CHAPTER 6: RATES OF REACTIONS

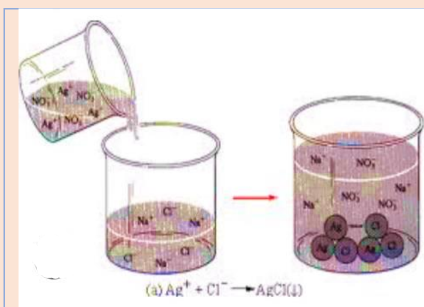
Learning Key words	Learning Outcomes
○ Chemical reaction ○ Reactants ○ Products ○ Catalysts ○ Activation energy ○ Reversible reaction ○ Equilibrium ○ Rates ○ Chemical kinetic	○ Understand and appreciate that chemical reactions take place at different rates. ○ Understand the effects of various factors on the rate of reaction and recognize that many of these reactions are reversible. ○ Understand the importance of reversible reaction in the Haber process (manufacture of ammonia) and contact process (manufacture of Sulphuric acid) in industrial processes.

COMPETENCY:

You will understand the effects of external conditions on the rate of reaction and how they can be explained in terms of kinetic particle model.

1.1. INTRODUCTION

In this chapter, you will not only understand what chemical reactions are but also to know that some reactions are **fast** and some are **slow,** hence they take place at different rates. Look at these examples.

		
When silver nitrate solution, is added to sodium chloride, precipitation of silver chloride occurs. This is a very fast reaction	Concrete setting. This reaction is very slow, since it will take some days for the concrete to fully harden	Rusting of an old car. This is usually a very slow reaction. It may take years for the car to rust completely away

But it is not always sufficient enough to know that a reaction is fast or slow. In industries that involve manufacturing of products from chemical reactants, there is need to know exactly how fast or slow a reaction proceeds and how long it will take to complete.

1.2. CHEMICAL REACTIONS

These are the process of breaking down or combining chemical reactants to form new substances called products.

Reactants are chemical substances that take part in a chemical reaction.

Products are chemical substances formed during a chemical reaction.

Rates of chemical reactions:

Refers to how fast the reactants are converted into products within unit time.

OR refers to how fast the products are formed during chemical reaction within unit time.

1.2.1. How to measure the rates of reaction

The table below shows how the rates of reaction can be measured with its standard units.

Measurement of rates of reaction	**Unit scale in:**
Mass of reactant used up per minute	grams per second (gs^{-1})
Volume of reactant used up per minute	Cubic centimeters per second (cm^3s^{-1})
Mass of product formed per minute	grams per second (gs^{-1})
Volume of product formed per minute	Cubic centimeters per second (cm^3s^{-1})

ACTIVITY: 6.2(a) To explore rates of chemical reaction.

What you need:

Access to textbooks, or internet

What to do: In groups

Brainstorm on the rates of chemical reactions of the following. Share and compare your findings with the rest of the class.

Discussion Questions:

a) Here are some chemical reactions that take place in the home. Arrange them in the decreasing order of rate of reaction (from the fastest to the slowest one)

(i) Boiling raw eggs to harden. (iii) Bread baking (v) a metal tin rusting

(ii) Fruit going rotten. (iv) Cooking gas burning (vi) A child growing up.

b) Arrange the following rates of travel beginning with the lowest rate.

(i) 80 kilometers per hour (ii) 2 kilometers per hour (iii) 30 kilometers per hour

1.2.2. Measuring the rates of reaction that involved volume of a gas

ACTIVITY: 6.2(b) To measure the rates of reaction that involves volume of a gas.

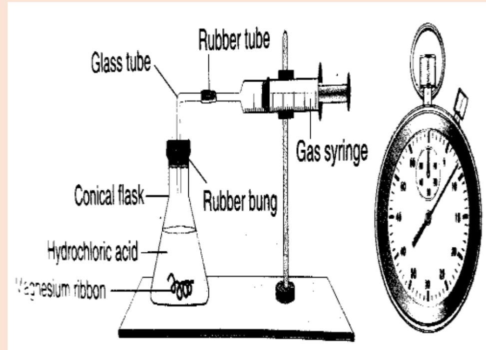
What you need:

Magnesium ribbon, dilute hydrochloric acid, conical flask, gas syringe, measuring cylinder, rubber, stop – clock, thermometer, glass tube and sand paper.

What to do: In groups

1. Arrange the set-up of the apparatus as shown in figure 6.2.2 below.

2. Measure 50 cm^3 of dilute hydrochloric acid using a measuring cylinder into a clean conical flask.
3. Clean 10g of magnesium with sand paper, drop it into the flask, and insert the stopper to the syringe and immediately start the stop – clock
4. Read and record the volume of the gas evolved in the syringe for every five seconds until the reaction stops.
5. Record your experimental results in the table below.



Time taken in seconds (s)	0	5	10	15	20	25	30
Volume of hydrogen gas in (cm^3)							

6.. Plot a graph of volume of hydrogen gas evolved with time taken.

Discussion Questions:

a) Identify the gas that evolved from the chemical reaction above.
b) Write an equation for the reaction leading to the formation of the gas in (a)
c) What was observed when cleaned piece of magnesium ribbon was added to the acid?
d) Why was magnesium ribbon cleaned before dropping it into the flask with acid?
e) From the graph. Determine the:
(i) Maximum volume of hydrogen gas evolved.
(ii) Time taken to produced 18.5cm^3 of hydrogen gas.

1.3. RATE CURVES

Most of the graphs in chemical kinetics are curves. This makes it difficult to determine the rate of reaction. This problem is sorted out by drawing a tangent along the curves at that given time taken. So, the gradient of the tangent is obtained as

Gradient (Rate of reaction) = $\frac{\textit{change in y-axis}}{\textit{change in x-axis}} = \frac{\Delta y\text{-}axis}{\Delta x\text{-}axis}$

Worked out example 1

Cleaned magnesium ribbon reacted with dilute hydrochloric acid to produce hydrogen gas according to the equation below.

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

The volume of hydrogen gas produced was collected in a gas syringe for every interval of five minutes as recorded in the table of results below.

Time taken in seconds (s)	0	5	10	15	20	25
Volume of hydrogen gas in (cm^3)	**0.0**	**1.5**	**2.5**	**3.0**	**3.0**	**3.0**

a) Plot a graph of volume of hydrogen gas evolved with time.

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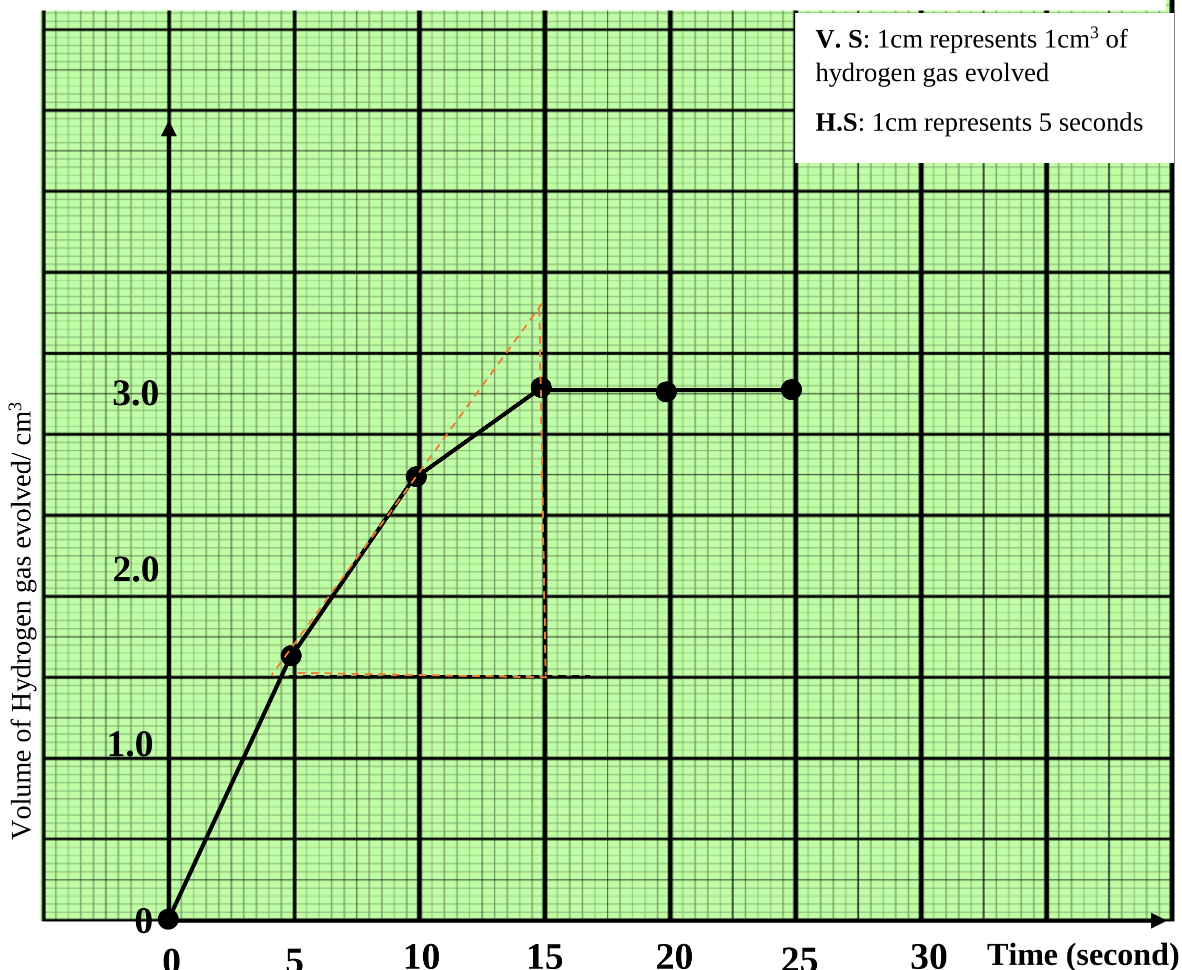
A GRAPH SHOWING VARIATION IN VOLUME OF HYDROGEN GAS EVOLVED WITH TIME



b) Use your graph to determine the rate of reaction that took place.

Gradient (Rate of reaction) = $\frac{\Delta y-axis}{\Delta x-axis} = \frac{15-10}{3.0-2.5} = \frac{5}{0.5}$ = 10 cm^{3}s^{-1}

1.4. Factors affecting the rates of chemical reactions.

In this section, you will understand the effects of various external conditions (factors) on the rates of chemical reactions. You will appreciate that these factors may increase or decrease rates of chemical reactions.

These factors that may increase or decrease the rates of chemical reactions include:

- Temperature of the medium
- Concentration of particles
- Surface area to volume ratio
- Presence of catalyst

1.4.1. Effect of temperature on Rates of reaction.

The rate of reaction increases with increase in temperature by increasing the frequency of particle collision, leading to faster rate of reaction. The reverse is true.

ACTIVITY:6.4(a) To investigate the effect of temperature on rates of reactions.

What you need:

Thermometer, white filter papers, stop – clock, measuring cylinders, tripod stand, 2molar hydrochloric acid, wire gauze, graph papers, 0.25molar sodium thiosulphate solution.

What to do: In groups

1. Obtain $20cm^3$ of sodium thiosulphate solution using a measuring cylinder into each of the four conical flasks labeled A, B, C and D.
2. Place a clean thermometer, into a conical flask A, warm the solution on a flame up to 25^oC.
3. Transfer the warmed conical flask A and place it on a white paper marked X.
4. Obtain $10cm^3$ of 2molar hydrochloric acid and add carefully into the warmed conical flask A content. Immediately start the stop – clock.
5. Note and record the time taken for the cross X on the white paper to disappear as viewed through the mouth of the flask.
6. Repeat procedures (2), (3), (4) and (5) at different temperature intervals of 30^oC, 35^oC, 40^oC, 45^oC and 50^oC. Use the table below to record your results.

Temperature / ^oC	25	30	35	40	45	50
Time taken for the cross to disappear (s)						
Reciprocal of time, $\frac{1}{t}$ (s^{-1})						

Discussion Questions:

a) Plot a graph of temperature against the reciprocal of time on a graph paper.
b) Describe the trend of your graph in (a)
c) Explain the trend of the graph described in (b).
d) Which other factors were kept constant during this experiment?
e) From your graph, state the relationship between rate of reaction and effect of temperature.

Self – checked Questions:

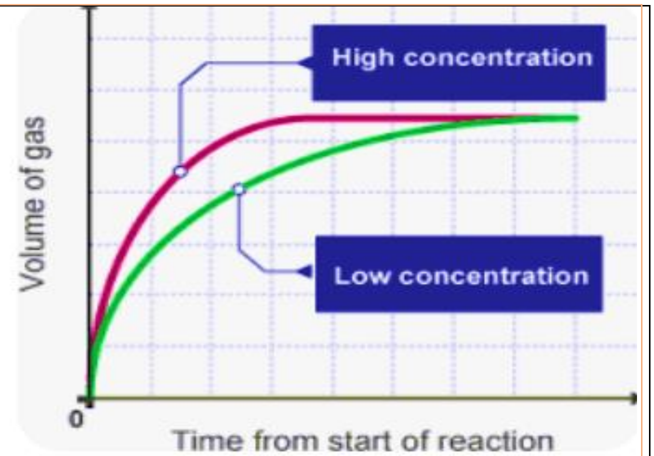
1. Using the knowledge of the effects of temperature on the rates of chemical reactions. Explain the following observations.

a) Smoked fish take many days before it starts decaying compared to fresh fish left outside at room temperature and pressure.
b) Frying pieces of cassava in cooking oil takes shorter time than cooking pieces of cassava with boiled water.
c) Milk easily gets spoilt and becomes sour when left at room temperature than the one stored in the refrigerator.

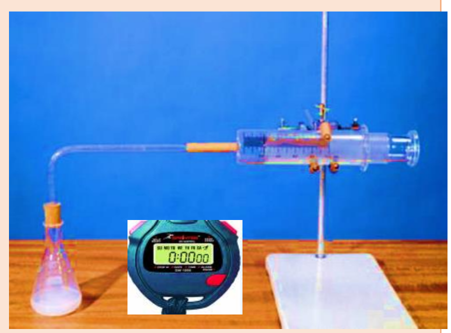
6.4.2. Effects of concentration of reactants on rates of reactions.

The closer the solute particles to another, the higher the concentration, the higher the frequency of particle collision.

The rate of reaction increases with increase in the concentration of solute (reactant) particles, the higher the frequency of collision of solute particles.



ACTIVITY: 6.2(b) To measure the rates of reaction that involves volume of a gas from different concentration of hydrochloric acid.

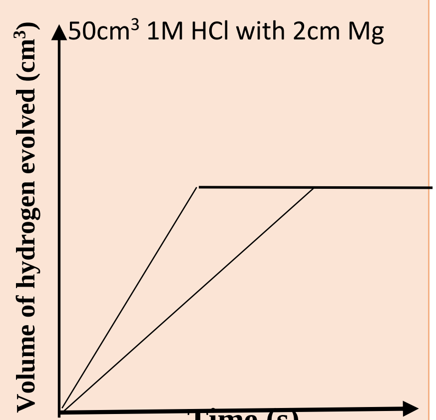


What you need

2 cm of clean magnesium ribbon, 1M hydrochloric acid, 0.5M hydrochloric acid, measuring cylinders, conical flasks, gas syringe, rubber bung, stop – clock.

What to do: In groups

1. Arrange the set – up of the apparatus as shown in the figure 6.2(b).
2. Measure $50cm^3$ of 1M hydrochloric acid using a measuring cylinder and transfer it into a clean conical flask.
3. Put 2cm of clean magnesium ribbon into the flask with the acid well fitted with a gas syringe. Immediately start the stop – clock.
4. Read and record the volume of the gas evolved in the syringe for every five seconds until the reaction stops.
5. Repeat procedures 2, 3 and 4 with 0.5M hydrochloric acid.
6. Use the table below to record your results.



Time taken in (s)	20	40	60	80	100	120	140
Volume of hydrogen gas in (cm^3) **in 1M HCl**							
Volume of hydrogen gas in (cm^3) **in 0.5M HCl**							

Discussion Questions:

a) Plot a graph of volume of a gas evolved from the concentration of the two hydrochloric acids used against time taken on the same graph paper.

b) State which of the two curves formed in (a) levels off:

(i) First before the other.

(ii) Last after the other.

c) Explain the reasons for your answers in (b) above.

d) Which other factors were kept constant during this experiment?

ACTIVITY: 6.2 (b) To investigate the effect of concentration of sodium thiosulphate on hydrochloric acid used.

What you need:

Apparatus: Stop – watch, conical flask, measuring cylinder, white tile, beaker.

Chemicals: Sodium thiosulphate solution, dilute hydrochloric acid.

What to do: In groups

1. Make a cross mark on a white tile or a white sheet of paper using a marker. Make sure the cross is visible when viewed from the top of the conical flask.
2. Measure $20cm^3$ of **0.02M** sodium thiosulphate solution using a measuring cylinder and transfer it into a clean conical flask.
3. Using another measuring cylinder, obtain $10cm^3$ of dilute hydrochloric acid and transfer it into a beaker containing sodium thiosulphate solution.
4. Immediately start the stop – watch (clock) and make observation by viewing the visibility of the cross mark from the top of the flask.
5. Stop the stop clock as soon as the cross mark disappears as seen from the top of the flask.
6. Note and record the time taken (in seconds) for the cross mark to disappear.
7. Repeat the procedures 2, 3, 4, 5 and 6 using **0.04M, 0.06M, 0.08M**, and **0.1M** sodium thiosulphate solution of different concentrations respectively.
8. Record your results obtained in a suitable table as shown below.

Concentration of sodium thiosulphate (M)	**0.02**	**0.04**	**0.06**	**0.08**	**0.1**
Time taken for the cross mark to disappear (s)					
Reciprocal of time, $\frac{1}{t}$ (s^{-1})					

Discussion Questions:

a) Plot a graph of concentration of sodium thiosulphate against the reciprocal of time.

b) Describe the shape of your graph in (a)

c) Explain the shape of the graph plotted in (a).

IMPORTANT NOTES.

The reaction of dilute hydrochloric acid with sodium thiosulphate, produces yellow precipitates of Sulphur. The amounts of Sulphur produced can be investigated by placing a flask of this mixture on a cross mark on a white tile. The time taken for the mark to disappear when viewed from the top is note and recorded.

Equation:

$Na_2S_2O_{3\,(aq)} + 2HCl_{(aq)} \longrightarrow S_{(s0} + SO_{2\,(g)} + 2NaCl_{(aq)} + H_2O_{(l)}$

If the concentration of dilute hydrochloric acid or sodium thiosulphate is manipulated, the time taken for the cross mark to disappear is noted and recorded.

ACTIVITY:6.2(c) To investigate the effect of concentration of sodium thiosulphate on hydrochloric acid used.

What you need:

Apparatus: Stop – watch, conical flask, measuring cylinder, white tile, beaker.

Chemicals: Sodium thiosulphate solution, dilute hydrochloric acid and water.

What to do: In groups

Make a cross mark on a white tile or a white sheet of paper using a marker. Make sure the cross is visible when viewed from the top of the conical flask.

1. Using a measuring cylinder, place 50cm^3 of sodium thiosulphate solution in a clean beaker
2. Measure 5cm^3 of 2M hydrochloric acid using a measuring cylinder and transfer it into a beaker containing sodium thiosulphate solution.
3. Immediately start the stop – watch (clock) and make observation by viewing the visibility of the cross mark from the top of the flask.
4. Stop the stop clock as soon as the cross mark disappears as seen from the top of the flask.
5. Note and record the time taken (in seconds) for the cross mark to disappear.
6. Repeat the procedures 2, 3, 4, 5 and 6 using 40cm^3, 30cm^3, 20cm^3, 10cm^{3}of sodium thiosulphate solution followed by adding 10cm^3, 20cm^3, 30cm^3, and 40cm^{3}of distilled water into the thiosulphate solution respectively.
7. Record your results obtained in a suitable table as shown below.

Experiment number	**1**	**2**	**3**	**4**	**5**
Volume of sodium thiosulphate (cm^3)	50	40	30	20	10
Volume of distilled water (cm^3)	0	10	20	30	40
Volume of hydrochloric acid /cm^3	5	5	5	5	5
Time taken / s					
Reciprocal of time, $\frac{1}{t}$ (s^{-1})					

Discussion Questions:

a) Plot a graph of volume of sodium thiosulphate against time.
b) Describe the shape of the graph.
c) Explain the shape of the graph.

d) Using your graph, determine the gradient (rate of reaction) from the graph and state its unit

e) Why does the cross mark disappear in the experiment?

1.3. SURFACE AREA (SIZE OF THE PARTICLES)

In this section, you will appreciate, the concept of "surface area". For instance, Irish potatoes, one piece of it is sliced into eight (8) pieces and the other unsliced. Which of the two potatoes will take shorter time to get ready when they fried in cooking oil.

The rate of reaction increases with increase in the surface area. This is because, increase in surface increases the frequency of particle collision, leading to faster rate of reaction.

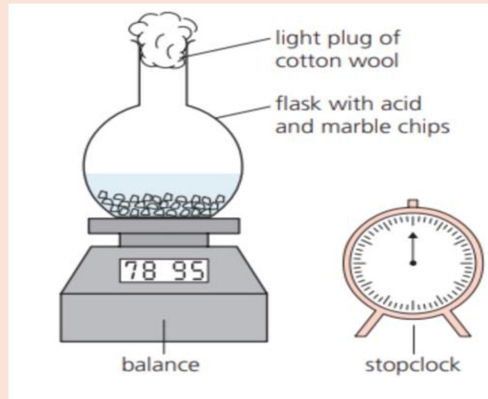
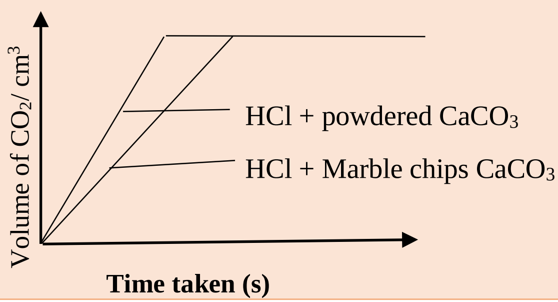
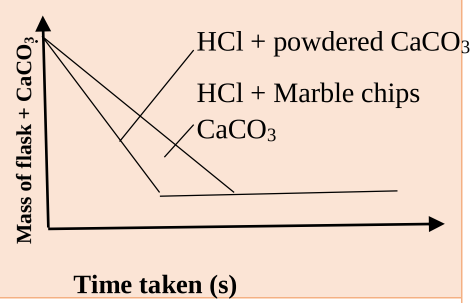


Fig 6.3. Surface area of Marble chip with acid

For example:

Calcium carbonate, exists as powdery solids and large lumps (Marble chips). Calcium carbonate reacts more rapidly with dilute hydrochloric acid when in powdery form than when in form of marble chips. This is because, powdery form increases surface area increasing the frequency of its particles colliding with the acid.





ACTIVITY: 6.3(a) Exercise 1

What to do: In groups

Discuss the how the particle size of calcium carbonate (lumps and powdery forms) affects the rates of reaction when dilute hydrochloric acid is added to each size of the salt. Share and compare your findings with the rest of the class.

Discussion Questions:

Below is a table of results obtained from the reaction between calcium carbonate of different sizes with dilute hydrochloric acid.

Time taken (s)	0	5	10	15	20	25	30	35
Volume of CO_2 from lumps of $CaCO_3$ + hydrochloric acid	0	8	11.5	16.5	18.5	20.5	20.5	20.5

Volume of CO_2 from powdery $CaCO_3$ + hydrochloric acid	0	13	16.5	18.5	20.0	20.5	20.5	20.5

a) Plot on the same axis a graph to show variation in the volume of carbon dioxide evolved with time.

b) Describe the changes in the rate of reaction during the experiment for the two curves.

c) Explain the changes in the rate of reaction as reflected by two curves obtained in (a).

d) From the graph:

(i) What volume of carbon dioxide gas in both lumps and powdery calcium carbonate would be produced in 23 seconds?

(ii) What would the rate of reaction in both curves after 15 seconds?

(iii) What would be the time taken for 17 volume of carbon dioxide gas to be formed in both lumps and powdery calcium carbonate?

ACTIVITY: 6.3(b) Exercise 2.

What to do: In groups

Discuss the how the particle size of calcium carbonate (lumps and powdery forms) affects the rates of reaction when dilute hydrochloric acid is added to each size of the salt. Share and compare your findings with the rest of the class.

Discussion Questions:

Below is a table of results.

Time (minutes)	0	1	2	3	4	5	6	7
Mass of flask +powdery $CaCO_3$ + HCl (g)	30.0	23.1	16.5	12.2	7.0	5.5	4.0	4.0
Mass of flask +lumps $CaCO_3$ + HCl (g)	30.0	25.2	18.0	10.5	8.5	6.8	4.0	4.0

a) Plot on the same axis a graph to show the changes in the mass for lumps and powdery calcium carbonate with time.

b) From the graph in (a) above. State which curves indicates the:

(i) slower rates of reaction between the limestone and the acid.

(ii) faster rates of reaction between the limestone and the acid.

c) Explain your answers in (b).

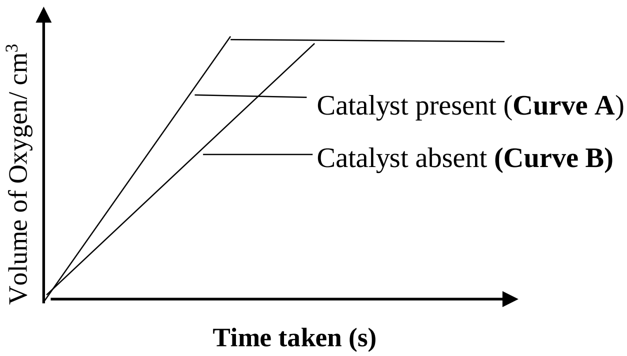
d) Compare the trends in the shapes of the two graphs (curves).

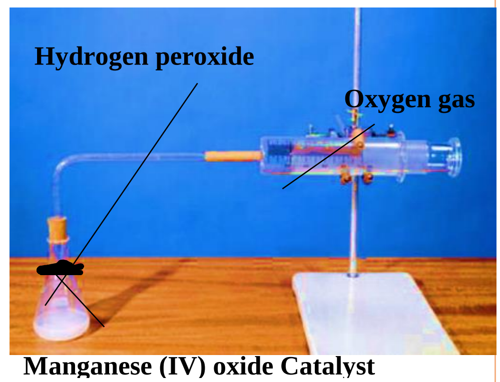
e) Explain the effect of surface area affect the rate of chemical reaction above.

f) Write an equation for the reaction that took place.

1.4. PRESENCE OF CATALYST

CATALYST: Is a substance which speeds up the rate of chemical reaction and remains chemically unchanged at the end of the reaction. The role of a catalyst is to lower the **activation energy** required to initiate a chemical reaction and to consequently increase the frequency of particle collision leading to faster rate of reaction.





From the graphs above we can therefore, say for:

Curve A (Description)

The rate of decomposition of hydrogen peroxide increases rapidly at faster rate until to a point where the reaction levels off early before curve B and remains constant.

Explanation:

The faster rate of reaction, is due to the presence of a catalyst that facilitated faster frequency in the particle collision hence increasing the rate of reaction rapidly.

Curve B (Description)

The rate of decomposition of hydrogen peroxide increases gradually at slower rate until to a point where the reaction levels off after the curve A and remains constant.

Explanation:

The slower rate of reaction is due to the decrease in the concentration of hydrogen peroxide.

ACTIVITY:6.4(a) To investigate the effect of manganese (IV) oxide catalyst on the rate of decomposition of Hydrogen peroxide.

What you need:

Measuring cylinder of 10mls, conical flasks, gas syringe, manganese (IV) oxide, hydrogen peroxide, stop – watch, spatula and a wooden splint.

What to do: In groups

a) Obtain two cleaned conical flasks and label them as flask A and Flask B.
b) Put one spatula endful of manganese (IV) oxide in to a conical flask A and close it.

c) Measure $5cm^3$ of hydrogen peroxide and add it into a flask A fitted with a gas syringe. Immediately start the stop – watch and record the volume of the gas evolved for every 20 seconds until the reaction stops.

d) To the flask B, add $5cm^3$ of hydrogen peroxide and connect it to the gas syringe. Immediately start the stop – watch and record the volume of the gas evolved for every 20 seconds until the reaction stops.

e) Record your results for the two conical flasks in the table below:

Time in seconds (s)	20	40	60	80	100	120
Volume of a gas from flask A/cm^3						
Volume of a gas from flask B/cm^3						

Discussion Questions:

a) Write an equation for the decomposition of hydrogen peroxide.

b) Plot a graph on the same axis to show the changes in the volume of oxygen gas evolved from the decomposition of hydrogen peroxide in the two flasks with time.

c) From the graph in (b) above. State which curves indicates the:

(i) Slower rates of reaction.

(ii) Faster rates of reaction

d) Explain your answers in (c).

e) State the role of manganese (IV) oxide in the flask A.

1.5. COLLISION THEORY AND ACTIVATION ENERGY

As you have already seen, you can increase the rate of reaction by increasing the:

- Temperature of the system.
- Concentration of a reactant.
- Surface area of a solid reactant

In some situations, as seen in **flour mills**, an increase in any of these factors above can lead to faster and dangerous reaction that can lead to **explosion**. Flour particles are very tiny, so giving them large surface area, if there is a lot of flour dust in the hot air, a spark may occur from the machine causing an explosion.

Fig 6.5 An explosion in flour mills

The Collision Theory: It states that for a chemical reaction to occur, the reactant molecules must collide with sufficient energy in the correct orientation. The reactant molecules need to overcome certain energy barrier called the **Activation energy.**

If the collision between reactant molecules does not provide sufficient energy to overcome this barrier, then products will not be formed. In essence, the products are only formed if the collision provides enough energy and the reactant molecules are in the correct orientation.

Activation energy: Is the minimum amount of energy required to be overcome by reacting molecules to form products.

It is also referred to the minimum amount of energy required to break the bonds in the reacting molecules to form products.

The collision theory is governed but not limited by the following principles:

1. **Collision:** Reactant molecules must collide with each other in order for a chemical reaction to occur.
2. **Orientation:** Each colliding particle must have the correct orientation relative to the other for the reaction to proceed.
3. **Energy:** Each colliding molecule must have enough energy to overcome the activation energy barrier for the reaction to occur.

ACTIVITY: 6.5(a) To explore particle collision Theory and Activation energy.

What you need:

Reference textbooks or internet.

What to do: In groups

Use reference textbooks or internet to search for information about particle collision theory and activation energy during chemical reactions. Share and compare your findings with the rest of the class.

Discussion questions:

a) What is meant by:
(i) Collision theory? (ii) Activation energy?
b) What are the three major principles of collision theory?
c) Explain using the knowledge of collision theory why any decrease in:
(i) concentration of reactant molecules lowers the rate of reaction.
(ii) temperature decreases the rate of reaction.
(iii) surface area decreases the rate of the reaction.

1.6. REVERSIBLE AND EQUILIBRIUM REACTIONS

Reversible reactions: Refers to the reaction that proceeds in both forward and reverse directions. The reverse of this type of reaction is called **Irreversible reaction.**

Irreversible reactions: Refer to the reaction which proceeds in only forward direction to form products.

Equilibrium reactions: Refer to the reaction where the rate of forward reaction equals the rate of reverse reaction with no net change in the concentrations of both reactants and products. An equilibrium state known as **dynamic Equilibrium** is established**.**

Factors that affect the positions of Equilibrium reactions.

The state of Equilibrium system, for the rate of forward and reverse reactions is disturbed by changes in the following factors such as **temperature, pressure and concentration** hence the positions of Equilibrium will adjust itself in order to manage the disturbance and restore equilibrium back to normal. Check for details in book four (4) under **Heat of reactions**.

Equilibrium (Reversible) reaction of Anhydrous copper (II) sulphate.

Hydrated copper (II) sulphate, $CuSO_4.5H_2O$ are blue crystalline solids, because it contains water of crystallization. When heated, it forms Anhydrous copper (II) sulphate as white powdered solids, because it lost its water of crystallization.

Hydrated Copper (II) Sulphate $\rightleftharpoons$ Anhydrous Copper (II) sulphate + Water

$$CuSO_4.5H_2O_{(s)} \rightleftharpoons CuSO_{4\,(s)} + 5H_2O_{(l)}$$

When water is added to white anhydrous copper (II) sulphate, blue crystalline solid of hydrated copper (II) sulphate is regenerated. This is a reversible reaction.

$$CuSO_{4\,(s)} + 5H_2O_{(l)} \rightleftharpoons CuSO_4.5H_2O_{(s)}$$

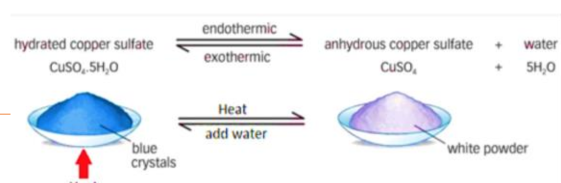
ACTIVITY 6.6(a) To investigate the reversible reaction of anhydrous copper (II) sulphate with water.

What you need:

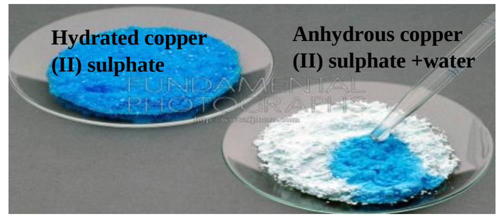
Hydrated copper (II) sulphate, heat source, spatula, boiling tube, petri dish, dropper, water.

What to do: In groups

1. Heat one spatula end of hydrated copper (II) sulphate crystals in a dry boiling tube, for five minutes. Note what happens.
2. Pour the solid residue obtained in a clean dry petri dish, allow it to cool down. Note the colour.
3. Add 4 drops of water to the cooled solid residue in the dish. Note what happens.



Discussion Questions:



a) State what was observed when:

(i) Hydrated copper (II) sulphate crystals we

(ii) Water was added to white anhydrous cop

b) Explain your observations in (a) (i) and (ii)

c) What is meant by the term "**Reversible reaction?**"

Applications of Reversible reactions during Industrial processes.

a) Manufacture of Sulphuric acid (By Contact process)

In this section, you will appreciate the reversible reactions in the manufacture of Sulphuric acid. Different uses of Sulphuric acid, put high demands, no wonder it should be produced on large scale to meet its demand.

ACTIVITY 6.6(b) To explain how Sulphuric acid is manufactured on large scale by Contact process

What you need:

Textbooks, or internet.

What to do: In groups

1. Research using any reference textbook, or internet to explain how pure Sulphuric acid is manufactured by Contact process. Share and compare your findings with the rest of the class.
2. **Discussion Questions**:

Manufacturing plant for Sulphuric acid

a) Explain how Sulphuric acid is manufactured by contact process in the following stages:
   (i) Stage 1: Combustion of Sulphur:
   (ii) Stage 2: Conversion and purification of Sulphur dioxide to Sulphur trioxide.
   (iii) Stage 3: Production of oleum.
   (iv) Stage 4: Production of pure Sulphuric acid from oleum.

b) Design a flow chart to show the different stages involved during contact process of Sulphuric acid.

Uses of Sulphuric Acid:

ACTIVITY 6.6(c) To explore the different uses of Sulphuric acid.

What you need:

Access to textbooks or internet.

What to do: In groups

Research using reference textbooks or internet to explore the different uses of Sulphuric acid in nature. Share and compare your findings with each other.

Discussion Questions:

a) State the different uses of Sulphuric acid in nature.
b) Explain how an acid rain is formed from Sulphur dioxide.

b) Manufacture of Ammonia (By Haber Process)

ACTIVITY 6.6(d) To explain the formation of ammonia by Haber process.

What you need:

Access to textbooks or internet.

What to do: In groups

Research using any reference textbook, or internet to explain how pure Ammonia is manufactured by Haber process. Share and compare your findings with the rest of the class.

Discussion Questions:

a) Name the raw materials used in the Haber process of Ammonia.
b) Explain how ammonia is manufactured on large scale.
c) Write word and chemical equations for the reversible reaction leading to the formation of ammonia.

Uses of Ammonia

ACTIVITY: 6.6(e) To explore the different uses of Ammonia compounds.

What you need:

Access to textbooks or internet.

What to do: In groups

Research using reference textbooks or internet to explore the different uses of ammonia compounds in nature. Share and compare your findings with each other.

Discussion Questions:

a) State the different uses of ammonia compounds in nature**.**

(i) Fertilizer production
(ii) Cleaning products.
(iii) Industrial applications.
(iv) Water treatment**.**
(v) Textile industry.
(vi) Food preservation

b) Explain how pure Nitric acid is manufactured from ammonia gas.

SAMPLE OF ACTIVITY OF INTEGRATION

SCENARIO:

Most of the sweet bananas, mangoes, and pineapples are largely grown in the central, south western and Eastern areas of Uganda. These fruits are always harvested when they are still green in colour for easy transportation. In the market places, these fruits slowly begin to turn yellow, orange or red indicating that they have ripened. But during the rainy season, these fruits tend to take long to ripen compared to dry season. This affects the rate at which they are sold to consumers, so the sellers and farmers are affected financially

TASK

Guide both the sellers and farmers of these fruits on how they should quicken fruits ripening during the rainy season. In your guidance.

a) Help them to know the factors that are affecting the delay of fruits ripening.
b) Guide them on how these factors can be improved in order for their fruits to ripen faster during rainy season.