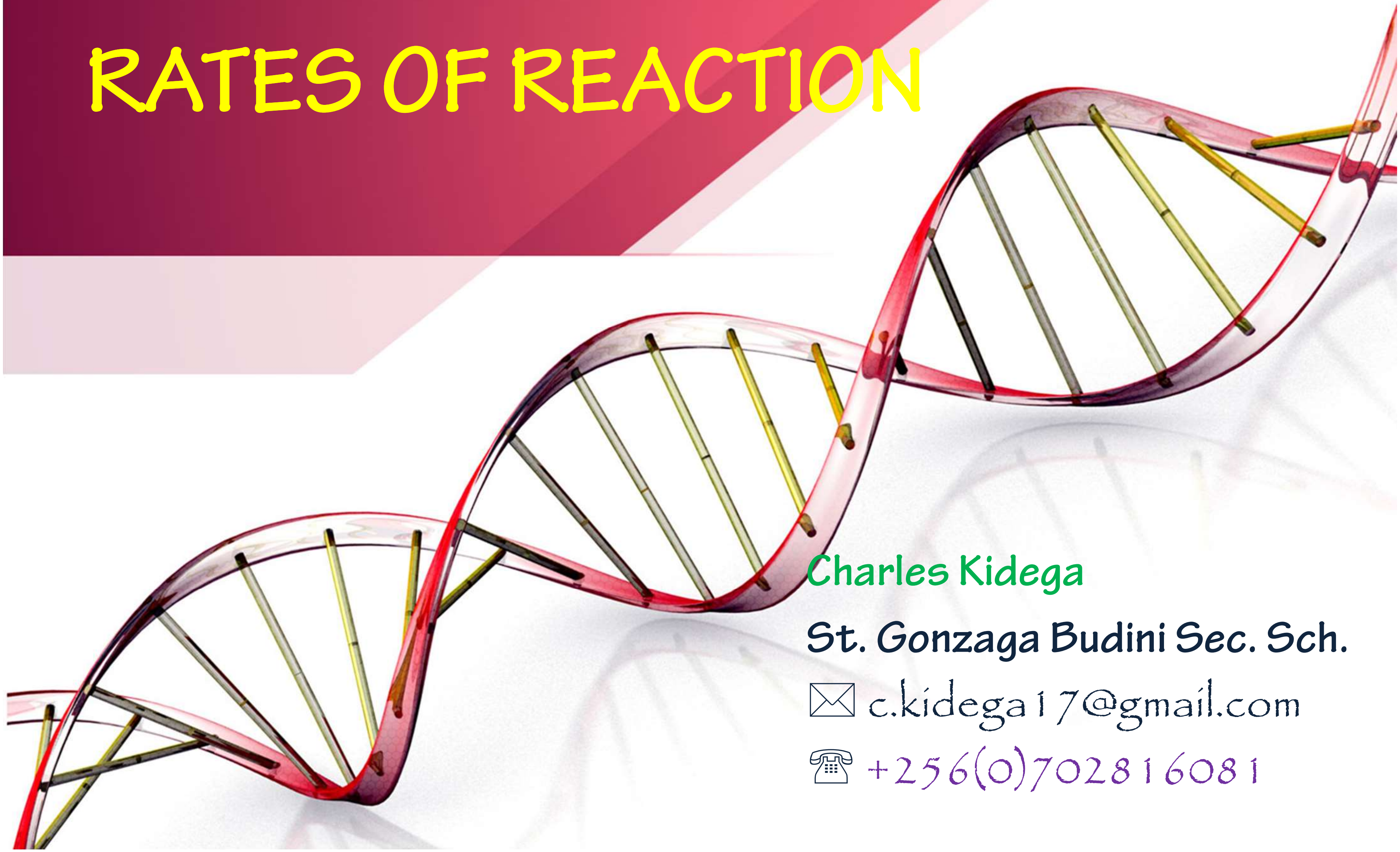
RATES OF REACTION



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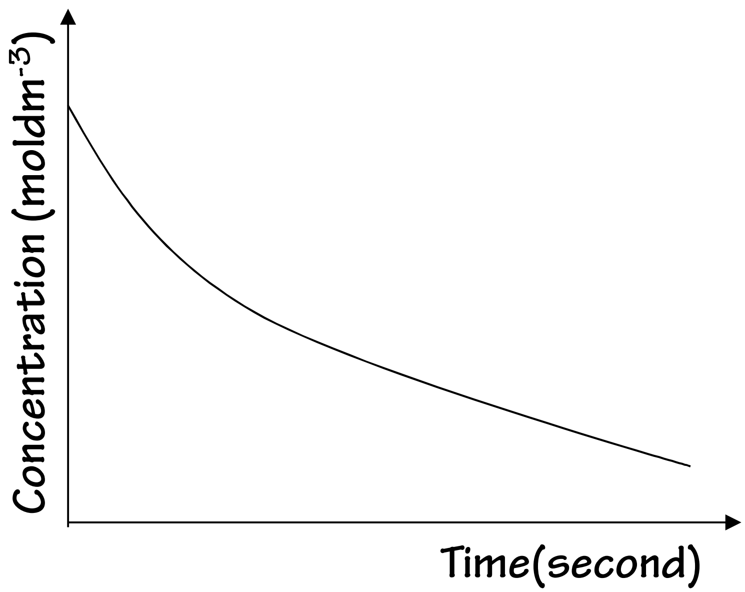
Introduction to rate of reaction

Rate of reaction is the change in the mass/ concentration/ volume of a given reactant/ product during the course of reaction per unit time.

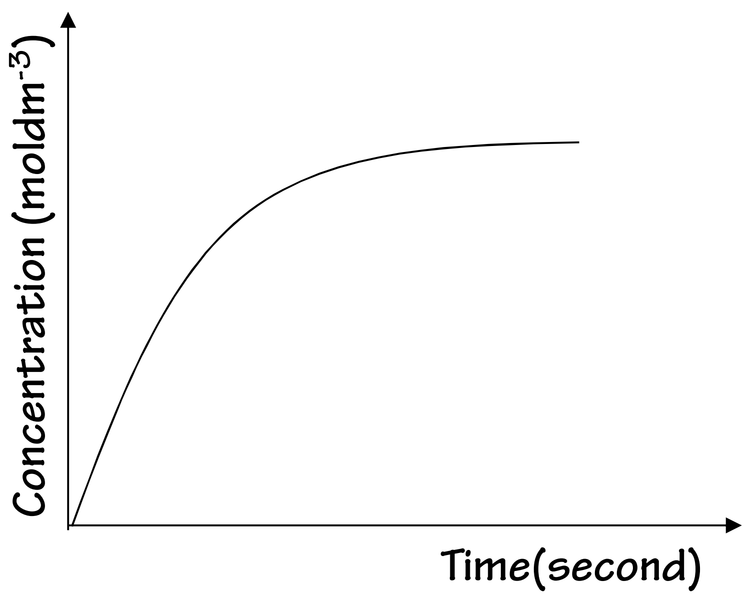
$$\text{Rate of reaction} = \frac{\text{change in amount}}{\text{change in time}}$$

The rate of reaction with respect to reactant is NEGATIVE because its amount decreases with time.

A graph of concentration of reactant against time



A graph of concentration of product against time



Rate of reaction is inversely proportional to the time of reaction.

$$\text{Rate} \propto \frac{1}{\text{time}}$$

Thus a fast reaction takes a short time and a slow reaction takes a long time to reach completion.

Determination of rate of reaction

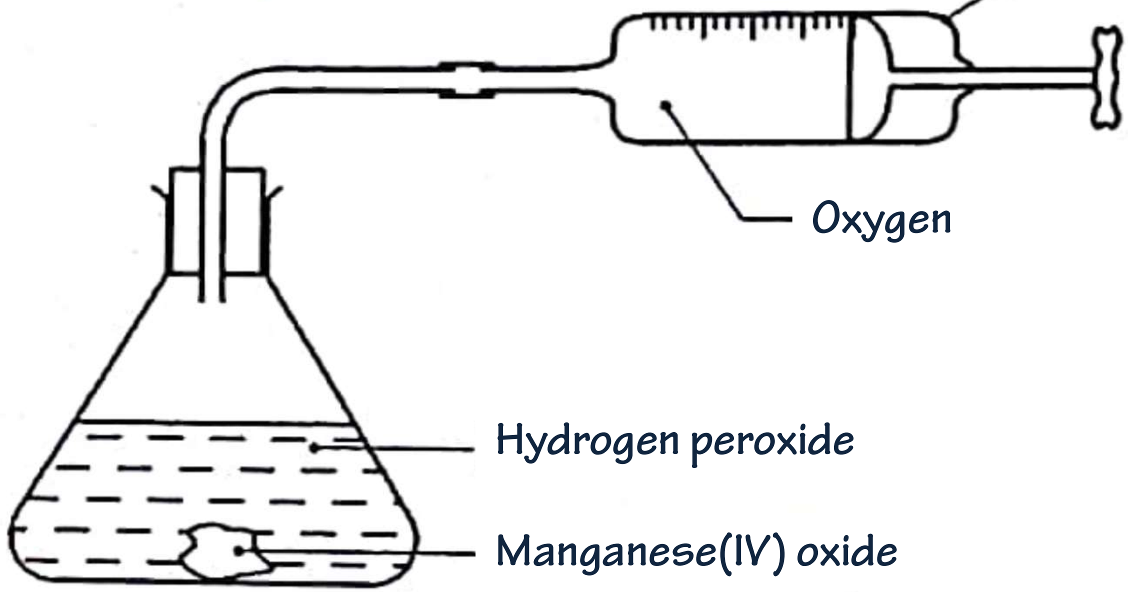
The rate of reaction can be determined by:

(a) Measuring the volume of gaseous product at different time intervals, for example:

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

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

The volume of gas can be measured using a gas syringe with a frictionless piston.



Manganese(IV) oxide is taken in the flask in the set up of apparatus shown in the diagram above.

The stopper is opened and hydrogen peroxide is added onto the manganese(IV) oxide.

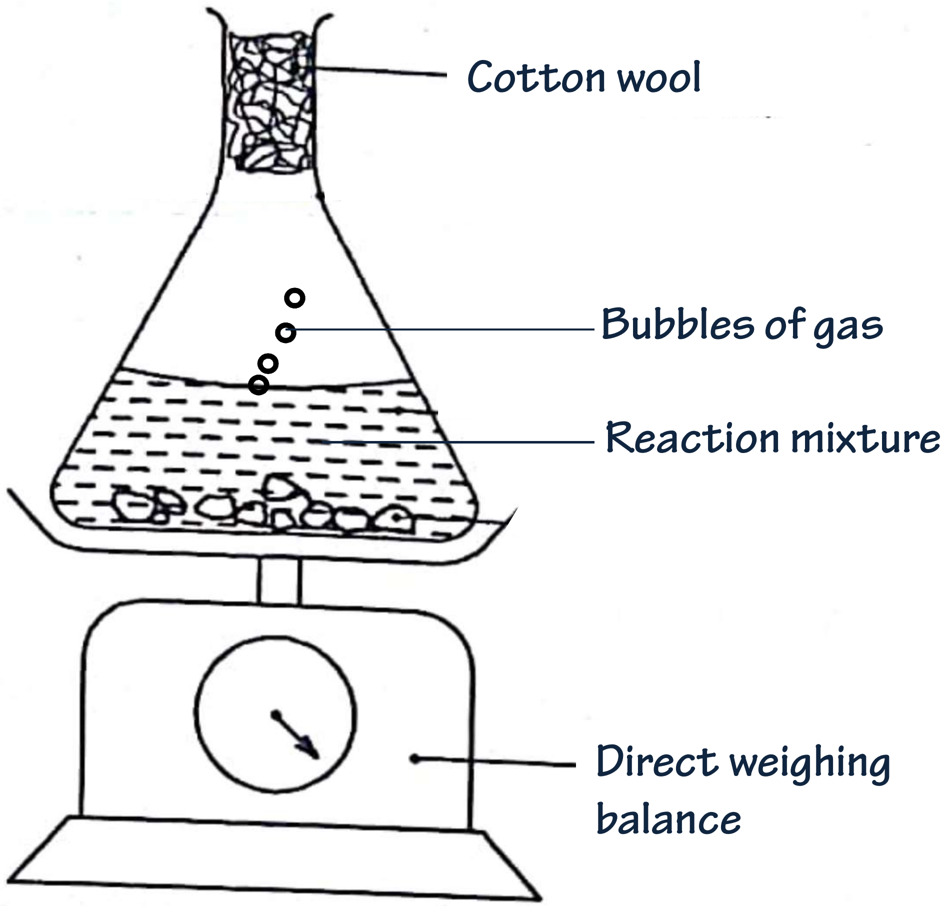
The stopper is quickly fitted back. Volume of oxygen evolved is measured from the gas syringe.

(b) Measuring mass of the reaction mixture at different time intervals.

This method applies to reactions in which a gas is given off and it is allowed to escape. For example,

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

Set up of apparatus



(c) Measuring the time taken for the reaction to reach a specific stage.

(i) Time taken for effervescence to stop e.g.

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

(ii) Time taken to precipitate sulphur.

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

(iii) Time taken for the brown color to appear in solution.

$$H_2O_2(aq) + 2I^-(aq) + 2H^+(aq) \longrightarrow I_2(aq) + 2H_2O(l)$$

Check Point-1

Question 1

1999/P2/11(a)(i) What is the rate of reaction? OR

2003/P2/11(a) What is meant by rate of a chemical reaction? OR

2007/P2/13(a) Define the term rate of reaction.

Question 2 (2002/P2/12)

When a certain volume of 0.1M hydrochloric acid was reacted at room temperature with excess of iron fillings, $120cm^3$ of a gas were produced.

(a) Draw a labelled diagram to show how the rate of reaction was determined.

(b) Write equation for the reaction that took place.

(c) Draw a sketch graph of the volume of the gas against time.

Question 3 (1987/P2/7(a))

A given mass of magnesium strips was reacted with dilute hydrochloric acid at room temperature. The volume of the gas produced was measured at various intervals.

(i) Write equation for the reaction.

(ii) Sketch a graph to show variations of the volume of the gas produced with time.

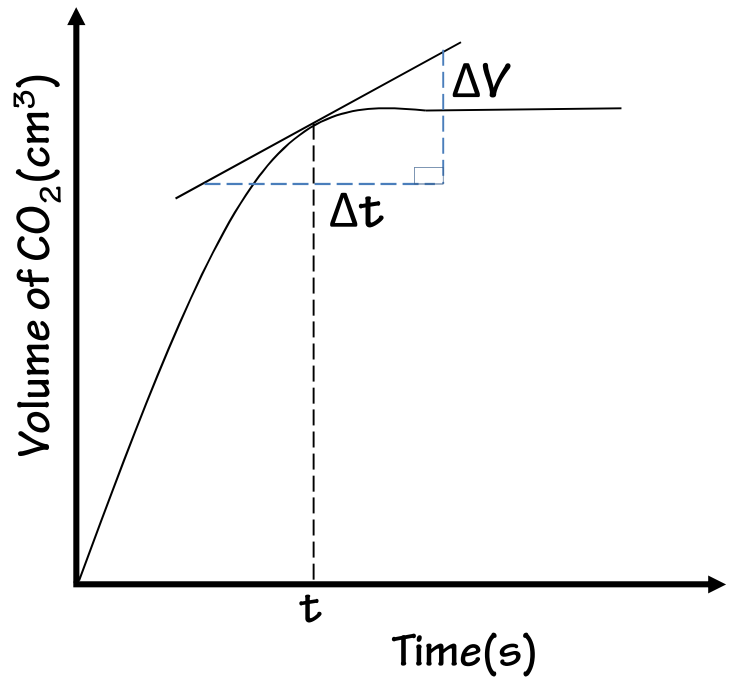
Rate curves

A rate curve is a graph obtained when volume of gas/ concentration of the reactant/ mass of the reaction mixture / loss in mass/ temperature of reaction is plotted against time/ reciprocal of time or vice versa.

To determine the rate of reaction at any time, t, on the graph:

- Draw a tangent line at the time given.
- Complete it to obtain a right-angled triangle.
- Then calculate the gradient/slope of the tangent and this is equal to the rate of reaction at time, t.

Calculating rate of reaction at time, t, from the graph



$$Rate = \frac{\Delta V}{\Delta t} \quad cm^3s^{-1}$$

- To find the volume of a gas/ mass of reactant/ concentration of reactant at a given time from the graph, draw a perpendicular line from the time axis at the given time, to meet the curve. Read off the value where the perpendicular line meets the rate curve.

Note

- Show evidence of value obtained from the graph, use dotted lines or triangle in case of slope.
- To explain the shape of the graph: State the relationship and give reason(s) for the shape.

- Where necessary, tabulate the rate of reaction, $\frac{1}{t}$, giving the values in scientific notation, for example, 0.0068 is recorded as 6.8×10^{-3}.

Checkpoint-2

Question 1

The table below shows the variation in the concentration of hydrogen peroxide with time when a sample of hydrogen peroxide was mixed with iron(III) chloride at room temperature.

Concentration of $S_2O_3^{2-}$ (M)	0.2	0.6	0.8	1.2	1.6
Time for sulphur to form (s)	60	20	15	10	7.5
$\frac{1}{t}(s^{-1})$					

(a) Copy and complete the table above by computing and filling in the values of $\frac{1}{t}$

(b) Plot a graph of $\frac{1}{t}$ against concentration of hydrogen peroxide.

(c) Using a graph, deduce how the rate of the reaction varies with the concentration of hydrogen peroxide.

(d) Determine the slope of the graph.

Question 2

The table below shows the variation in volume of hydrogen evolved with time when dilute hydrochloric acid was added to excess zinc.

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

(a) Plot a graph of volume of hydrogen evolved against time.

(b) Using the graph, determine the time taken to collect 60 cm^3 of hydrogen gas.

(c) Draw tangents on your graph at points when the time is 20 and 60 seconds and determine the gradient of each tangent.

(d) Compare the rate of reaction at 20 seconds and 60 seconds.

Question 3

Dilute hydrochloric acid reacts with zinc according to the equation.

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

25cm^3 of dilute hydrochloric acid of different concentration were separately placed in six beakers. The same mass of zinc powder was added to each beaker and the time taken for effervescence to stop was recorded as shown in the table below.

Concentration of HCl ($moldm^{-3}$)	0.25	0.50	0.75	1.00	1.25	1.50
Time taken(s)	32.0	22.0	14.5	9.5	6.0	3.0

(a) Plot a graph of concentration against time.

(b) Use the graph to determine the

(i) concentration of hydrochloric acid

for which the reaction stops after 10s.

(ii) rate of reaction with 0.50M and 1.25M hydrochloric acid.

(c) Compare the reaction rate in (b)(ii). Explain your answer.

Factors affecting rate of reaction

The factors that affect the rate of chemical reactions are concentration, temperature, pressure, catalyst, surface area and light.

The effect of each factor on the rate of reaction is explained by the collision theory which states that chemical reactions take place when reactant particles collide.

Thus rate of reaction is directly proportional to the number of collisions per unit time (frequency of collisions).

- Effect of temperature

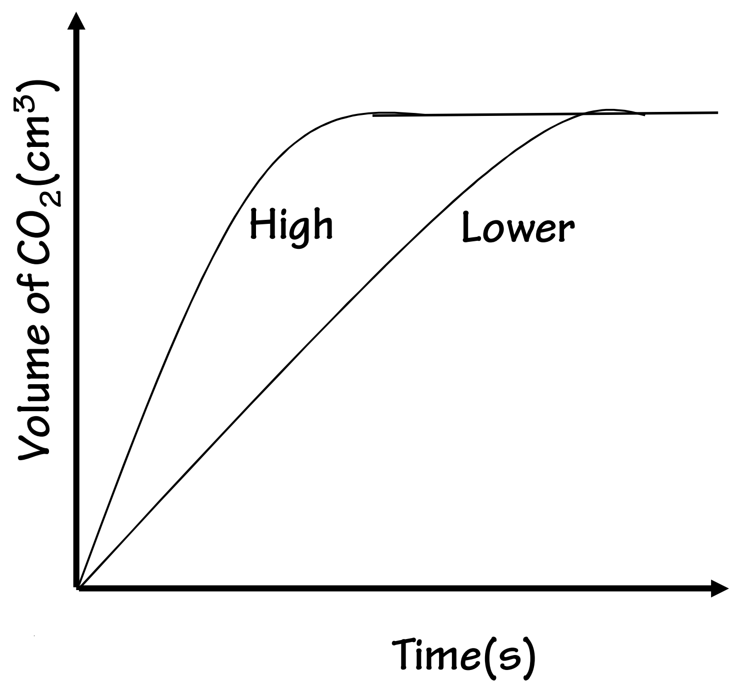
The rate of reaction increases with increase in temperature OR

The higher the temperature, the faster the rate of a chemical reaction

At high temperature, more particles acquire more kinetic energy.

Particles with high kinetic energy move at increased velocities resulting into more particles colliding with one another more frequently and vigorously. Hence the reaction proceeds at a faster rate.

A graph showing volume of gas evolved against time at different temperatures.



Note

The curves eventually end at the same level.

This is because the same mass of reactant is used and all have been used up.

Expt. To show the effect of temperature on the rate of reaction between sodium thiosulphate and dilute acid

Standard solutions of thiosulphate and hydrochloric acid are made.

Equal volumes of thiosulphate solutions are taken in five different conical flasks.

Temperature of the solution in one of the conical flasks is recorded and the conical flask placed over a cross marked on a piece of paper on a table.

A known volume of hydrochloric acid is added and at the same time a stop is started.

Time taken for the cross to just be covered is noted.

The procedure is repeated with the second conical flask heated to a temperature slightly above room temperature. This is done with the 3rd,4th and 5th conical flasks with temperature in each case higher than the preceding one.

Graph of temperature against reciprocal of time is plotted which gives a straight line. This shows that rate of reaction increases with increase in temperature.

- Effect of concentration of reactants

The rate of reaction increases with increase in concentration of reactants OR

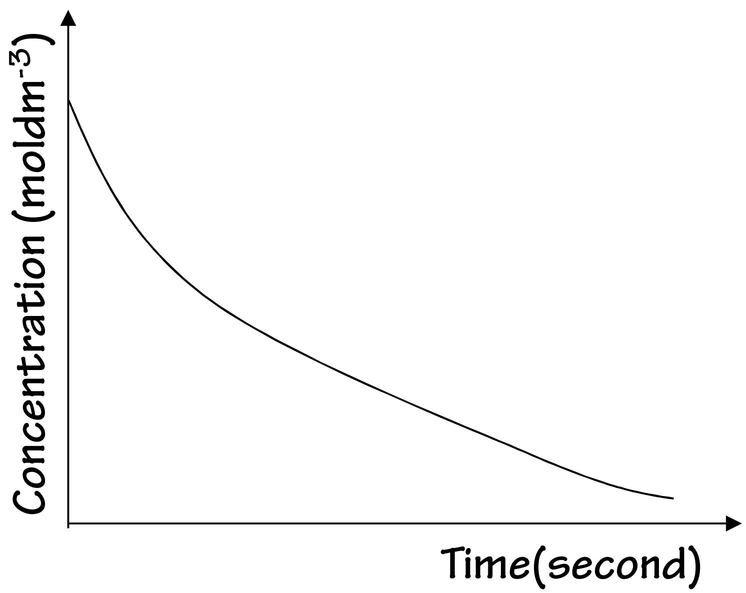
The higher the concentration of reactants, the faster the rate of reaction.

At higher concentration, more particles are brought together and are close to each other. Thus the chances of collisions are higher and more frequent.

For example: when equal masses of excess magnesium ribbon were separately reacted with sulphuric acid of different concentration, different volume of hydrogen was evolved.

Graphs showing effect of concentration on rates of reaction.

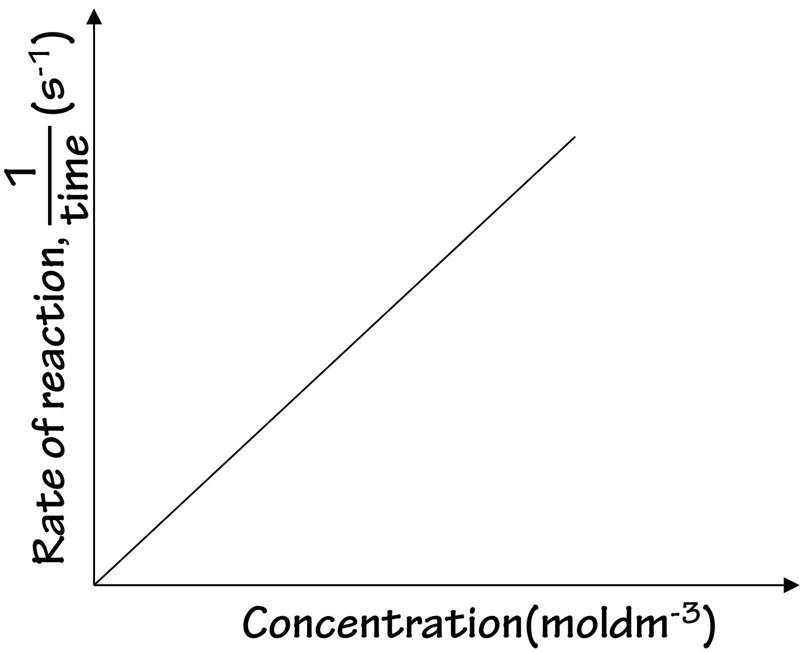
(i) Concentration against time



Initially the concentration of reactants is high leading to a faster rate of reaction.

As the reactants get used up, their concentration decreases and the rate of reaction slows down since there are fewer particles colliding until it finally stops and becomes constant.

(ii) Rate against concentration



The graph is a straight line passing through the origin.

It shows that the rate of reaction is directly proportional to the concentration of reactants.

- **Effect of surface area/particle size**

The smaller the particle size the faster the rate of reaction.

Small sized particles e.g. powdered, provide large surface area on which reaction can occur. This provides greater chances for more particles to participate in the reaction.

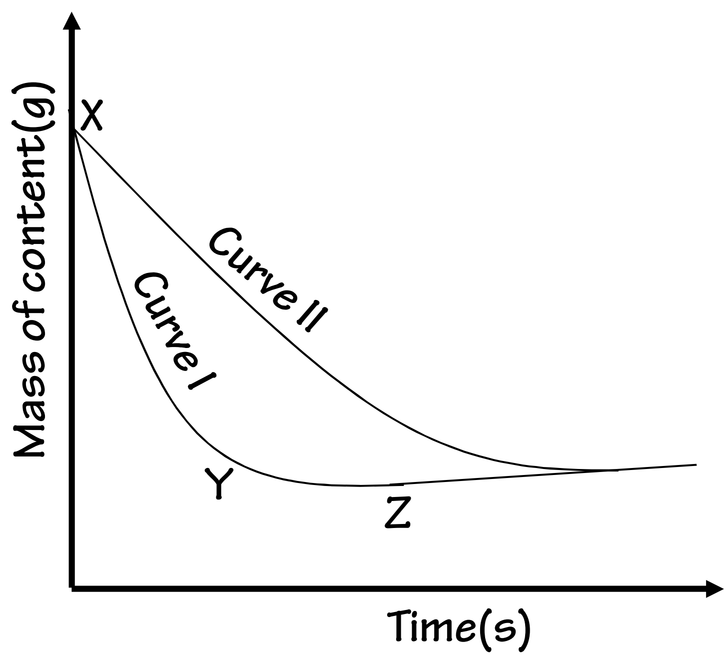
Note

- Surface area affects the rate of reaction involving solid reactant(s) only.
- For example, reaction between zinc and dilute hydrochloric acid shows that the reaction is faster with zinc powder than with zinc granules.

- A reaction between marble chips (lumps of calcium carbonate) and dilute hydrochloric acid is slow compared to that with powdered calcium carbonate.

- Thus the rate of reaction is measured as the rate of change in mass of the apparatus and its content.

A graph of change in mass against time



Curve I: Powdered calcium carbonate

Curve II: Marble chips

In curve I, there is a rapid decrease in the mass from X to Y due to large surface area provided by the powdered calcium carbonate hence the reaction rate is very fast.

From Y to Z, the mass remains constant because the reaction has reached completion.

In curve II, the mass decreases slowly and the rate of the reaction is slower hence less carbon dioxide is produced and it takes a longer time for the reaction to reach completion.

This is because of the small surface area of the marble chips which are in large lumps.

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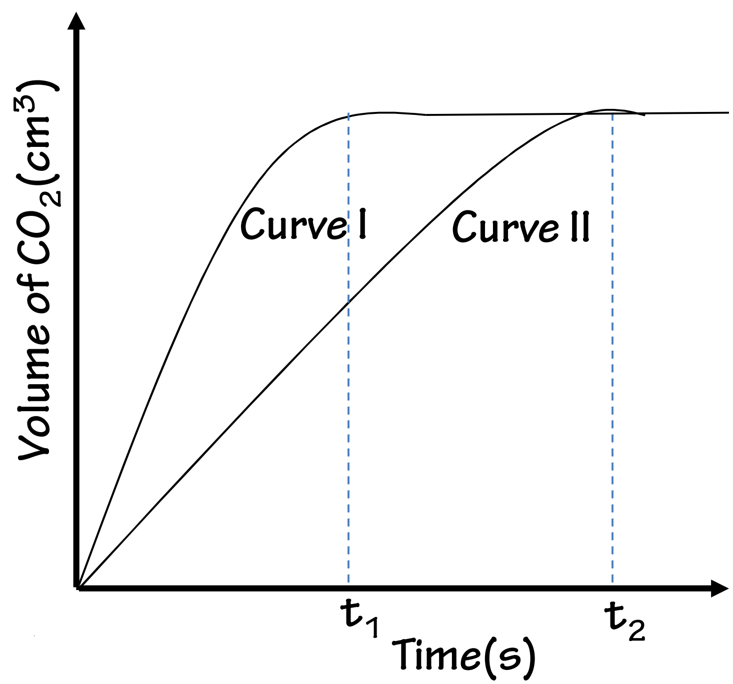
A graph of volume of carbon dioxide evolved against time



Curve I: Powdered calcium carbonate

Curve II: Marble chips

In curve I, there is a rapid increase in the volume of carbon dioxide evolved due to the large surface area provided by the powdered calcium carbonate, hence the reaction rate is very fast until it reaches its completion and becomes constant.

In curve II, there is a gradual increase in the volume of carbon dioxide produced and the rate of reaction slower and it takes a lot of time for the reaction to reach its completion.

This is because of the smaller surface area of the marble chips which are in large lumps.

Expt. To show how surface area can affect the rate of the reaction between calcium carbonate and hydrochloric acid

A given volume of hydrochloric acid of known concentration is put in a conical flask and the flask is loosely plugged with a piece of cotton wool.

The flask with its contents is placed on a direct weighing balance and the mass recorded.

The cotton wool is slightly displaced and a known mass of lumps of calcium carbonate is added to the acid and simultaneously a stop clock is started.

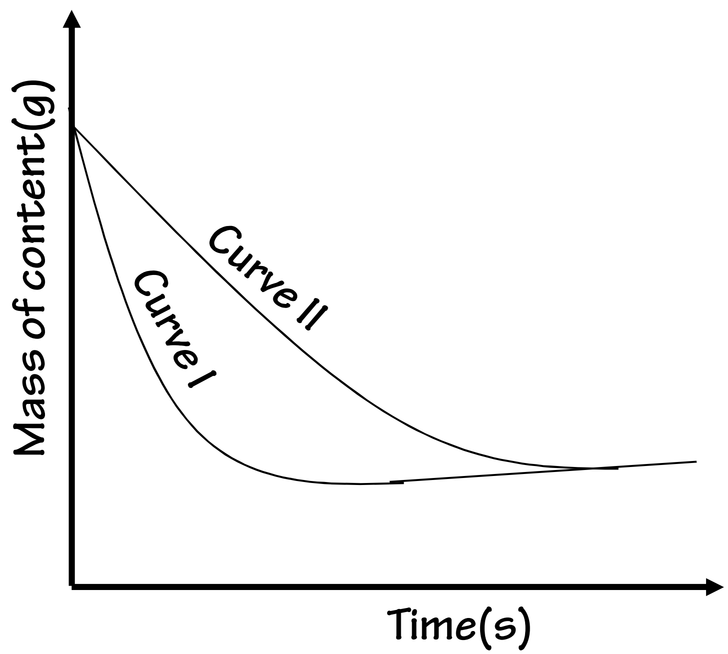
The mass of the flask and its content is recorded after fixed intervals of time until there is no further change.

The experiment is repeated using fresh acid of the same volume and same mass of calcium carbonate powder.

The change in mass is due to the carbon dioxide lost during the reaction.

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

On the same axes, a graph of mass against time for the reaction between the acid and lumps of calcium carbonate(curve II), the acid and calcium carbonate powder(curve I) is plotted.



The graph shows that the big size of lumps of calcium carbonate offers a small surface area for the reaction with the acid compared to the powdered calcium carbonate with small sized particles that offer a large surface area and the latter reaction reaches completion in a shorter time.

Therefore, the larger the surface area, the faster is the rate of reaction.

- Effect of catalyst

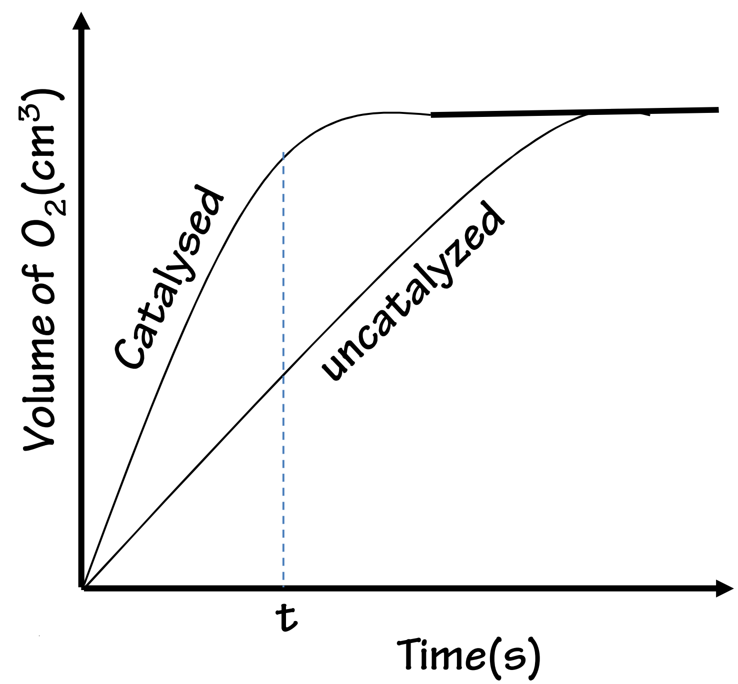
The rate of reaction increases with addition of a catalyst.

This is because in the presence of a catalyst, the number of collisions per unit time increases and the activation energy is lowered, hence increased rate of reaction.

Note

- A catalyst is a substance that alters the rate of chemical reaction but it remains chemically unchanged.
- Powdered catalyst are more effective because the powder presents a large surface area compared to lumps.

A graph showing catalysed and uncatalysed reaction



At any time, t, the volume of oxygen produced is higher in catalysed reaction than in uncatalysed reaction.

- Effect of pressure

A change in pressure affects reactions which occur in gaseous state.

When pressure of a gas mixture is increased, the gases are compressed and this brings the gas particles closer together thus increasing the frequency of collision between the reacting particles, therefore rate of reaction increases.

Examples of reactions affected by pressure are:

$$N_2(g) + H_2(g) \rightleftharpoons 2NH_3(g)$$

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

- Effect of light

Some reactions are photosensitive. For example silver chloride darkens when exposed to light due to its decomposition to silver metal and chlorine gas.

$$2AgCl(s) \longrightarrow 2Ag(s) + Cl_2(g)$$

Checkpoint-3

Question 1 (2013/P2/14)

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

(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 not required)

Question 2 (2010/P2/14)

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.

Question 3 (2005/P2/13)

(a) Describe an experiment to show how surface area can affect the rate of the reaction between calcium carbonate and 2M hydrochloric acid. Your answer must include:

a labelled diagram of the apparatus.

sketch of expected graphs.

mention of how the graphs can be used to make conclusions.

(b) Briefly explain why, when a 4M hydrochloric acid was used instead of 2M acid, the rate of the reaction was faster. Explain the observation.

(c) State one factor other than those mentioned above that can affect the rate of the reaction between hydrochloric acid and calcium chloride.

UNEB Questions

2003/P2/11

2007/P2/13

2009/P2/13

2013/P2/14

2015/P2/9

2018/P2/11

2019/P2/9

END