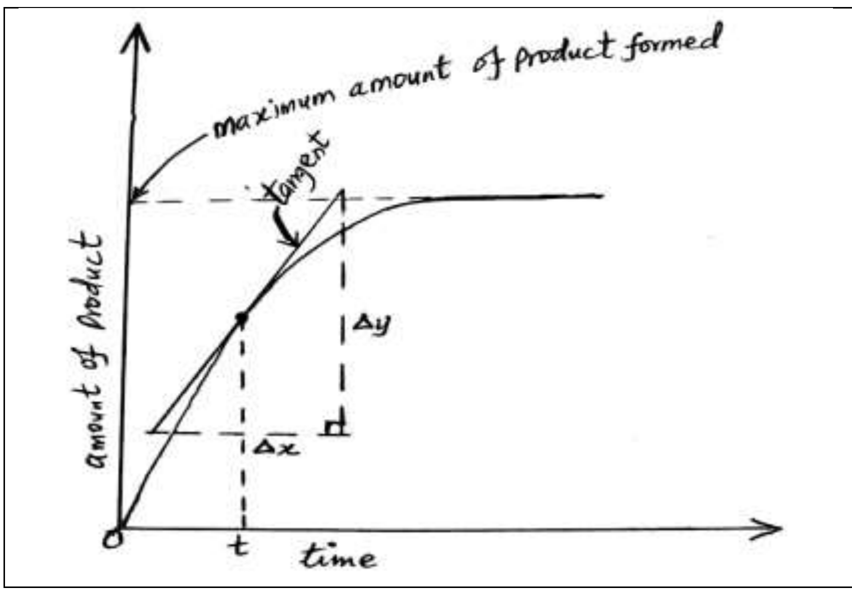
Rates of Chemical Reactions

- Different chemical processes (reactions) take different times to go into completion.
- For example,
  (i) rusting of an iron nail left in the open air takes some days, months or years.
  (ii) formation of precipitate takes few seconds or minutes.
- We can know how slow or fast chemical processes (reactions) are by:
  (i) **either** measuring the amounts of **product formed and time taken to form it**.
  (ii) **or** measuring the amounts of **reactants used up and time taken to use it up**.
- Thus, *the **rate of a chemical reaction** refers to the amount of products formed per unit time in a chemical reaction.*
  ***Or** we can define it as the amount reactants used up per unit time in a chemical reaction*.
- The rates of chemical reactions can be followed by considering some **physical property** of the reaction which **changes with time**.
  Some of the physical properties in reaction that may change with time include:
  (i) Volume of gas formed/produced.
  (ii) Amount of precipitate formed.
  (iii) Intensity of colour formation/disappearance.
  (iv) Mass of product formed/reaction mixture (reactants) used up.
- The units of rates of chemical reactions entirely depends on the physical property of reaction being investigated with time. Thus we have $cm^3 s^{-1}, gs^{-1}, mol\ s^{-1},\ mol\ min^{-1}$ etc.
- Graphically, we can illustrate the changes in amounts (mass, volume or moles) of products or reactants with time for a particular reaction.
  (a) For the case of changes in the amount of products with time, we have;

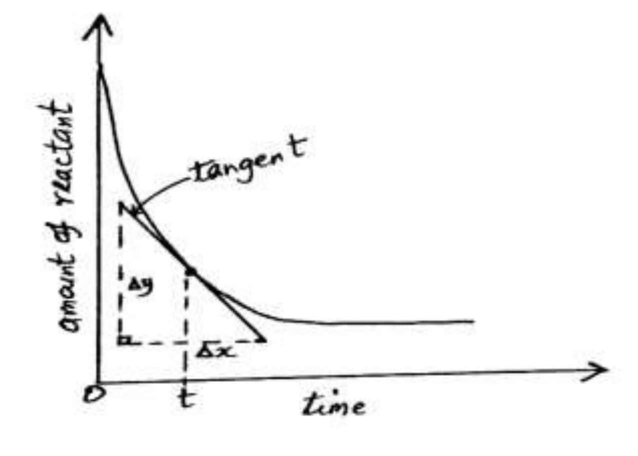


The rate of the reaction at time, t, is obtained by drawing a tangent at time, t, and obtaining its slope/gradient. Slope gives the rate of the reaction at time, t.

Rate $= \frac{\Delta y}{\Delta x}$

$= \frac{change\ in\ amount}{change\ in\ time}$

(b) Similarly, for changes in amounts of reactants with time, we have;



The rate of the reaction at time, t, is obtained by drawing a tangent at time, t, and obtaining its slope/gradient. Slope gives the rate of the reaction at time, t.

$$\text{Rate} = \frac{\Delta y}{\Delta x}$$
$$= \frac{change\ in\ amount}{change\ in\ time}$$

- ❖ Factors which affect the rate of chemical reactions are:
  - *(i) Concentration of the reactants*
  - *(ii) Temperature of the reaction mixture*
  - *(iii) Surface area (particle size) for* ***solid*** *reactants*
  - *(iv) Presence of a catalyst*
  - *(v) Pressure for* ***gaseous*** *reactants*
  - *(vi) Light for* ***photochemical*** *reactions*

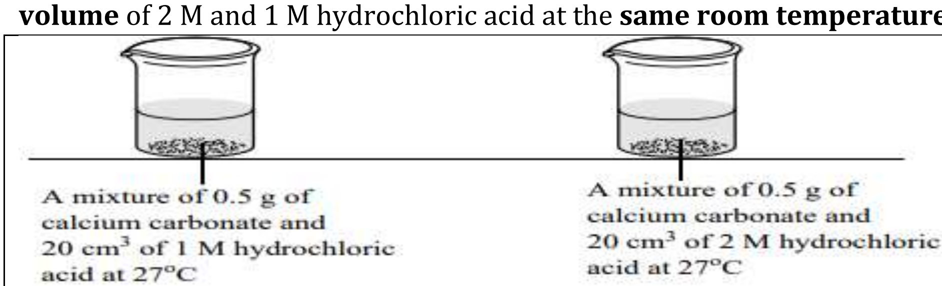
- ❖ Detailed discussions on each of the above factors is as below:

A. Effect of concentration of the reactants

- The **higher** the concentration of the reactants, the **faster** the rate of reaction.
- This is because at higher concentrations, **more** reactant particles are **so close** to each other.
- This results into the chances of them colliding being **higher** and more **frequent.**
- Hence faster reaction rate.

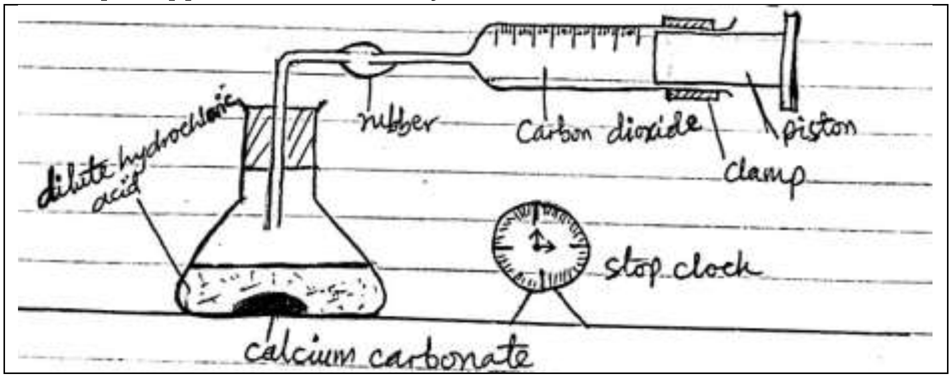
- ❖ The effect of concentration of the reactants can be illustrated using the following reactions:
  - *(i) Reaction of calcium carbonate with dilute acids at different concentrations of the acid.*
  - *(ii) Reaction of metals (e.g. Mg and Zn) with dilute acids at different concentrations of the acid.*
  - *(iii) Reaction of sodium thiosulphate with dilute acids at different concentrations of the acids.*
  - *(iv) Decomposition of the* ***same volume*** *of hydrogen peroxide of* ***different concentrations****.*

- ❖ For example;
  - Consider the reaction of the **same mass** of calcium carbonate **separately** with the **same volume** of 2 M and 1 M hydrochloric acid at the **same room temperature**.

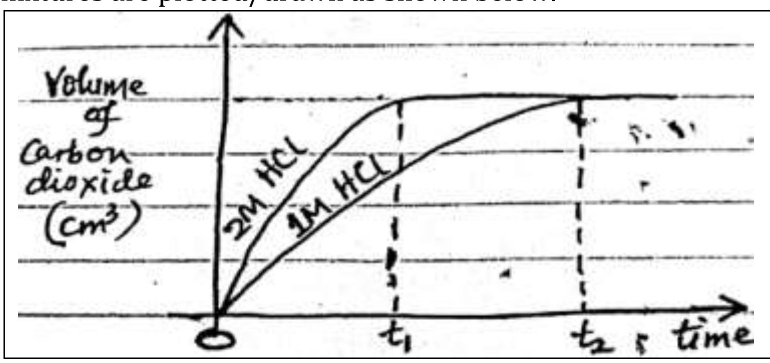


- The carbonate and the acid react as follows:
  $CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
- The rate of reaction can be followed by any of the following methods:

(a) ***Either***; measuring the volume of carbon dioxide formed at regular time intervals separately for each mixture until the reaction is **complete**.
The setup of apparatus for the study is as follows:

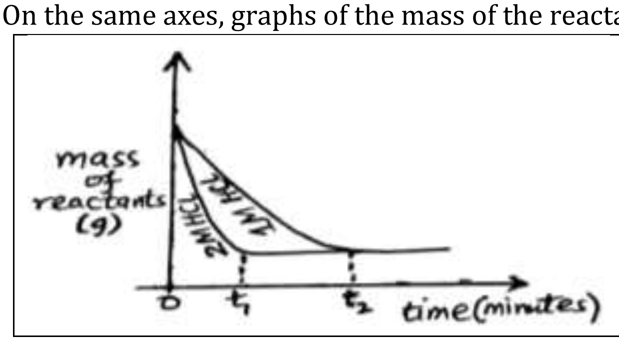


On the **same axes**, graphs of volumes of carbon dioxide collected against time for the mixtures are plotted/drawn as shown below.



When 2 M hydrochloric acid is used, the reaction takes a shorter time to reach completion than when 1 M hydrochloric acid is used. Hence, the higher the concentration, the faster the reaction rate.

(b) ***Or;*** measuring the mass of the reactants (reaction mixture) at regular time intervals separately for each mixture until reaction is **complete**.
On the same axes, graphs of the mass of the reactants against time are plotted.



When 2 M hydrochloric acid is used, the reaction takes a shorter time to reach completion than when 1 M hydrochloric acid is used.
Hence, the higher the concentration, the faster the reaction rate.

❖ In all cases (i.e. in (a) and (b)), the two graphs level off at the **same point** because the **same mass of calcium carbonate** is used.

Trial questions

1. Magnesium reacts with dilute acids according to the following equation.

$$Mg(s) + 2H^{+}(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$$

   (a) State what would be observed if dilute hydrochloric acid was added to magnesium ribbon.

   (b) The rate of the reaction is affected by the concentration of dilute hydrochloric acid.

      (i) State **one** other factor that can affect the rate of the reaction

      (ii) Describe an experiment that can be carried out in the laboratory to show that concentration of dilute hydrochloric acid affects the rate of the reaction.
      (You answer should include suitable diagram of the setup of apparatus that can be used and relevant graphs)

2. Explain each of the following observations

   (a) 100 cm^3 of 2 M hydrochloric acid reacts with 2.0 g of calcium carbonate more vigorously than the same volume of 1 M hydrochloric acid.

   (b) 2 M nitric acid reacts with magnesium more vigorously than a 2 M ethanoic acid dose.

   (c) When a 4 M hydrochloric acid was added to calcium carbonate, the rate the reaction was faster than when a 2 M hydrochloric acid was used.

3. Sodium thiosulphate reacts with dilute acids according to the following equation.

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

   The time taken for the formation of sulphur seen as a yellow precipitate indicates the rate of the reaction.

   (a) Name **one** reagent that would be used to confirm the presence of sulphur dioxide, and state what would be observed if the reagent you have named was treated with sulphur dioxide.

   (b) Define the term **rate of a chemical reaction**.

   (c) The table of results below shows the time taken for sulphur to form when various concentration of sodium thiosulphate were reacted with a fixed volume of 2 M hydrochloric acid.

Concentration of thiosulphate (M)	0.2	0.6	0.8	1.2	1.6
Time for sulphur to form, t(s)	60	20	15	10	7.5
$\frac{1}{t}(s^{-1})$					

      (i) Determine the values for, $\frac{1}{t}$, copy the table and enter your answers in the spaces provided in the table.

      (ii) Plot a graph of $\frac{1}{t}$ against concentration of thiosulphate

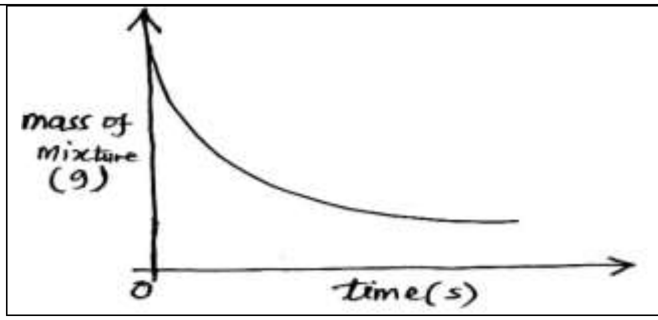
      (iii) State the relationship between a rate of the reaction and $\frac{1}{t}$.

      (iv) Using your graph, deduce how the rate of the reaction varies with the concentration of thiosulphate.

      (v) Determine the slope of the graph.

4. (a) State the factors that can affect the rate of a chemical reaction.

   (b) A mixture of a known mass of zinc and a certain volume of 2 M hydrochloric acid was put in a conical flask and the mass of the mixture was recorded at various intervals. The results of the experiment were represented in the graph below.



(i) On the same axes, draw a graph that would be obtained when the same mass of zinc was reacted with the same volume of 1 M hydrochloric acid.
(ii) Mark the graphs, the times when the two reactions reach completion.
(iii) Compare the time the reaction took to reach completion when 1 M hydrochloric acid was used to that when 2 M hydrochloric acid was used.

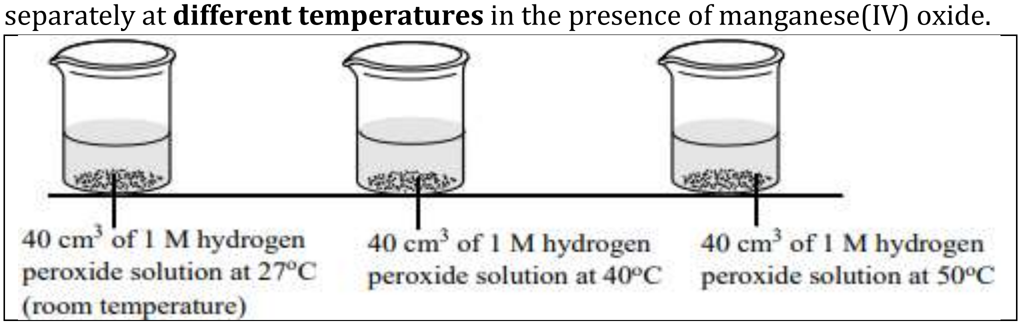
(c) Determine the number of moles of hydrogen ions that is contained in 250 cm^3 of:
(i) a 1 M hydrochloric acid.
(ii) a 2 M hydrochloric acid.
(iii) a 2 M sulphuric acid.

B. Effect of temperature on the reaction

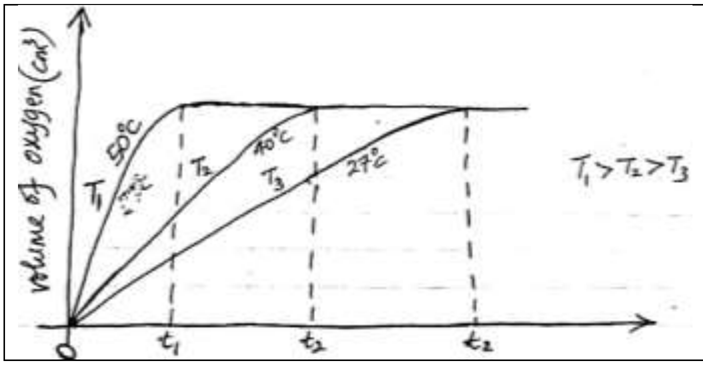
- The **higher** the temperature of the reaction, the **faster** the rate of reaction.
- This is because at higher temperatures, heat is supplied to the reactant particles. Many reactant particles have more kinetic energy and therefore more at increased speeds/velocities (or move faster).
- This results into more energetic and more frequent collisions between the reactant particles**.**
- Hence faster reaction rate.

❖ The effect of temperature on the reaction rate can be illustrated by:

(a) Allowing **equal** volumes of hydrogen peroxide of the **same** concentration to decompose separately at **different temperatures** in the presence of manganese(IV) oxide.



- Hydrogen peroxide decomposes according to the following equation:
  $2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(g)$
- The volume of oxygen produced and collected in a graduated gas syringe is **recorded/noted** at regular time intervals in **each case** (i.e. for each particular temperature) until the reaction is **complete**.
- On the same axes, graphs of volumes of oxygen collected against time at the different temperatures are plotted (or drawn) as shown below.



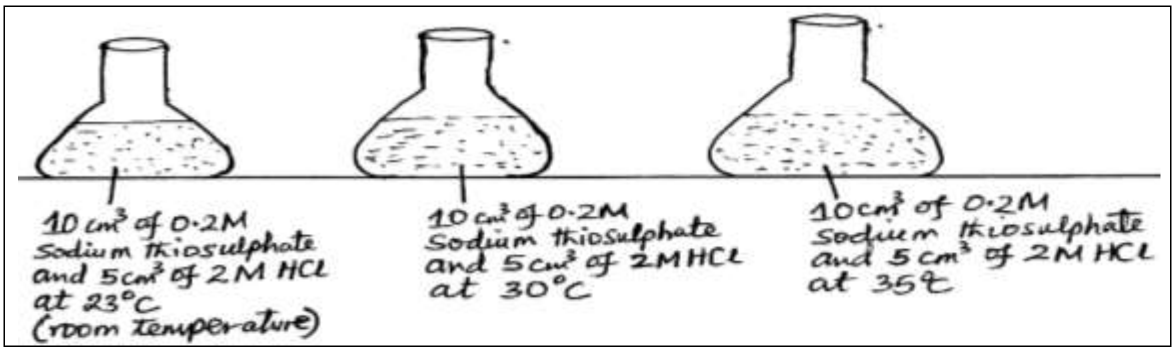
The graphs show that a reaction at a higher temperature takes a shorter period of time to go to completion while that at a lower temperature takes a longer period of time to go to completion. Thus, the rate of the reaction is faster at T_1 than at T_2. Similarly, the rate of the reaction is faster at T_2 than at T_3.

Note: The setup of apparatus that can be used to follow this reaction is similar to the one already drawn on page 3.

(b) Allowing mixtures of the **same** volumes of **standard** sodium thiosulphate solution and **standard** hydrochloric acid to react at **different temperatures** in **separate flasks**.

- For each temperature, the stop clock/stopwatch is started at the **same time** when the **two solutions** are mixed.

Note: The reaction mixture forms a yellow precipitate of sulphur. The intensity of the precipitate (which determines the rate of the reaction) is measured by placing the flasks/beakers on a white sheet of a paper with a cross marked on it.



- Each flask is swirled and placed on a white sheet of paper with a cross marked on it on a table. The cross is then viewed from above the mixture.
- Time taken for the cross to just be covered (i.e. 'disappear') in each case is noted and recorded.
- A graph of temperature against reciprocal of time is plotted. The plot gives a straight line through the origin.
- This shows that the rate of the reaction which is proportional to $\frac{1}{t}$ increases with increase in temperature.

Trial questions

1. Sodium thiosulphate reacts with dilute acids according to the following equation.

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

   (a) State what would be observed if dilute hydrochloric acid was added to sodium thiosulphate solution.

(b) The rate of the reaction is affected by the concentration of sodium thiosulphate.

(i) State **one** factor other than concentration that can affect the rate of the reaction.

(ii) Briefly explain the effect of the factor you have stated in (b)(i) on the rate of the reaction.

(iii) Describe an experiment that can be carried out in the laboratory to show the effect of the factor you have stated in (b)(i) on the rate of the reaction. (*Diagram(s)* ***not*** *required*)

(c) The table below shows the variations in the concentration of sodium thiosulphate with time.

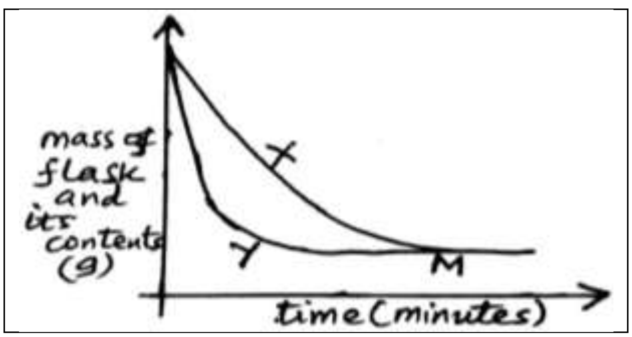
Time (s)	200	100	40	20	10
Concentration of thiosulphate (mol dm^{-3})	0.05	0.09	0.15	0.20	0.25
$\frac{1}{\text{Concentration of sodium thiosulphate}}$ $(mol^{-1}dm^{3})$					

(i) Copy and complete the table by entering values for $\frac{1}{\text{Concentration of sodium thiosulphate}}$ in the spaces provided in the table.

(ii) Plot a graph of $\frac{1}{\text{Concentration of sodium thiosulphate}}$ (vertical axis) against time (horizontal axis).

(iii) State any conclusions that can be drawn from the shape of the graph.

2. Curve **X** in the diagram below shows the results that were obtained during the investigation of the rate of the reaction between iron and dilute hydrochloric acid under normal conditions.
Curve **Y** was obtained when some conditions of the experiment were changed.



(a) (i) List the conditions that were changed to obtain curve **Y**.

(ii) Give a reason for your answer.

(iii) State what point M represents.

(b) Explain why the curves eventually end at the same level.

(c) State **one** other method that can be used to measure the rate reaction between iron and dilute hydrochloric acid.

(d) Suppose during the investigation, the reaction between iron and dilute hydrochloric acid was first carried out 60°C and the reaction was repeated at 40°C.

(i) State which curve shows the reaction taking place at
60°C.. 40°C..

(ii) Give a reason for your answer in (d)(i).

C. Effect of surface area (particle size) of the solid reactant on the reaction

- The **smaller** the particle size of the solid reactant, the **faster** the rate of reaction.
- Small sized particles (powder) provide a large surface area on which a reaction can occur.
- This provides greater chances for many (more) reactant particles to collide.
- Hence frequent and vigorous collisions per unit time which give rise to faster reaction rate.

Alternatively the effect of surface area of solid reactant can be explained as follows:
Powdered reactants have ***many reactant particles*** *than lumps of the* ***same mass*** *of the reactant. Thus, there are greater chances of collision per second between the reactant particles in powder than in the lumps.*

❖ The effect of particle size of the solid reactants on the rates of chemical reactions can be demonstrated as follows:

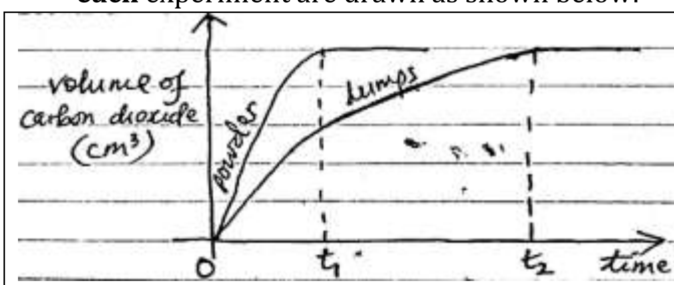
Procedure

I. A given mass of **powdered** calcium carbonate is allowed react with a given volume of standard hydrochloric acid.

II. The volumes of carbon dioxide evolved are collected and noted at regular time intervals until the reaction is **complete**.

III. Procedures **I** and **II** are repeated using the same mass of lumps of calcium carbonate and same volume of the same standard hydrochloric acid at the same temperature.

IV. On the same axes, graphs of volumes of carbon dioxide evolved at regular time intervals in **each** experiment are drawn as shown below:

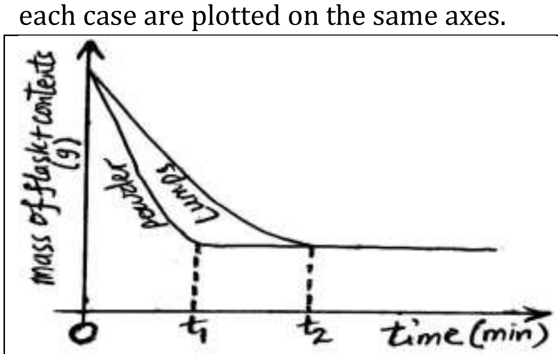


Note: The setup of apparatus that can be used to follow this reaction is similar to the one already drawn on page 3.

From the graphs, it's noted that the reaction involving powdered calcium carbonate and acid comes to completion faster than that involving lumps of calcium carbonate and the same acid. Thus, the rate of the reaction is faster when the surface area is large.

❖ **Alternatively**, the rate of the reaction above can be followed by recording the mass of the flask and reactants (its contents) at regular time intervals until the reaction is complete when both the powdered and lumps of calcium carbonate are separately used.

❖ The graphs of masses of flask and reactants (its contents) recorded at regular time intervals for each case are plotted on the same axes.



From the graphs, it's noted that the reaction involving powdered calcium carbonate and acid comes to completion faster than that involving lumps of calcium carbonate and the same acid. Thus, the rate of the reaction is faster when the surface area is large.

Trial questions

1. (a) Describe an experiment to show how surface area (particles size) can affect the rate of the following reactions.
   (i) Calcium carbonate and 2 M hydrochloric acid.
   (ii) Zinc and dilute sulphuric acid
   In each case, your answer must include:
   - A well labelled diagram of the setup of apparatus used.
   - Sketch of expected graphs.
   - Mention how the graphs can be used to reach conclusions.
   - Ionic equation for the reaction that takes place.

   (b) State **two** factors other than the one mentioned above that can affect the rate of the reaction in (a)(i).
2. When a certain volume of 0.1 M hydrochloric acid was reacted at room temperature with excess iron filings, 120 cm^3 of a gas were produced.
   (a) (i) Identify the gas that was produced during the reaction.
   (ii) Briefly describe how the gas can be identified in the laboratory.
   (b) (i) Draw a well labelled diagram to show the rate of the reaction was determined.
   (ii) Write the equation for the reaction that took place
   (c) Calculate the:
   (i) volume of 0.1 M hydrochloric acid required to produce 120 cm^3 of the gas.
   (ii) mass of iron filings that reacted.
   (d) Draw a sketch graph to show how the volume of the gas produced varies with time.
   (e) (i) State how the rate of reaction would change if the reaction was carried out at a temperature above room temperature.
   (ii) Explain your answer in (e)(i).
3. Magnesium ribbon reacts with 2 M hydrochloric acid according to the following equation.

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

   (a) State **two** ways through which the reaction can be made to proceed faster.
   (b) The table below shows the volume of hydrogen collected at various time intervals when magnesium was reacted with 2 M hydrochloric acid.

Time(s)	0	1	2	3	4	5	6	7
Volume of hydrogen collected (cm^3)	0	25	45	60	70	75	77	77

   (i) Plot graph of volume of hydrogen collected against time.
   (ii) From your graph; determine;
   I. the rate of reaction at 3 seconds.
   II. the volume of hydrogen evolved at 3.5 seconds.
   (c) State how the rate of reaction at 3 seconds would be affected if a 1 M hydrochloric acid was used.
   (d) In another experiment, excess zinc granules were added to dilute hydrochloric acid and the volume of hydrogen evolved with time noted as showed in the table below.

Time(s)	0	10	20	30	40	60	80	90
Volume of hydrogen collected (cm^3)	0	20	35	46	56	72	79	79

   (i) Plot a graph of volume of hydrogen evolved against time.
   (ii) From your graph, determine;
   I. the time taken to collect 60 cm^3 of hydrogen.
   II. the rates of the reaction at 20 s and 60 s.

(iii) How does the rate of the reaction at 20 s compare with that of 60 s. Explain your answer? (**Or** Comment on the rate of the reaction after 20 s and 60 s. Explain your answer.) (iv) Suggest a reason why the volume of hydrogen evolved remains constant after 80s. (v) Explain what would happen to the reaction if zinc granules were replaced with zinc powder. 4. (a) Write equation for the reaction between calcium carbonate and dilute nitric acid. (b) State and explain the effect of the following factors on the rate of the reaction in (a). (i) Temperature. (ii) Large surface area of calcium carbonate.

D. Effect of a catalyst on the rate of reaction

❖ **Recall;** A catalyst is a substance which alters the rate of a chemical reaction but remains unchanged in chemical nature and in mass (or amount) at the end of the reaction.

❖ Examples of catalysts for some common reactions include:

Catalyst	**Reactions they catalyze**	**Reaction equation involved**
Manganese(IV) oxide	Decomposition of hydrogen peroxide (*Lab preparation of oxygen*)	$2H_2O_2(aq) \longrightarrow 2H_2O(l) + O_2(s)$
	Decomposition of potassium chlorate(V)	$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(s)$
Copper(II) sulphate	Reaction between zinc and dilute sulphuric acid (*Lab preparation of hydrogen*)	$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$
Vanadium(V) oxide (or sometimes platinum)	Synthesis of sulphur trioxide in the Contact process during the manufacture of sulphuric acid.	$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
Finely divided iron	Synthesis of ammonia in the Haber process.	$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
Platinum or platinum-rhodium gauze	oxidation of ammonia during the manufacture of nitric acid	$2NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$

Note: Platinum /iron(III) chloride/copper(II) oxide can also be used to catalyze the decomposition of hydrogen peroxide.

❖ In the presence of a catalyst, the rate of reaction is faster than when the catalyst is absent. This is because a catalyst lowers the activation energy by providing an alternative reaction path. This results into more frequent collisions between the reactant particles and hence faster reaction rate.

❖ Most catalysts used in industry are solids and therefore they must be finely divided (e.g. as powders, pellets, fibres or gauze). This is because the reactions only place on their surfaces and thus they are finely divided to provide a larger surface area for maximum yield of the required product.

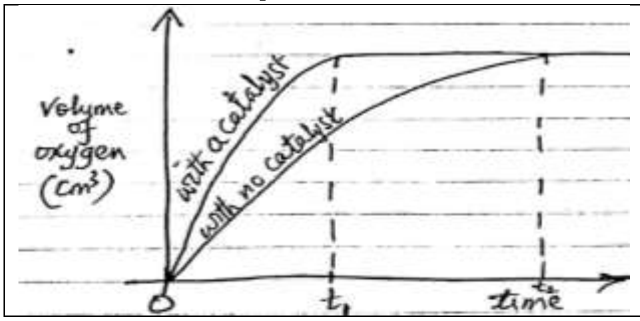
❖ The effect of a catalyst on the rates of chemical reactions can be demonstrated as follows:

I. A given volume of standard hydrogen peroxide solution is allowed to decompose in the presence of manganese(IV) oxide catalyst at room temperature.

II. The volume of oxygen produced and collected in a gas syringe is noted and recorded at regular time intervals until the reaction is complete.

III. Procedures I and II are repeated with the same volume of the same standard hydrogen peroxide at the same temperature but with no catalyst used.

IV. The graphs of volumes of oxygen recorded against time are plotted/drawn on the **same axes** for the two experiments.



From the graphs, it is noted that the decomposition involving a catalyst comes to completion faster than that without a catalyst.

Note: The setup of apparatus for use in this study is similar to the one already drawn on page 3.

E. Effect of light on the rate of reaction

❖ Light provides energy necessary for frequent collisions between the reactant particles for some reactions to proceed.

❖ **For example;**

(i) Production of starch by green plants from carbon dioxide and water. This process is a called **photosynthesis**.

(ii) Reaction between chlorine and methane (or hydrogen) occurs explosively in sunlight (at room temperature).

(iii) Decomposition of hydrogen peroxide.

Trial questions

1. Hydrogen peroxide was exposed sun light.
   (a) State:
      (i) what was observed.
      (ii) the role of sunlight.
   (b) Write equation for the reaction that took place.
   (c) Calculate the volume of the gaseous product that would be collected at room temperature when 20.0 g of hydrogen peroxide was allowed to decompose completely.
   (d) (i) State what would be observed if manganese(IV) oxide was added to the hydrogen peroxide solution.
      (ii) Give a reason for your answer in (d)(i).
   (e) State the condition(s) and write equation for the reaction between magnesium and one of the products in (b).
2. Oxygen can be formed from hydrogen peroxide in the presence of manganese(IV) oxide according to the following equation.
   (a) State:

(i) the role of manganese(IV) oxide.
(ii) **two** ways by which the formation of oxygen can be made to proceed at a faster rate than the normal rate in the absence of manganese(IV) oxide.

(b) In an experiment, a certain volume of hydrogen peroxide was used to prepare oxygen at room temperature. With the aid of a suitable diagram, describe how the following can be determined:

(c) In another experiment, one half of the volume of hydrogen peroxide used in (b) was diluted with an equal volume of water. On the same axes, draw graphs to show the variation of volume of oxygen with time for experiments in (b) and (c).

(d) Oxygen produced from 200 cm^3 of 0.5 M hydrogen peroxide solution was reacted completely with strongly heated magnesium. Calculate the mass of magnesium that reacted.

3. (a) Hydrogen peroxide decomposes quite easily at room temperature.
(i) State what would be observed if a glass beaker containing hydrogen peroxide was allowed to stand at room temperature.
(ii) Write equation for the decomposition of hydrogen peroxide.
(iii) State **two** ways by which the decomposition of hydrogen peroxide can be made to proceed faster.

(c) On the same axes, sketch graphs of concentration of hydrogen peroxide versus time for the decomposition of the hydrogen peroxide
(i) at room temperature.
(ii) under **one** of the conditions you have stated in (a)(iii).

4. (a) The table below gives results from a reaction mixture containing zinc granules, dilute sulphuric acid and copper(II) sulphate solution.

Time (minutes)	0	5	10	15	20	25	30
Volume of gas (cm^3)	0	10	20	25.5	29.5	32	32

(i) Write the name and chemical formula of the gas formed from the reaction mixture.
(ii) Briefly describe how the named gas in (a)(i) can be identified in a laboratory.
(iii) State the role of copper(II) sulphate solution in the mixture.
(iv) Write an ionic equation for the reaction that took place in the mixture.
(v) Explain what would happen to the reaction mixture if zinc granules were replaced with zinc powder.
(vi) Plot a graph of volume of gas formed (vertical axis) against time (horizontal axis).
(vii) Describe how you would determine the rate of the reaction at 12 minutes.
(viii) Compare the rate at 12 minutes with that at 20 minutes. Give a reason for your answer.
(ix) What happens to the shape of the graph after 20 minutes? Explain your answer.

(b) A certain mass of zinc powder was reacted with hydrochloric acid at room temperature.
(i) Write equation for the reaction that took place.
(ii) Sketch a graph to show how the volume of the gaseous product varies with time.
(iii) What would be the effect of:
(I) adding copper(II) sulphate solution to the reaction mixture at room temperature?
(II) using the same mass of zinc granules instead of the zinc powder?

(c) Give a reason for your answer in (b)(iii)(II).

5. State and explain the effect of each of the following conditions on the rate of a chemical reaction.
(a) Particle size.
(b) Concentration of reactants.
(c) Temperature.