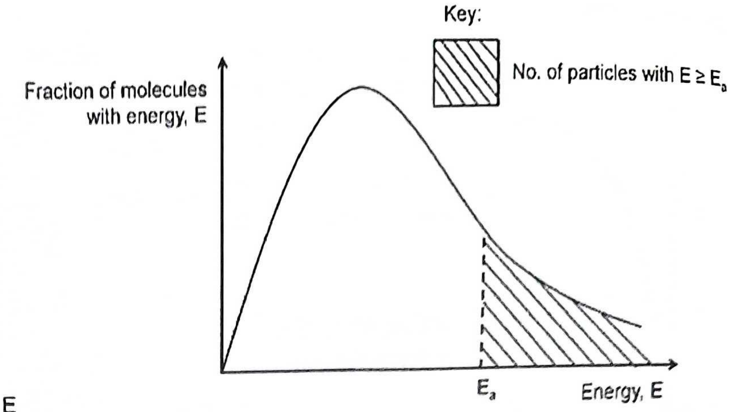


- The collisions that lead to a chemical reaction and formation of products are known as **effective collisions.**
- The rate of reaction is **proportional** to the **frequency of effective collisions.**

D1. Maxwell Boltzmann Distribution Curve

- In any system, the reacting particles present have a very wide range of <u>energies</u> as shown in the **Maxwell Boltzmann Distribution Curve**.

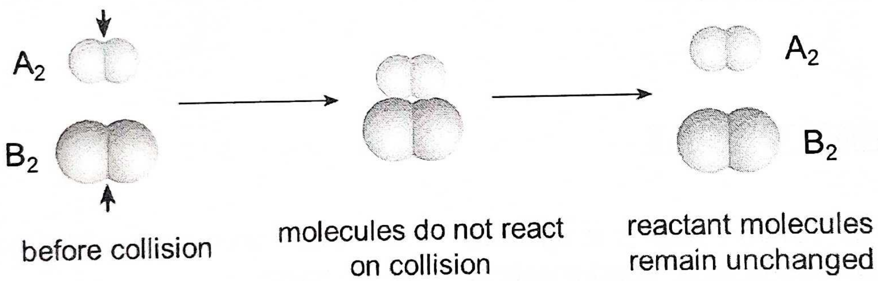




- Only a fraction of reacting particles has energy <u>equal to</u> or <u>more than</u> the activation energy.

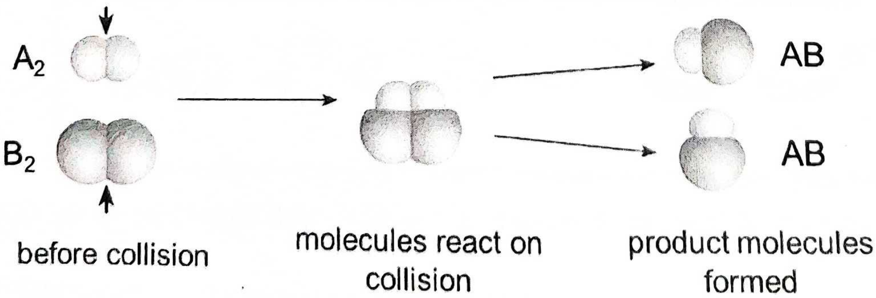
(a) Collision of insufficient energy (Energy of collision < Activation energy)

- Most reacting particles do not have <u>**sufficient**</u> energy to react when they collide.
- The reacting particles just bounce apart and do not react.



(b) Collision of sufficient energy (Energy of collision > Activation energy)

- Some reacting particles possess <u>**sufficient**</u> energy when they collide to provide the activation energy needed.
- Thus <u>**bond breaking**</u> in the reactants occur, a reaction starts.



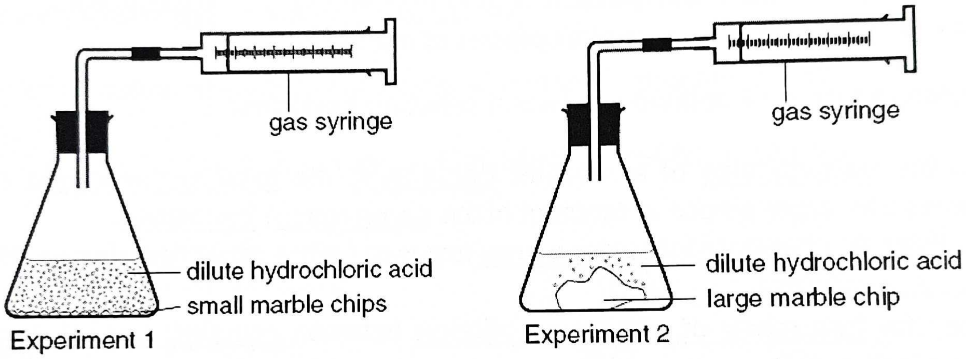
FACTORS AFFECTING SPEED OF REACTION

- There are several factors that can affect the speed of a chemical reaction. These include:
  1) Effect of **particle size** (of solid reactant)
  2) Effect of **concentration** (of solution)
  3) Effect of **pressure** (of gaseous reactant)
  4) Effect of **temperature** (of reaction mixture)
  5) Effect of **catalyst** (being present or not)

E1. Effect of Particle Size (of solid reactant)

- Small particle size of solid: powder
- Large particle size of solid: lumps, granules, ribbon
- Consider the reaction between marble (calcium carbonate) and dilute hydrochloric acid

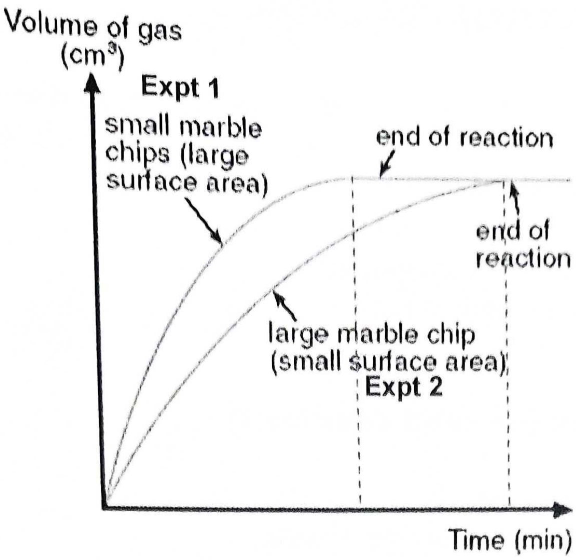
$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$



	Experiment 1	Experiment 2
Mass of $CaCO_3$	0.500 g	0.500 g
Particle size of $CaCO_3$	**Small chips**	**Large chips**
Concentration of acid	1.00 mol/dm^3	1.00 mol/dm^3
Volume of acid	100 cm^3	100 cm^3
Temperature	Room temperature	Room temperature

Results:

- We would observe that effervescence of colourless an odourless gas occurs <u>**more quickly**</u> in the first reaction. This indicates that the reaction is <u>**faster**</u> when <u>**smaller pieces**</u> of marble are used.
- The two graphs of 'volume of carbon dioxide' against 'time' are as shown.

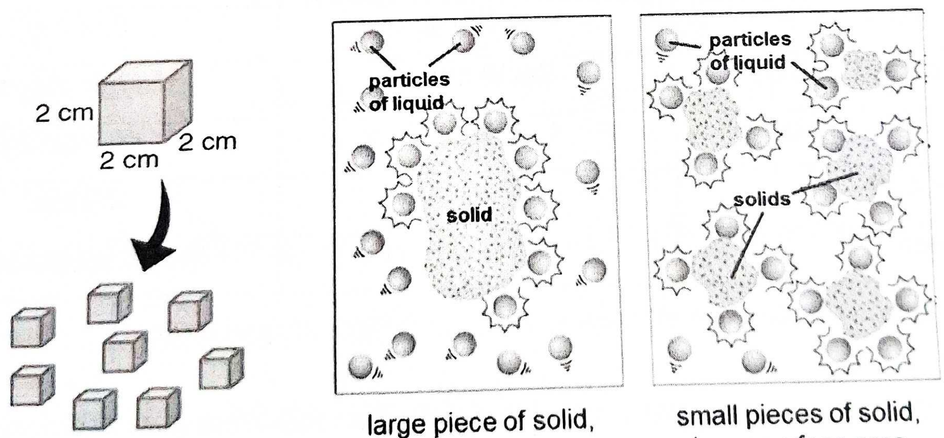


Interpretation of Graph:

1) In Experiment 1, the time taken for the reaction to complete is shorter than Experiment 2.
2) For Experiment 1, the initial gradient of graph is steeper than Experiment 2. Thus, the rate of the reaction is faster when **smaller pieces** of marble are used.

Explanation in terms of collision between reacting particles:

- When the particle size of a reactant decreases, the total surface area of the reactant (compared to larger pieces of reactant of the **same mass**) increases.
- Thus, there is a **larger total surface area** exposed for the **other reacting particles** to collide with at any one time.
- Hence, the frequency of effective collision between reacting particles increases, and speed of reaction increases.



large piece of solid, small surface area

small pieces of solid, large surface area

Applications:

- In modern coal-burning power stations, powdered coal is used for burning as powered coal has greater surface area, which makes it burns faster than bigger pieces of coal.
- In coal mines, if air contains too much coal dust, explosions can occur from the ignition of coal dust by a single spark from machinery

Note on Reacting Particles:

- Reacting particles refer to the particles that take part in a chemical reaction.
  - Example: Reaction between calcium carbonate and hydrochloric acid
  - Chemical equation:
    - $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
  - Ionic equation: $CaCO_3(s) + 2H^+ \cancel{aq} (aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$
  - The reacting particles refer to $CaCO_3$ and H^+
  - **Chloride, Cl^- ions are spectator ions** that do **not** take part in the reaction.

- **Question**:

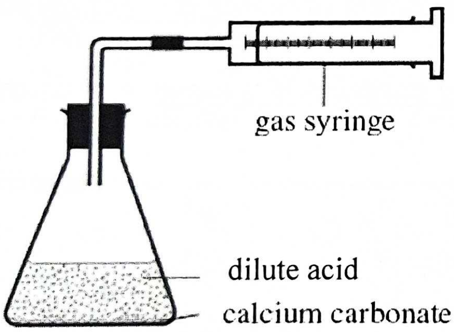
  Identify the reacting particles in the reaction between magnesium ribbon and dilute sulfuric acid.

  Mg and H^+

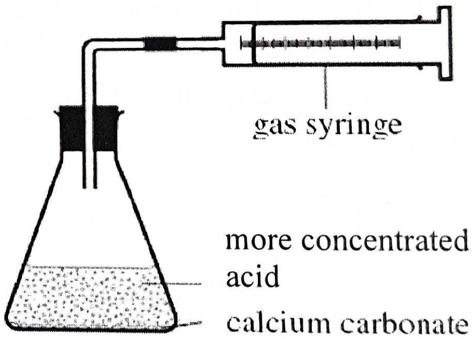
E2. Effect of Concentration (of solution)

- Consider the reaction of calcium carbonate with dilute hydrochloric acid

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



Experiment 1

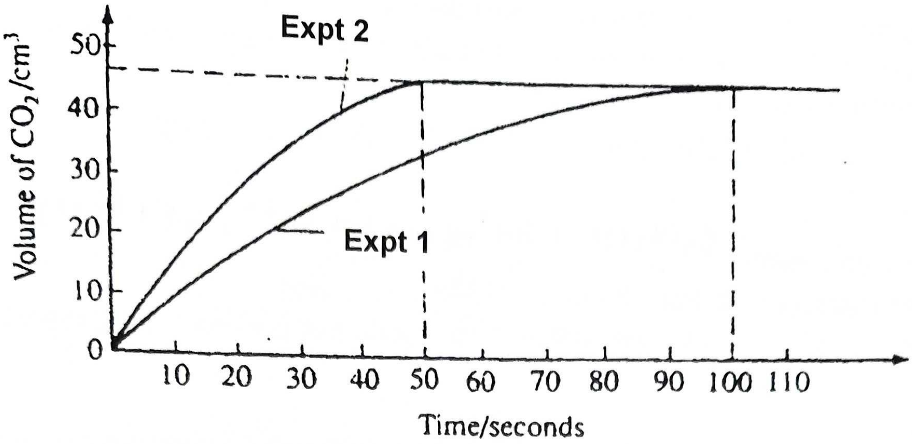


Experiment 2

	Experiment 1	Experiment 2
Mass of $CaCO_3$	0.200 g	0.200 g
Particle size of $CaCO_3$	powder	powder
Concentration of acid	**1.00 mol/dm³**	**2.00 mol/dm³**
Volume of acid	50.0 cm³	50.0 cm³
Temperature	room temperature	room temperature

Results: more concentration, end faster.

carbonate is limiting

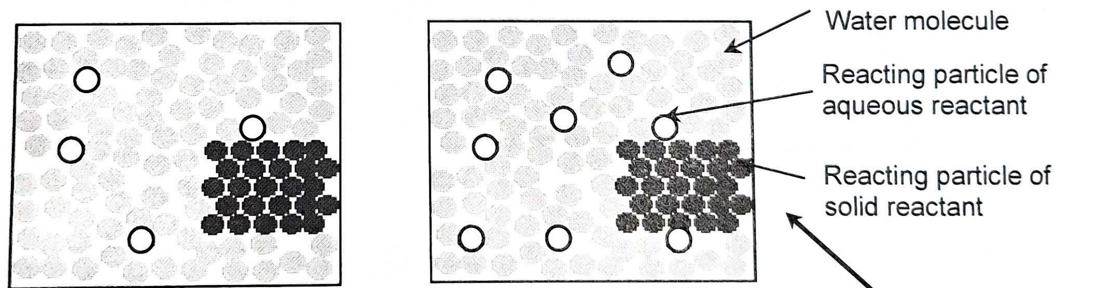


Interpretation of Graph:

1) In Experiment 1, the reaction lasts <u>**100 s**</u>, whereas in Experiment 2, it lasts <u>**50 s**</u>. This means that the rate of reaction in Experiment 1 is <u>**slower**</u> than that in Experiment 2.
2) For Experiment 2, the initial gradient of graph is <u>**steeper**</u> than Experiment 1. This means that when concentration of acid increases, rate of reaction <u>**increases**</u>.

Explanation in terms of collisions between reacting particles:

- When concentration of a reactant is increased, the <u>**number**</u> of <u>**reacting particles**</u> per unit <u>**volume increases**</u>.
- <u>**Frequency**</u> of <u>**effective collision**</u> between <u>**reacting particles increases**</u>. Hence, the rate of reaction <u>**increases**</u>.

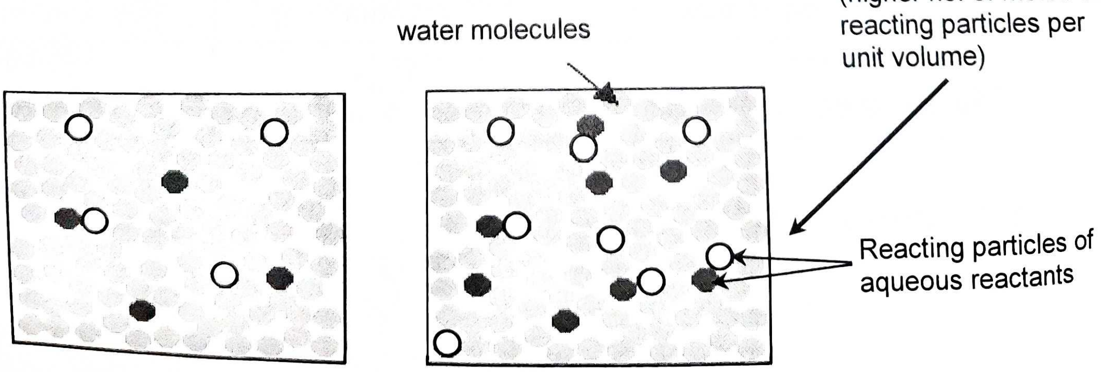


<u>**One reactant is a solid**</u>

E.g. $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$

Low concentration (lower no. of moles of reacting particles per unit volume)

High concentration (higher no. of moles of reacting particles per unit volume)



<u>**Both reactants in solution**</u>

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

Question:
From the graph, the volume of carbon dioxide produced in both Experiments 1 and 2 is the same. Explain why.

Answer:

Experiment 1

$CaCO_3(s)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$H_2O(l)$	+	$CO_2(g)$
0.200g M_r = 100.1		1.00mol/dm^3 50.0cm^3 = 0.005dm^3						24 dm^3 at r.t.p.
0.0019980 ~~0.0020~~ mol (limiting)		0.00500 mol (excess)						0.0019980 mol

Experiment 2

$CaCO_3(s)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$H_2O(l)$	+	$CO_2(g)$
0.200g		2.00mol/dm^3 50.0cm^3						dependent on limiting 24 dm^3 at r.t.p.
0.0019980 mol (limiting)		0.100 mol (excess)						0.0019980 mol

- For both Experiment 1 and 2, HCl is in excess and the limiting reactant is $CaCO_3$.
- This means that for the same number of moles of $CaCO_3$ used in both experiments, the same number of moles of CO_2 will be produced.
- Hence, volume of carbon dioxide produced will be the same.

Question:
If the acid is the limiting reactant instead of calcium carbonate, as shown in the experiments 3 and 4 below.

	Experiment 3	Experiment 4
Mass of $CaCO_3$	excess	excess
Particle size of $CaCO_3$	powder	powder
Concentration of acid	**1.00 mol/dm^3**	**2.00 mol/dm^3**
Volume of acid	50.0 cm^3	50.0 cm^3
Temperature	room temperature	room temperature

(i) Explain how the **volume** of carbon dioxide gas produced would differ.
(ii) Explain how the **rate** of reaction would differ.
(iii) Sketch the graphs of rate of reaction for Experiments 3 and 4 on the same axes.

Answer:

(i) **Experiment 3**

$CaCO_3(s)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$H_2O(l)$	+	$CO_2(g)$
excess		1.00 mol/dm^3 50.0 cm^3						0.6 dm^3 at r.t.p
		0.0500 mol						0.0250 mol

Experiment 4

$CaCO_3(s)$	+	$2HCl(aq)$	→	$CaCl_2(aq)$	+	$H_2O(l)$	+	$CO_2(g)$
excess		2.00 mol/dm^3 50.0 cm^3		half no. of mols produced.				1.20 dm^3 at r.t.p
		0.100 mol						0.0500 mol

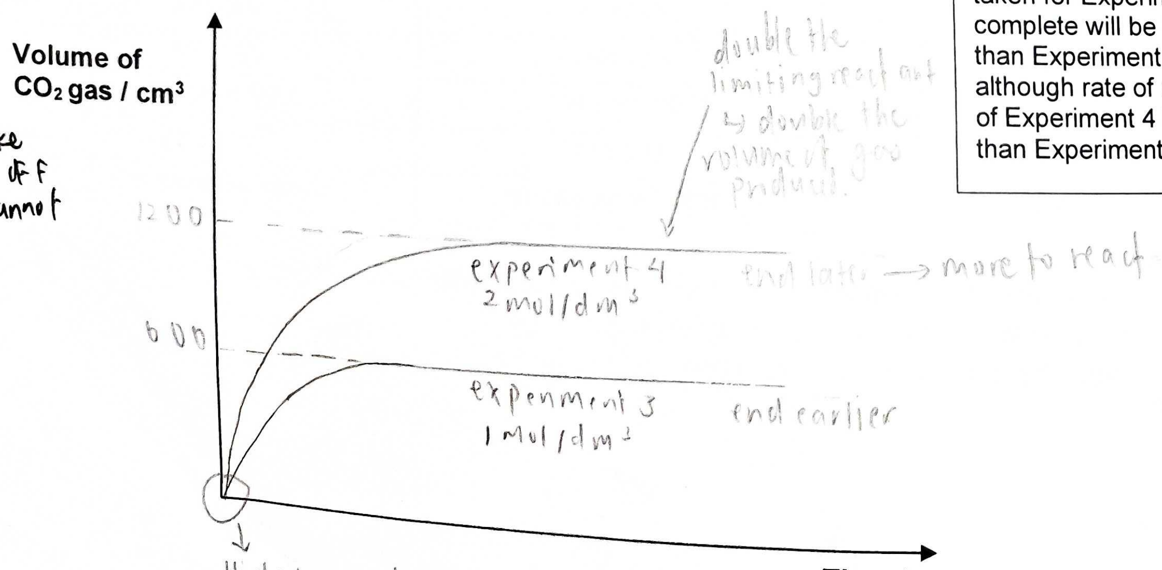
- In Experiments 3 and 4, $CaCO_3$ is in excess and hydrochloric acid is the limiting reactant.
- Since concentration of acid used in Experiment 4 is <u>twice</u> that in Experiment 3. For the same volume of acid, the number of moles of acid used in Experiment 4 will be <u>twice</u> that of Experiment 3.
- This means the number of moles of carbon dioxide produced in Experiment 4 will be also be **twice** that of Experiment 3.
- Hence, volume of carbon dioxide gas produced in Experiment 4 will be **twice** that of Experiment 3.

(ii) Rate of reaction for Experiment 4 is <u>faster</u> than Experiment 3 as the **concentration of acid** used is <u>higher</u>.

Thinking Question:
Suggest why the time taken for Experiment 4 to complete will be longer than Experiment 3 although rate of reaction of Experiment 4 is faster than Experiment 3?

(iii)

common mistake
→ do not taper off
→ gradient cannot coincide



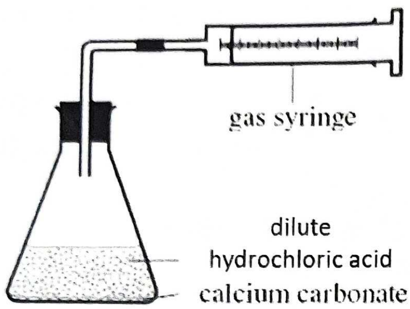
Higher concentration
→ Faster reaction
→ Steeper gradient

E2.1 Effect of Basicity of Acid

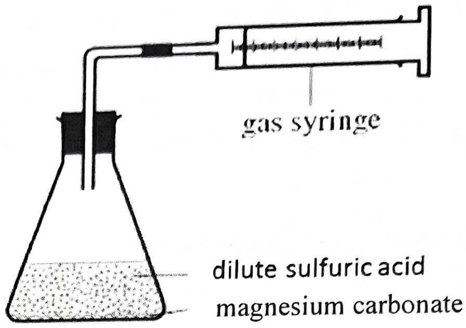
- Consider the reaction between magnesium carbonate with different types of acids.

Experiment 1: $MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$

Experiment 2: $MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$



Experiment 1



Experiment 2

	Experiment 1	Experiment 2
Mass of $MgCO_3$	1.00 g	1.00 g
Particle size of $MgCO_3$	powder	powder
Type of Acid (Basicity)	**Hydrochloric acid, $HCl(aq)$ (monobasic)**	**Sulfuric acid, $H_2SO_4(aq)$ (dibasic)**
Concentration of acid	Same for both experiments	
Volume of acid	Same for both experiments	
Temperature	Room temperature	

Case 1: Limiting reactant is $MgCO_3$

Explain how (i) volume of carbon dioxide gas produced and (ii) the rate of reaction would differ (if any) for Experiment 1 and 2 if the limiting reactant is $MgCO_3$.

(i)

Experiment 1

$\mathbf{MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)}$

$MgCO_3$: 1.00 g; 0.011862 mol (limiting)

HCl: excess

CO_2: 0.28469 dm³ at r.t.p ↑ × 24.0; 0.011862 mol

Experiment 2

$\mathbf{MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)}$

$MgCO_3$: 1.00 g; 0.011862 mol (limiting)

H_2SO_4: excess

CO_2: 0.28469 dm³ at r.t.p ↑ × 24.0; 0.011862 mol

- Volume of carbon dioxide gas produced in both experiments is the <u>same</u>.
- This is because for the <u>same</u> **number of moles** of $MgCO_3$ used, the number of moles of carbon dioxide produced will be the <u>same</u>. Hence volume of carbon dioxide produced for both experiments will be the <u>same</u>.

(ii) Rate of reaction in Experiment 2 is **faster** than in Experiment 1.

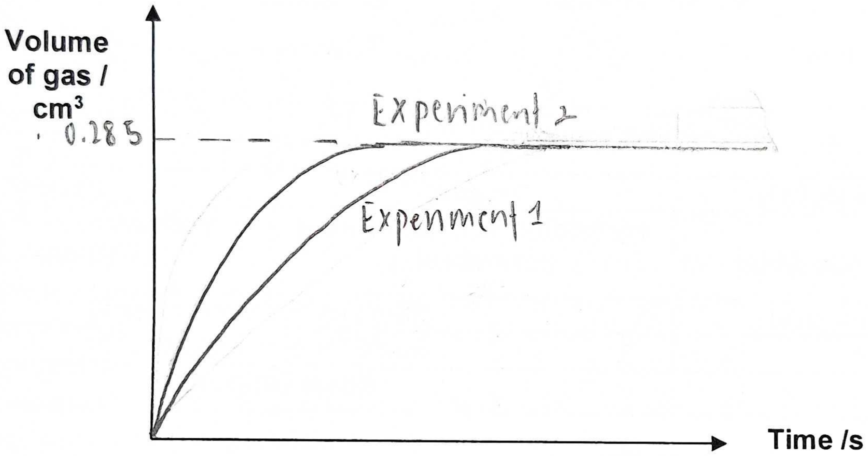
- This is because sulfuric acid is a **dibasic** acid and hydrochloric acid is a **monobasic** acid.

$HCl(aq) \rightarrow$ ......$H^+(aq) + Cl^-(aq)$......

$H_2SO_4(aq) \rightarrow$ ......$2H^+(aq) + SO_4^{2-}(aq)$......

- For the same concentration of acid used in both experiments, the **concentration** of ....H^+ ions.... in sulfuric acid used in Experiment 2 is **twice** that of hydrochloric acid used in Experiment 1.
- This explains why rate of reaction is **faster** in Experiment **2**.

(iii) On the same axes, sketch the graphs of volume of carbon dioxide gas produced against time for Experiment 1 and 2 when the limiting reactant is $MgCO_3$.



Case 2: Limiting reactant is Acid

Question:
If the acid is the limiting reactant instead of magnesium carbonate, as shown in the experiments 3 and 4 below.

	Experiment 3	Experiment 4
Mass of $MgCO_3$	excess	excess
Particle size of $MgCO_3$	powder	powder
Type of Acid (Basicity)	**Hydrochloric acid, $HCl(aq)$ (monobasic)**	**Sulfuric acid, $H_2SO_4(aq)$ (dibasic)**
Concentration of acid	1.00 mol/dm³ for both experiments	
Volume of acid	50.0 cm³ for both experiments	
Temperature	Room temperature	

(i) Explain how the **volume** of carbon dioxide gas produced would differ.
(ii) Explain how the **rate** of reaction would differ.
(iii) Sketch the graphs of rate of reaction for Experiments 3 and 4 on the same axes.

Answer:

(i)

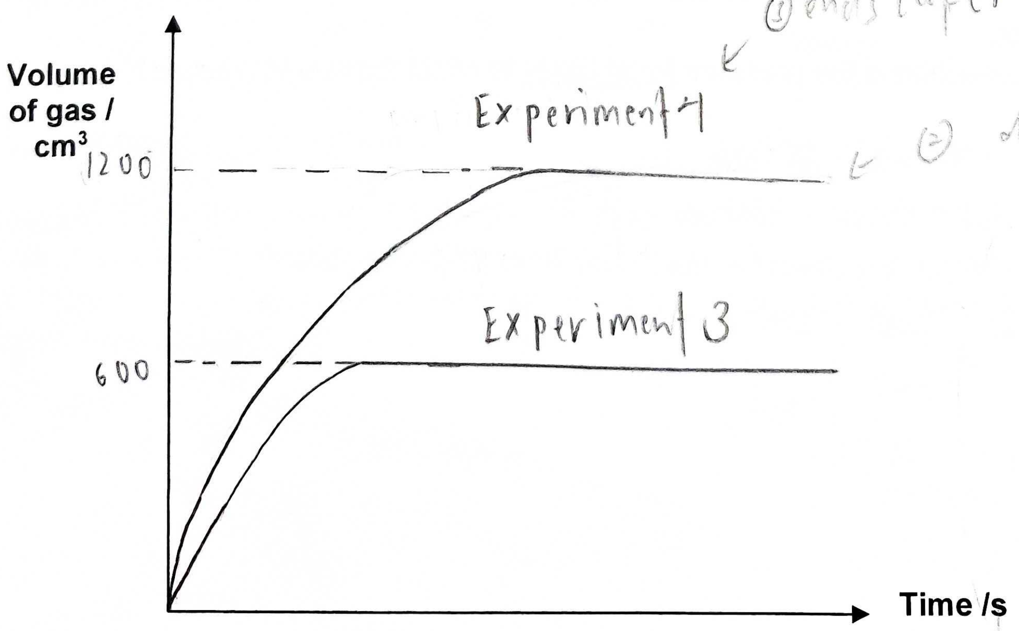
Experiment 3								
$MgCO_3(s)$	+	$2HCl(aq)$	→	$MgCl_2(aq)$	+	$H_2O(l)$	+	$CO_2(g)$
excess		1.00 mol/dm^3 50.0 cm^3						0.600 dm^3 at r.t.p
		0.0500 mol						0.0250 mol

Experiment 4								
$MgCO_3(s)$	+	$H_2SO_4(aq)$	→	$MgSO_4(aq)$	+	$H_2O(l)$	+	$CO_2(g)$
excess		1.00 mol/dm^3 50.0 cm^3						~~0.600~~ dm^3 at r.t.p ~~0~~ 1.200
		0.0500 mol						0.0250 mol

- Volume of carbon dioxide gas produced in Experiment 4 is **twice** that of Experiment 3.
- The ionic equation for both reactions in Experiment 3 & 4 is
  $MgCO_3 + 2H^+ \rightarrow Mg^{2+} + CO_2 + H_2O$
- Hydrochloric acid used in Experiment 3 is **monobasic** acid and sulfuric acid used in Experiment 4 is **dibasic** acid.
- This means that for the **same number of moles** of **acids** used, the number of moles of hydrogen ions in sulfuric acid is **twice** that of hydrochloric acid.
- Hence the number of moles and volume of carbon dioxide produced in Experiment 4 will also be **twice** that of Experiment 3.

(ii) Rate of reaction in Experiment 4 is **faster** than in Experiment 3.
- This is because sulfuric acid is a **dibasic** acid and hydrochloric acid is a **monobasic** acid.
- For the same concentration of acid used in both experiments, the **concentration** of H^+ ions in sulfuric acid used in Experiment 4 is **twice** that of hydrochloric acid used in Experiment 3.
- This explains why rate of reaction is **faster** in Experiment **4**.

(iii)



① ends later

② double the volume

E3. Effect of Pressure (of Gaseous Reactant)

- Pressure has very little effect on reactions involving only solid and liquid reactants.
- Changes in pressure will only affect those reactions involving gaseous reactants.
- Increase in pressure of the gaseous reactants **increases** the rate of reaction.

Explanation in terms of collision of reacting particles:

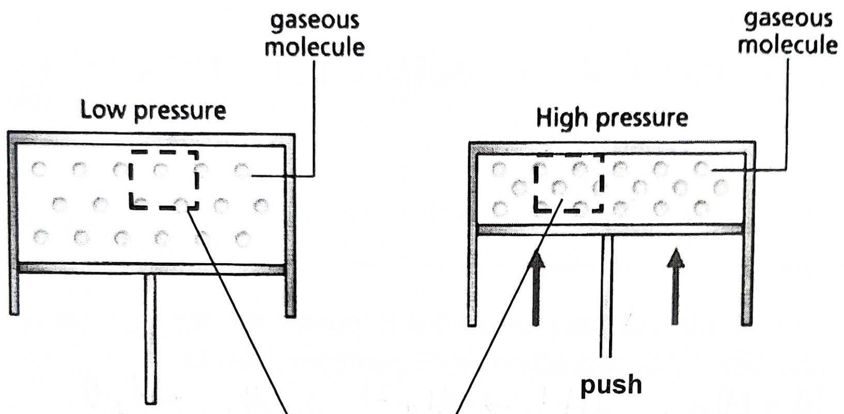


Fig. 18.12 At low pressure, particles are spread far apart.

Fig. 18.13 At high pressure, the same number of particles are closer together.

When pressure is increased, we can say that there are **same number reacting particles in a smaller volume.**

OR

more reacting particles in the same volume (per unit volume).
(explanation is similar to the effect of increase in **concentration**)

- When pressure of reaction mixture involving gaseous reactants is increased, the **number of reacting particles per unit volume increases.**
- **Frequency** of **effective collision** between reacting particles **increases.**
- Hence, rate of reaction **increases**.

Question:
In which reaction is the pressure **least likely** to affect the rate of reaction?

A $C(s) + CO_2(g) \rightarrow 2CO(g)$

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

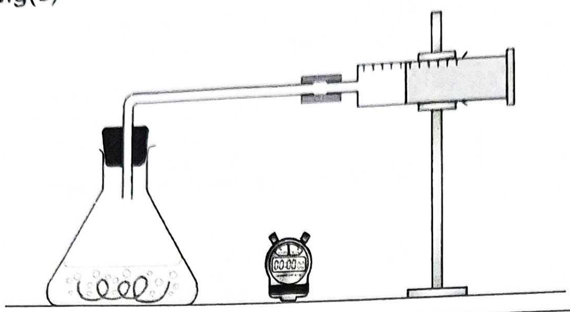
C $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

D $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

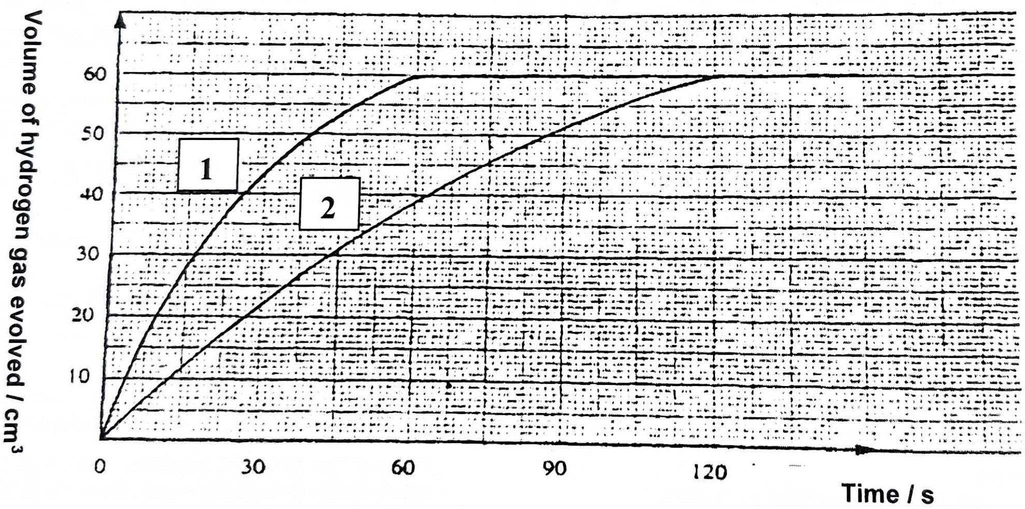
E4. Effect of Temperature

- Consider the reaction between magnesium and dilute hydrochloric acid

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$



	Experiment 1	Experiment 2
Mass of Mg ribbon	1.00 g	1.00 g
Concentration of acid	0.500 mol/dm^3	0.500 mol/dm^3
Volume of acid	30.0 cm^3	30.0 cm^3
Temperature	**50°C**	**30°C**



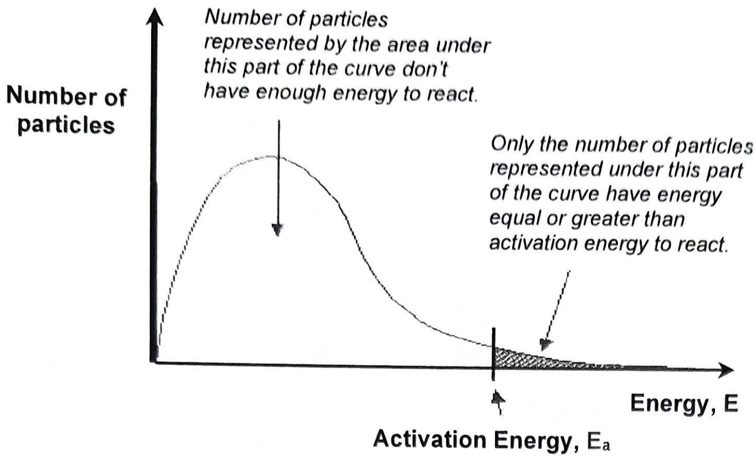
Interpretation of Graph:

1) Initial gradient of graph for Experiment 1 is <u>steeper</u> than that of Experiment 2. This shows that rate of reaction is <u>faster</u> when a higher temperature is used in Experiment 1.
2) Time taken for reaction to complete in Experiment 1 is <u>60 s</u> and in Experiment 2 is <u>120 s</u>.
3) Both reactions produce <u>60 cm^3</u> of hydrogen gas.

Explanation in terms of collision of reacting particles:

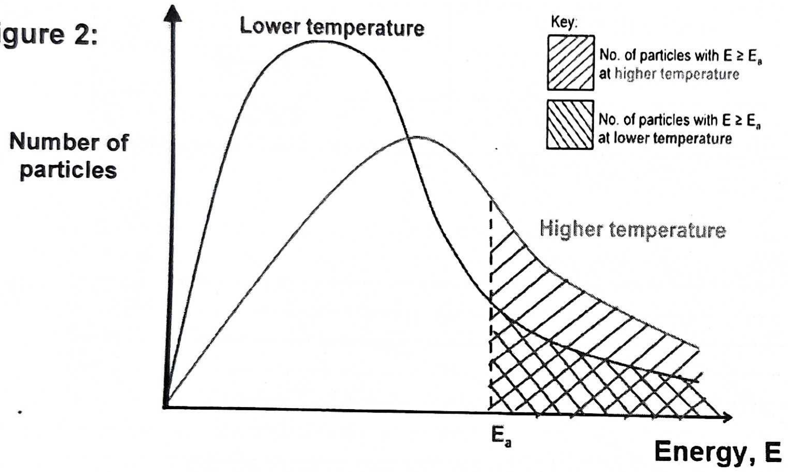
We make use of Maxwell Boltzmann Distribution Curve:

Figure 1:

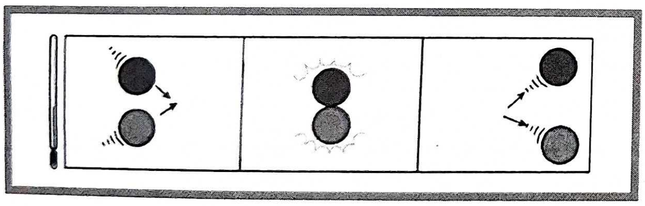


In **Figure 1**, the area under the graph represents the total number of particles involved in a chemical reaction.

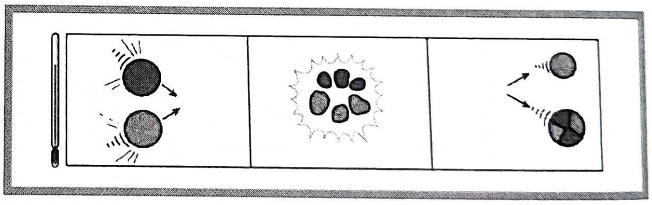
Figure 2:



In **Figure 2**, at higher temperature, there will be an increase in the number of particles possessing energy equal or greater than activation energy, E_a.



At low temperature, reacting particles **bounce apart** without a reaction.



At higher temperature, reacting particles **gain energy** and **move faster** and collide more often with greater force.

- When **temperature** of the reaction mixture is **increased** by heating, the **reacting particles** will **gain kinetic energy** and **move faster**.
- There will be an **increase** in **number of reacting particles** with energy **equal to** or **more than** activation energy.
- **Frequency** of **effective collisions** between **reacting particles**. Hence rate of reaction **increases**.

Note: Increase in temperature has **no effect** on the E_a value of the reaction.

Importance of Temperature:

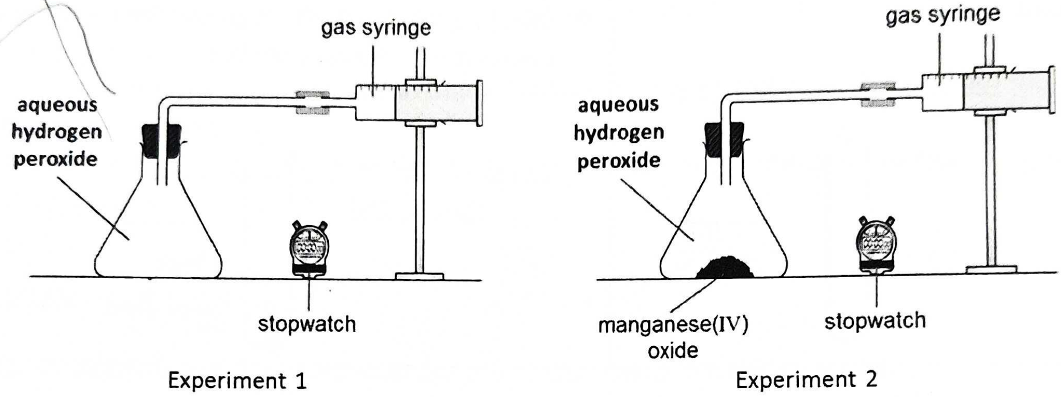
- Temperature is important in preserving food. The decay and decomposition of food is a chemical process. Food kept a low temperature slow down the reaction.
- Heat is commonly used in industries to speed up reactions. High temperatures are often used to make products as quickly as possible. Slow reactions are not economical.

E5. Effect of Catalyst

- A catalyst is a substance that **increases** the speed of a chemical reaction, without itself being chemically changed or used up at the end of the reaction.
- Other characteristics of catalysts include:
  - A **small** amount of a catalyst has a big effect on the speed of a reaction.
  - Many catalysts are **transition metals** or **compounds** of transition metals.
  - Most catalysts are specific in their action, i.e., only catalyse one type of reaction.

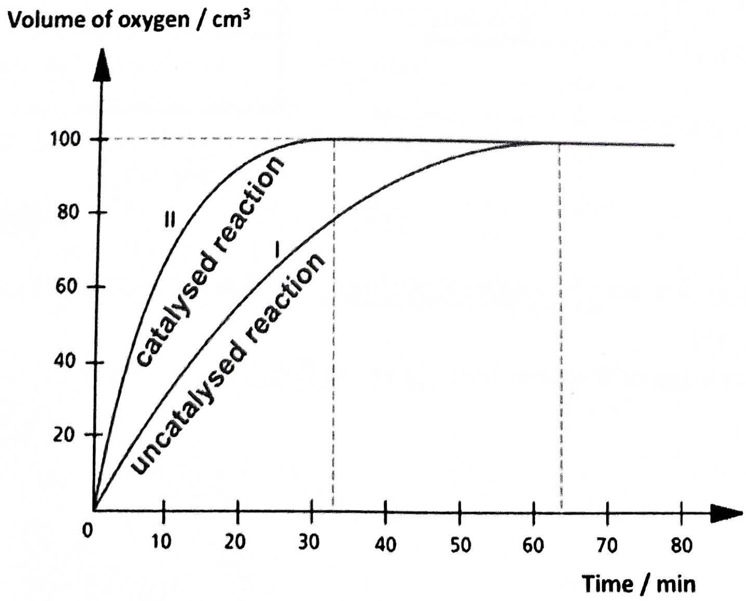
Example: Decomposition of aqueous hydrogen peroxide

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$



	Experiment 1	Experiment 2
Concentration of H_2O_2	1.00 mol/dm^3	1.00 mol/dm^3
Volume of H_2O_2	100 cm^3	100 cm^3
Temperature	Room temperature	Room temperature
Presence of MnO_2 (catalyst)	**No**	**Yes**

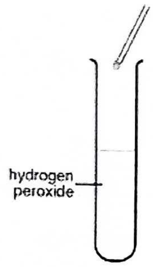
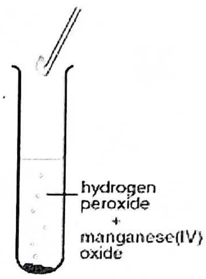
Results:



Interpretation of Graph:

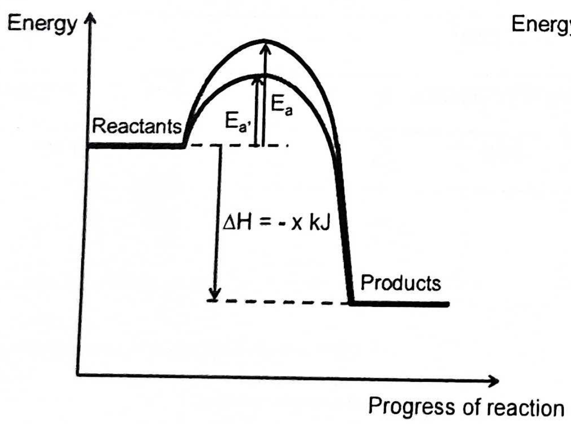
1) The **initial gradient** of graph for the **catalysed** reaction is **steeper**. Thus the speed of the catalysed reaction is **faster**.
2) Also, the catalysed reaction is **faster** because the **time taken** for completion of reaction is **shorter**.

Alternative way:

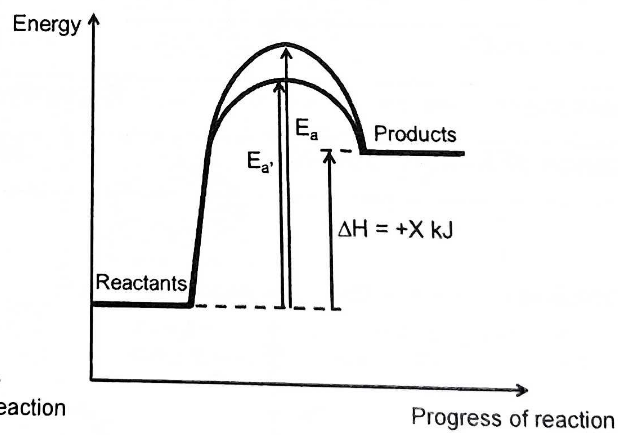
Without catalyst	With catalyst, manganese (IV) oxide
• If aqueous hydrogen peroxide is left standing, it decomposes into water and oxygen gas very slowly. 	• Upon addition of the **MnO_2** which acts as a catalyst, the decomposition of hydrogen peroxide is sped up. • The oxygen produced per unit time is higher, thus causing the glowing splint to burn more brightly. 

Explanation in terms of collision between reacting particles:

Exothermic



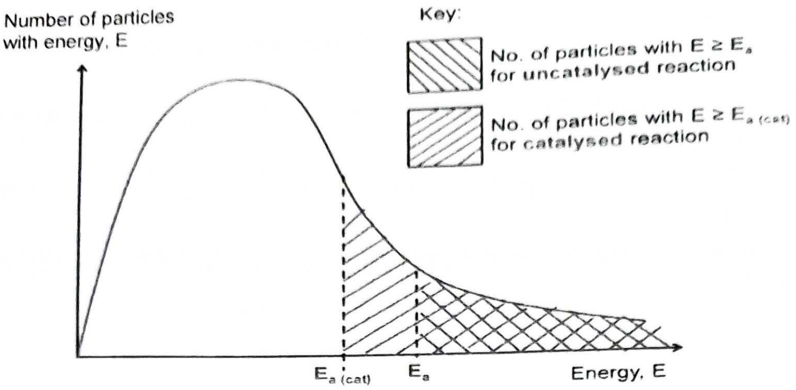
Endothermic



Note:

- The catalyst provides an **alternative pathway** for the reaction to occur with **lower activation energy**.
- The enthalpy change of the reaction **remains the same**.

Maxwell Boltzmann Distribution Curve on Effect of Catalyst



- A catalyst provides an alternative pathway for the reaction to occur lower activation energy.
- Thus the **number** of **reacting particles** having energy equal to or more than the activation energy increases.
- Hence, the frequency of effective collisions between **reacting particles** increases, and speed of catalysed reaction increases.

Applications of catalysts:

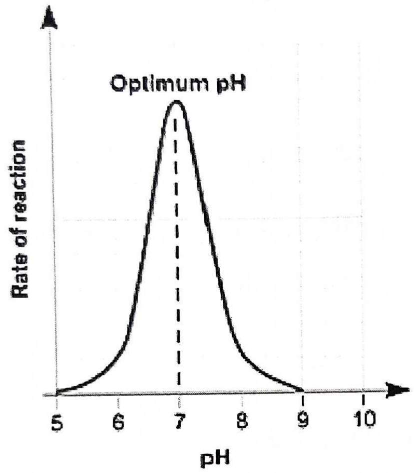
- Many catalysts are transition metals or compounds of transition metals.
- Catalysts are commonly used in industry to speed up chemical reactions. This saves energy and money as catalysed reactions require lower temperature and pressure.

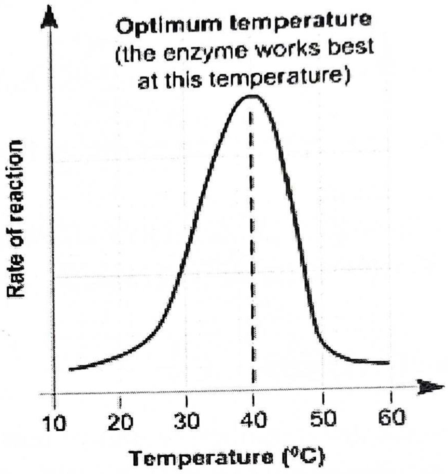
Chemical Reaction	**Catalyst**
Decomposition of hydrogen peroxide	Manganese(IV) oxide, MnO_2
Reaction between zinc and dilute sulfuric acid	Copper(II) sulfate, $CuSO_4$

Industrial Process	**Catalyst**
Haber Process *(industrial manufacture of ammonia)*	Finely divided iron, Fe
Contact Process *(industrial manufacture of sulfuric acid)*	Vanadium(V) oxide, V_2O_5
Hydrogenation *(addition of hydrogen to alkenes)*	Nickel, Ni
Cracking of petroleum *(industrial manufacture of alkanes and alkenes)*	Aluminium oxide, Al_2O_3

Enzymes (Biological Catalysts):

- Enzymes are biological catalysts, essentially proteins which catalyse biochemical reactions (reactions in living organisms).
- Enzymes are found in living things such as plants and animals. Most chemical reactions in plants and animals use enzymes.
- Most enzymes are highly specific in their action, i.e., only catalyse a very specific biochemical reaction.
- Enzymes work well only over a narrow pH range (pH 5 to 9) and over narrow temperature range (20°C to 40°C).



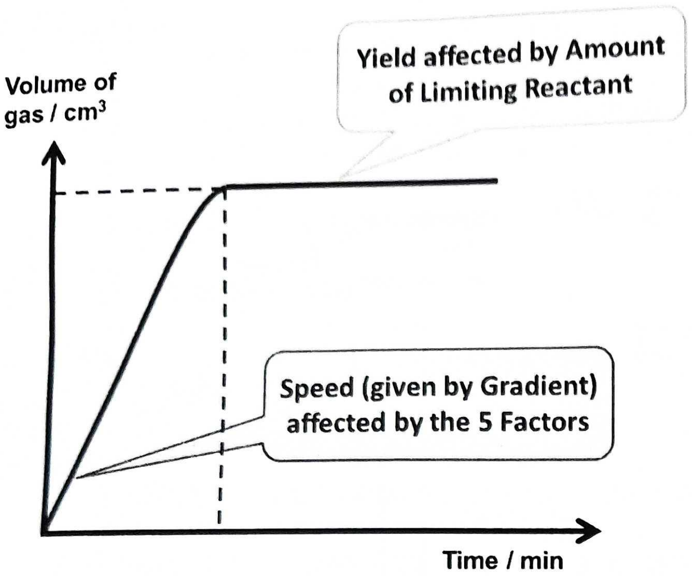


- When the temperature rises above the optimum temperature, the enzyme will be denatured, making the enzyme lose its ability to catalyse the reaction.
- When the temperature falls below the optimum temperature, the enzyme will become inactive, making the enzyme lose its ability to catalyse the reaction.

Examples & Uses of Enzymes:

- Fluid inside pitcher plants contains **certain enzymes** which speed up digestion of insects that fall into them.
- Digestion of food involves many enzymes, e.g. saliva contains an **enzyme, amylase** which speeds up digestion of starch in mouth.
- Antibiotics are compounds that kill harmful bacteria. They are produced by **enzymes in fungi**.
- Beer and wine are made from fermentation, a reaction that uses **enzymes in yeast** to convert starch or sugars into ethanol and carbon dioxide.
- Washing powders contain an **enzyme, subtilisin** which remove grease and blood stains by digesting them biologically. Heat cannot be used as this will denature the enzyme.

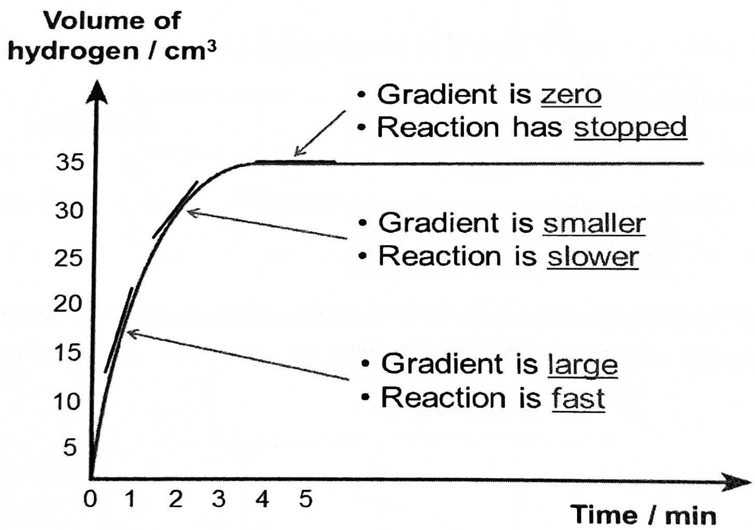
F. Important Notes



****Speed** of a reaction – is affected by any of the five factors (particle size, concentration, pressure, temperature, use of a catalyst)

****Yield** of a reaction – is affected by the **amount of limiting reactant** in the reaction.

Question:
Describe and explain how the speed/rate of reaction changes as the reaction progresses.



1st phase of graph:
Describe: At the **start** of the reaction, the **speed** of reaction is the **greatest**.
Explain: The **concentration** of the **reactants** is **highest** at the start.

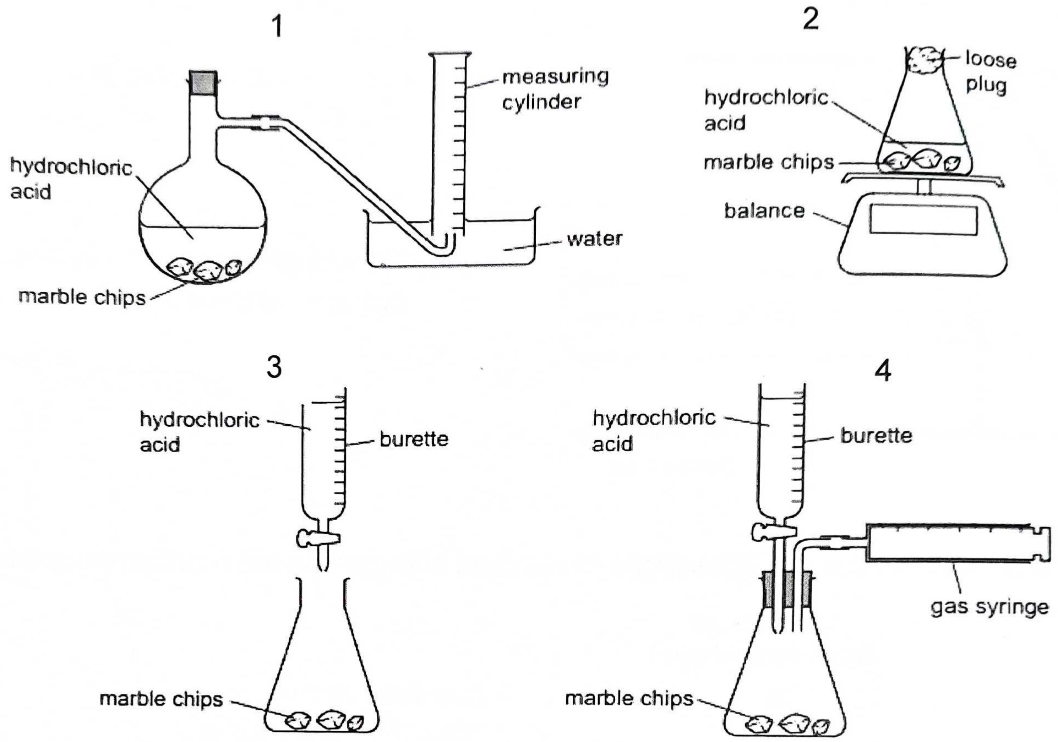
2nd phase of graph:
Describe: As the reaction **progresses**, the **speed** of reaction **decreases** with time.
Explain: The **concentration** of the **reactants decreases** as the reactants are being used up, thus **frequency of effective collisions between reacting particles decreases**.

3rd phase of graph:
Describe: Eventually, the reaction **stops**, the **speed** of reaction becomes **zero**.
Explain: The **limiting reactant** is **completely used up**.

Self-Check Exercise

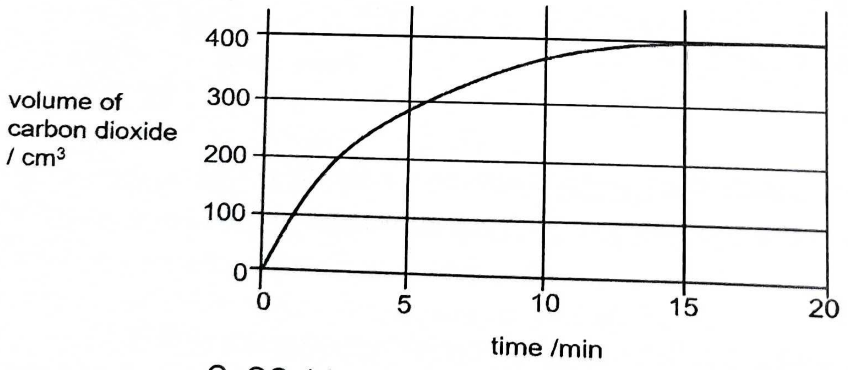
1. A student follows the rate of the reaction between marble chips, $CaCO_3$, and dilute hydrochloric acid.

   Which diagrams show apparatus that is suitable for this experiment?



A 1 and 2 B 1 and 3 C 1 and 4 D 1, 2 and 4

2. A student added excess hydrochloric acid to calcium carbonate. As the reaction proceeded, he measured the volume of carbon dioxide released, at room temperature and pressure and plotted a graph.



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

What can the student conclude from the information?

1. The number of moles of $CaCO_3$ used was 1.67×10^{-2}.
2. During the reaction, the rate steadily increased. ✗
3. After 15 minutes, the reaction has stopped. ✓

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

(C)

3. In the reaction between zinc and hydrochloric acid, the following changes could be made to the conditions.

1. increase the concentration of the acid
2. increase in particle size of the zinc
3. increase in pressure on the system
4. increase in temperature of the system

Which pair of changes will increase rate of reaction?

A 1 and 2
B 1 and 4
C 2 and 3
D 3 and 4

(B)

4. 1 mol of X reacts with 1 mol of Y to produce a gas.

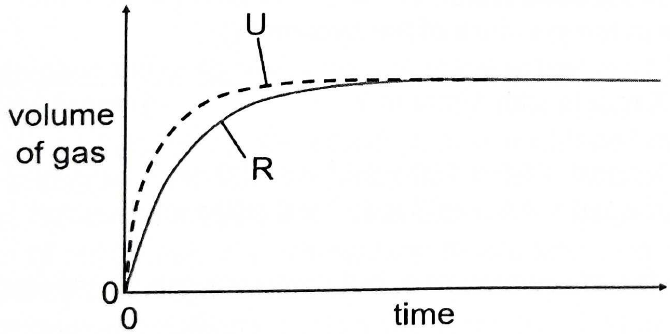
A student investigates the rate of this reaction by measuring the volume of gas produced at timed intervals.

He plots his results on a graph.

The table gives the volume and concentrations of X and Y that he uses to produce line R.

X		Y	
volume / cm^3	concentration in mol / dm^3	volume / cm^3	concentration in mol / dm^3
100	1.0	100	1.0

The student varies both volume and concentration of X and Y over a series of experiments.



The results of one of these experiments is plotted in the same graph.

Which row shows the volumes and concentrations that the student used to obtained line U?

	X		Y	
	volume / cm^3	concentration in mol / dm^3	volume / cm^3	concentration in mol / dm^3
A	25	2.0	100	1.0
B	100	1.0	50	2.0
C	100	1.0	200	0.5
D	400	0.5	400	0.5

(A)

5. How is the activation energy for a reaction between two gases changed when it is carried out either in the presence of a catalyst or when temperature is increased?

	change in activation energy	
	addition of catalyst	increase in temperature
A	decreases	decreases
B	decreases	stays the same
C	increases	decreases
D	increases	stays the same

(B)

....

	Answers	Explanation
1	D	1 & 4 – measuring of volume of gas at regular time interval 2 – measuring mass of reaction mixture at regular time interval
2	C	$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$ No. of moles of CO_2 formed at r.t.p. = 400 / 24000 =0.016667 mol $\frac{\text{No. of moles of } CaCO_3}{\text{No. of moles of } CO_2} = \frac{1}{1}$ 1. No. of moles of $CaCO_3$ givem = 0.016667 = 1.67×10^{-2} mol (√) 2. During the reaction, the rate steadily increased.(X) => rate decreases with time 3. After 15 minutes, the reaction has stopped. (√) => gradient reaches 0 at 15 minutes
3	B	1. increase the concentration of the acid (√) 2. increase in particle size of the zinc .(X) => decrease in total surface area of zinc Smaller total surface area exposed for H^+ ions to collide with, slower rate of reaction 3. increase in pressure on the system.(X) => no effect on rate as reactants are not in gaseous state 4. increase in temperature of the system (√)
4	A	Given 1 mol of X reacts with 1 mol of Y. For Graph R, no. of moles of X used = $1.0 \times 100 \times 10^{-3}$ = 0.100 mol no. of moles of X used = $1.0 \times 100 \times 10^{-3}$ = 0.100 mol Graph U shows faster reaction rate, but same volume of gas formed. For Option A: • higher concentration of Y, 2.0 mol/dm^3 used => faster rate of reaction • no of moles of X used = 0.100 mol ; no of moles of Y used= $0.5 \times 200 \times 10^{-3}$ = 0.100 mol => same as in Graph R => same volume of gas will be formed.
5	B	Catalyst provides alternative pathway by lowering activation energy of the reaction. Increase in temperature has no effect on the activation energy as well as the enthalpy change. The energy content of reactant(s) and product(s) also remain unchanged.