REACTION RATES
AND
REVERSIBLE REACTIONS

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Organizer

Objectives

By the end of the topic, the learner should be able to:

(a) Define rate of reaction and explain the term activation energy.

(b) Describe some methods used to measure the rate of reaction.

(c) Explain the effect of different factors on reaction rates.

(d) Illustrate reaction rates graphically and interpret experimental data.

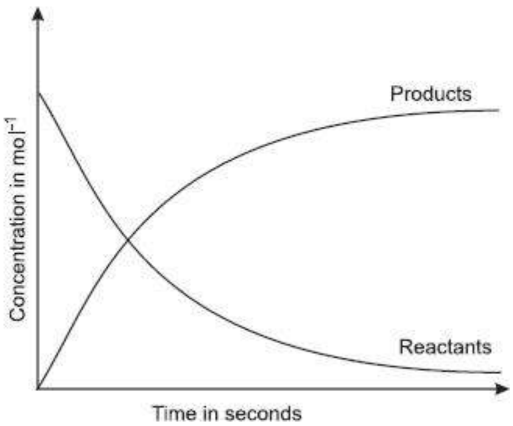
(e) State examples of simple reversible reactions.

(f) Explain chemical equilibrium and the effect of different factors on the position of equilibrium.

REACTION RATES AND REVERSIBLE REACTIONS

Reaction Rates

Reaction rate *is a measure of how much of the reactants are consumed or how much of the products are*



formed per unit time.

Relationship between the concentration of reactants and products during a reaction

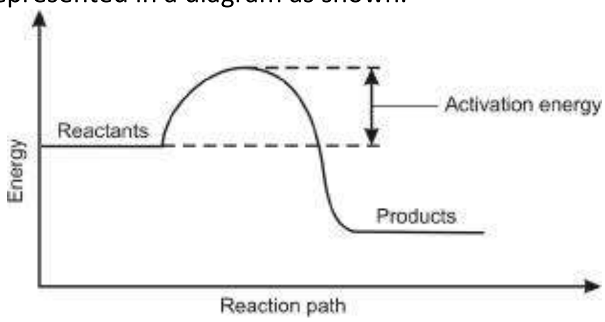
Collision Theory and Activation Energy

In order for a reaction to occur, reacting particles must collide.

Not all collisions results into products. For effective collisions, the reacting particles must have a certain minimum energy.

The minimum energy required by the reacting particles to cause a successful collision to form products is referred to as **Activation Energy**.

This information can be represented in a diagram as shown.



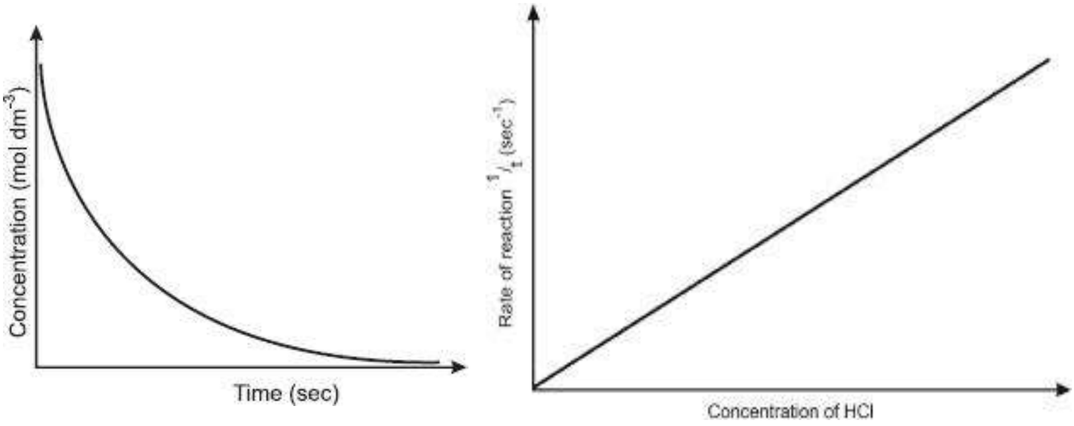
Energy changes during a reaction

Factors Affecting the Rate of a Reaction.

The rate of a reaction depends on concentration, pressure (for gaseous reactants), temperature, surface area(for solid reactants) and catalyst.

1. Concentration.

When the concentration of the particles of the reactants is high, the number of collisions is also high. This increases the number of successful collisions and hence, increases the rate of a reaction.



Graph of (a) concentration versus time and (b) rate of reaction vs concentration

In a reaction between an acid, say HCl, and a metal eg Magnesium ribbon, it follows that the higher the concentration of the acid, the faster the reaction.

A graph of $\frac{1}{t}$ (s^{-1}) against concentration of the hydrochloric acid shows that the rate increases with increase in concentration

In the experiment, the rate is measured by the time taken for the piece of magnesium ribbon to disappear.

2. Time.

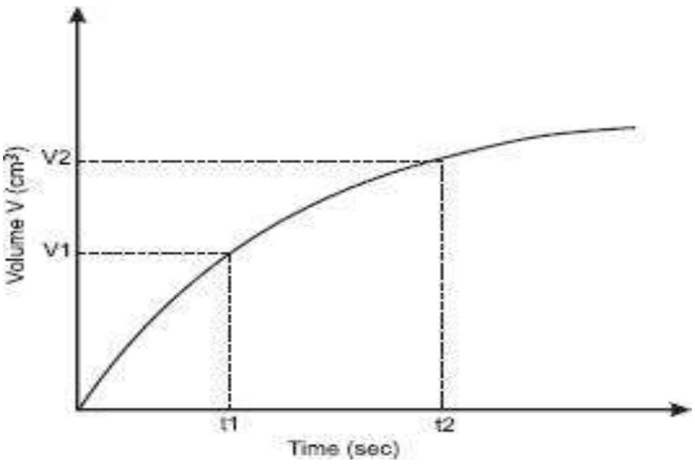
The rate of reaction is inversely proportional to the time taken.

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

Consider the reaction between magnesium and dilute hydrochloric acid whichproceeds with the production of hydrogen gas.

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

The graph obtained when the volume of the gas produced is plotted against time is a curve.



It is clear that as the reaction progresses, a lower volume of hydrogen gas is liberated within a given time interval. Both the amount of magnesium and the concentration of the acid are decreasing during the experiment. The rate of production of hydrogen gas is highest at the start of the experiment, then it slows down and eventually stops when one or all the reactants are used up.

The average rates of reactionscan be worked out by dividing the change in volume of gas at selected time intervals as follows;

$$\text{Av. rate} = \frac{\text{Volume of hydrogen produced}}{\text{Time taken}}$$

$$= \frac{(V_2 - V_1)\ \text{cm}^3}{t_2 - t_1\ \text{sec}}$$

3. Temperature

The rate of a reaction increases with a rise in temperature.

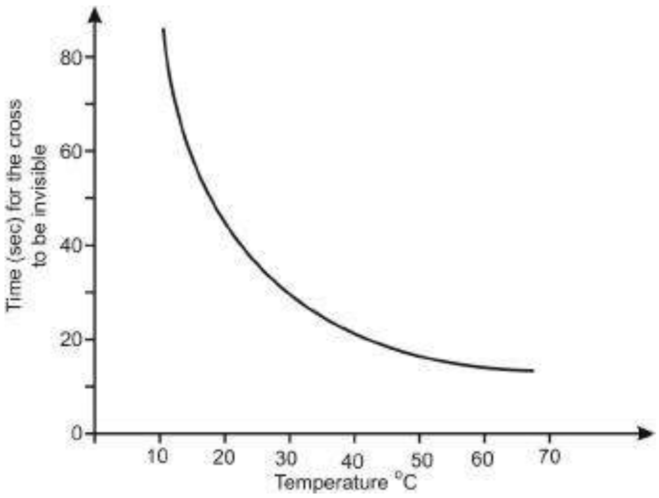
During a reaction, an increase in temperature results in an **increase in the kinetic energy of the particles.** This makes the particles to **move faster and collide more frequently** with sufficient energy to cause more effective collisions per given time. Thus the reaction takes place faster resulting in faster formation of the products.

For example, the reaction between sodium thiosulphate and hydrochloric acid forms a yellow precipitate of sulphur.

$$Na_2S_2O_3(aq) + 2HCl(aq) \longrightarrow 2NaCl(aq) + SO_2(g) + H_2O(l) + S(s)$$

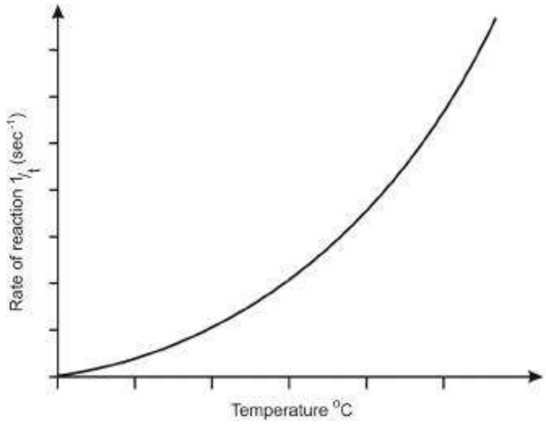
As more and more sulphur is formed, the cross becomes less and less visible until it disappears. The time taken for the cross to be invisible is shorter at higher temperature. Therefore the reaction is faster at higher temperature.

The graph of time versus temperature is a curve. At high temperature less time is required to produce the same amount of sulphur.



It can be seen that the rate of reaction almost doubles with every 10°C temperature rise.

The graph of $\frac{1}{t}$ (sec^{-1}) versus temperature t (°C) shows that the rate increases with a rise in temperature.



4. Surface area.

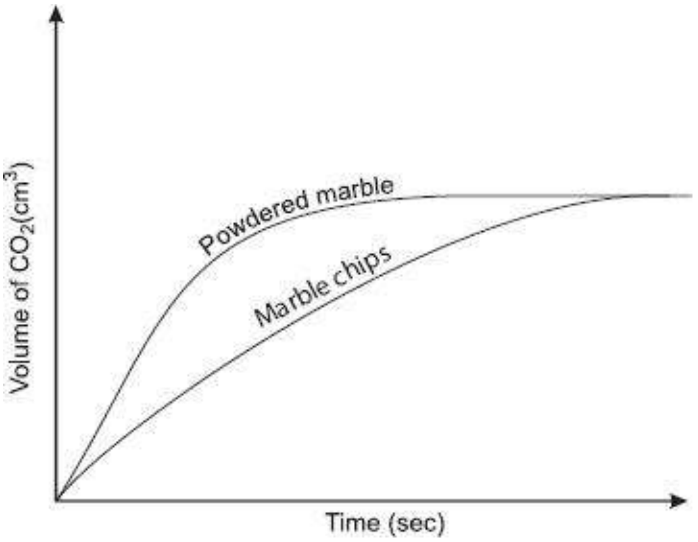
The **greater the surface area of contact between reacting particles, the higher the rate of reaction**.

An **increases in the surface area** of the particles **increases the area of contact** of the reacting particlesand hence the reaction proceeds faster.

For example, dilute hydrochloric acid reacts with marble to liberate carbon(IV) oxide gas.

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

- The volume of the gas produced per unit time is used to determine the rate of the reaction.
- The reaction is faster at the start due to the high concentration of the reactants.
- When a graphs of volume of CO_2 produced is plotted against time for marble chips and powdered marble, it is observed that the curve for powdered marble is steeper than that of the marble chips. This means that carbon(IV) oxide gas is being produced faster with marble powder.

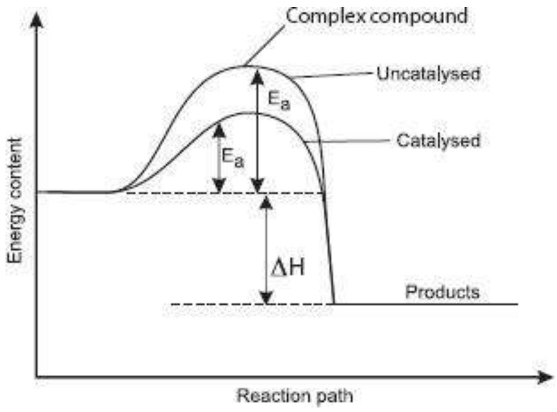


- Both reactions slow down as the reactants are used up and finally both curves flatten out at the same final volume. The total volume of carbon(IV) oxide gas produced in both experiments is the same because the mass of marble used is the same
- ❖ When solid substances are ground into a powder**, the sizes of the particles are reduced. This increases the surface area of the particles thus increasing the area of contact of the reacting particles.** The greater the surface area of contact between reacting particles, the higher the rate of reaction. This is why the reaction of powdered marble with acid takes place more vigorously and reaches completion faster.

5. Catalysts

A catalystis a substance which alters the rate of a chemical reaction but remains chemically unchanged at the end of the reaction.

Catalysts **lower the activation energy** of the of reaction.They **provide a shorter route** for reactions, hence **minimising the energy required for the reaction to proceed** and hence increases the initial rate of a reaction.

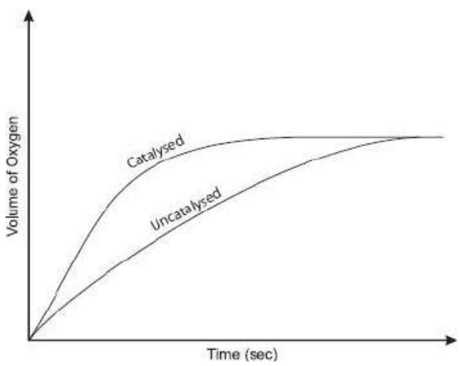


Comparison of catalysed and uncatalysed reactions

For example, Hydrogen peroxide decomposes to form water and oxygen gas.

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

The rate of the decomposition can be determined by using the volume of oxygen produced per unit time. The production of oxygen is very slow at room temperature. However, addition of manganese(IV) oxide,which acts as a catalyst, makes the reaction faster and a large volume of oxygen is collected within a short time. If the volume of oxygen collected in each case is plotted against time, the curves obtained are as shown below.



The mass of manganese(IV) oxide remains unchanged at the end of the experiment since it is not used up during the reaction.

A particular catalyst alters the rates of certain reactions but not all reactions. Some common catalysts are shown below.

Catalyst	Industrial process
Iron	Manufacture of ammonia in Haber process
Vanadium(V) oxide platinum	Manufacture of sulphuric acid contact process
Platinum	Manufacture of nitric(V) acid

The energies of the reactants and the products remain the same but the activation energy is lowered.

6. Light

Light influences the rate of some chemical reactions by **energising some of the particles** involved. This **increases the number of effective collisions per unit time thus increasing the rate of reaction.**

For example, Potassium bromide reacts with silver nitrate to form a pale yellow precipitate of silver bromide.

$$KBr(aq) + AgNO_3(aq)) \longrightarrow AgBr(s) + KNO_3(aq)$$

In the presence of light, silver bromide decomposes to form metallic silver and bromine. Inthe presence of light, the colour of the precipitate changes from pale yellow to grey due to the formation of metallic silver. In the absence of light, there is no noticeable colour change of the precipitate .

The degree of change depends on the intensity of light falling on the test-tube. The rate of decomposition of the silver bromide is greater when exposed to more light.

The following are some examples of photochemical reactions.

(a) **Photography**

Photographic plates are normally coated with silver bromide. Processing of films is done in a dark room to avoid decomposition of the silver bromide.

(b) **Reaction between chlorine and hydrogen**

Under ordinary temperature and pressure, chlorine and hydrogen do not react in darkness. However

$$Cl_2(g) + H_2(g) \xrightarrow{UV\ light} 2HCl(g)$$

exposed to direct light, the reaction is explosive.

(c) **Reaction between bromine and methane**

Bromination of methane does not take place in darkness. But in the presence of light the reaction takes place rapidly.

$$CH_4(g) + Br_2(l) \xrightarrow{UV\ light} CH_3Br(g) + HBr(g)$$

(d) **Photosynthesis**

During photosynthesis, carbon(IV) oxide and water are converted into sugar and oxygen which are useful for plant growth.

7. Effect of Pressure on the Rate of Reaction

Pressure **has no effect** on reactions involving solids and liquids. It **only affects reactions in which either the reactant or products are gases.**

Increasing the pressure of reacting gases by compressing **brings the gaseous particles closer** thus increasing

$$6CO_2(g) + 6H_2O(l) \xrightarrow{Light} C_6H_{12}O_6(aq) + 6O_2(g)$$

the rate of effective collisions. This results in an **increase in the rate of a reaction.**

Reversible Reactions

Some chemical reactions are reversible, i.e., they can go forward or backward depending on the prevailing conditions.

A reaction which can proceed in both directions is called a **reversible** reaction.

When white solid ammonium chloride is warmed gently in a test-tube, it decomposes into ammonia and hydrogen chloride gases. The process is called **thermal dissociation**.

$$NH_4Cl(s) \longrightarrow NH_3(g) + HCl(g)$$

On the cooler part of the test-tube, ammonia and hydrogen chloride **recombine** to form ammonia chloride.

$$NH_3(g) + HCl(g) \longrightarrow NH_4Cl(s)$$

Production of ammonia chloride when ammonia and hydrogen chloride recombine is a **reverse reaction** to the thermal dissociation of ammonia chloride. The **forward** and the **reverse** (backward) are represented by **double arrows** in an equation.

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl\ (g)$$

When hydrated copper(II) sulphate crystals are heated, a white powder is formed. The crystals lose water of crystallisation.

$$\underset{\textit{Blue crystals}}{CuSO_4.5H_2O(s)} \longrightarrow \underset{\textit{White powder}}{CuSO_4} + \underset{\textit{Colourless liquid}}{5H_2O(l)}$$

When the colourless liquid obtained is poured back into the solid residue, the white anhydrous copper(II) sulpahte turns blue.

$$\underset{\textit{White powder}}{CuSO_4(s)} + \underset{\textit{water}}{5H_2O(l)} \longrightarrow \underset{\textit{Blue crystals}}{CuSO_4.5H_2O(s)}$$

A reversible reaction like the one above is represented by double arrows ($\rightleftharpoons$).

$$CuSO_4.5H_2O(s) \underset{cool}{\overset{heat}{\rightleftharpoons}} CuSO_4\ (s) + 5H_2O(l)$$

Otherexamples of reversible reactions are:

(i) $\underset{\textit{pink}}{COCl_2.6H_2O(s)} \underset{cool}{\overset{heat}{\rightleftharpoons}} \underset{\textit{blue}}{COCl_2} + 6H_2O(g)$

(ii) $\underset{\textit{grey}}{3Fe(s)} + 4H_2O(g) \rightleftharpoons \underset{\textit{brown}}{Fe_3O_4(s)} + 4H_2(g)$

(iii) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

Equilibrium

Equilibrium is a **state of balance** that occurs when both the forward and backward reactions are taking place at the same rate but in opposite directions.

A dynamic equilibrium is attained when the rate of the forward reaction is equal to that of the backward reaction.

If a sample of ammonium chloride is heated in a closed container, both the forward and backward reactions take place.

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

If the reactions attain equilibrium, the sign $\rightleftharpoons$ replaces the sign $\rightleftharpoons$ for reversible reactions which cannot attain equilibrium.

The equations of the reaction can now be written as,

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)?$$

Factors Affecting Equilibrium

The state of balance in reversible reactions can be affected by changing physical conditions such as temperature, pressure and concentration. They affect the equilibrium by either favouring the forward or backward reaction.

Le Chatelier's principle.

"When a change in conditions is applied to a system in equilibrium, the system moves so as to oppose that change."

Effect of Change in Concentration on Equilibrium

If substance A and B react to form product C and D, then changing the concentration of any of the components will shift the equilibrium in such a way as to oppose the change.

$$A(aq) + B(aq) \rightleftharpoons C(aq) + D(aq)$$

If the concentration of B is increased, then the reaction will favour the direction which helps in reducing the concentration of B. This is achieved by more of B reacting with A to form C and D. If the concentration of D is increased, the backward reaction will be favoured.

The effect of change of concentration on a reaction at equilibrium can be demonstrated using the chromate/dichromate reaction.

$$\underset{\textit{Yellow}}{2Cr_2O_4^{2-}(aq) + 2H^+(aq)} \rightleftharpoons \underset{\textit{Orange}}{Cr_2O_7^{2-}(aq) + H_2O(l)}$$

Addition of acid causes more chromate (CrO_4^{2-}) ions to change to dichromate ($Cr_2O_7^{2}$) ions making the solution to become orange. Addition of hydroxide ions reverses the reaction leading to the formation of the yellow chromate solution.

$$Cr_2O_7^{2-}(aq) + OH^-(aq) \rightleftharpoons 2CrO_4^{2-}(aq) + H_2O(l)$$

Altering the concentration of any one of the components of the equilibrium mixture disturbs the equilibrium by **making the reaction rate in one direction greater than the other, hence the reaction proceeds predominantly in that direction until the equilibrium is re-established.**

The observations are in agreement with **Le Chatelier's Principle.**

Effect of Pressure on Equilibrium

A change in pressure will only affect equilibrium in reactions involving gases. It has no effect on those reactions in which both the reactants and products are either solids or liquids since they cannot be compressed or expanded.

Pressure is caused by collisions of the gas particles with the walls of the reaction vessel. The more the number of molecules present, the higher the number of collisions and hence the higher the pressure. I**f the pressure of a system in equilibrium is increased, the system will favour the direction that has fewer molecules.**

$$3A(g) + B(g) \longrightarrow 2C(g)$$

3 moles	1 mole	2 moles
(3 + 1) moles		2 moles
4 moles		2 moles
4 vol		2 vol

Consider the following reaction involving gases

The left-hand side has higher pressure (4 molecules) than the right hand side (molecules). Applying more pressure will favour the forward reaction since it reduces the pressure. A decrease in pressure favours the direction of reaction that forms more molecules.

The effect of changing pressure can be illustrated using the equilibrium established between nitrogen(IV) oxide and dinitrogen tetroxide.

$$\underset{\textit{Dark-brown}}{2NO_2(g)} \rightleftharpoons \underset{\textit{Pale yellow}}{N_2O_4(g)}$$

Consider what happens if the two gases are enclosed in a syringe. If the piston is pushed in, the volume of the gases is reduced hence increasing pressure. The contents of the syringe appear pale yellow. This shows that an increase in pressure favours the forward reaction which forms fewer molecules.

When the piston is withdrawn slowly, the contents of the syringe darken as the volume is increased therefore lowering the pressure. This shows that backward reaction which forms more molecules is favoured by lowering the pressure.

If the number of molecules of reactants equals number of molecules of products, a change in pressure does not have any effect on the equilibrium.

For example:

$$\underset{\textit{1 mol}}{I_2(g)} + \underset{\textit{1 mol}}{H_2(g)} \longrightarrow \underset{\textit{2 mol}}{2HI(g)}$$

Change in pressure will not affect the position of equilibrium but will help in the quick attainment of the equilibrium. The effects of change in pressure on equilibrium may be summarised as in table below.

Effect of pressure on equilibrium

Reaction	Effect of pressure equilibrium change on the mixture	
	Increase	**Decrease**
$N_2O_4(g) \rightleftharpoons 2NO_2(g)$	More N_2O_4 formed (backward reaction favoured).	More NO_2 formed (forward reaction favoured).
$N_2(g) + 2H_3(g) \rightleftharpoons 2NH_2(g)$	More of the reactants formed due to increase in no. of molecules.	More NH_3 formed due to decrease in volume.
$2SO_2(g)\ O_2(g) \rightleftharpoons 2SO_2(g)$	More SO_2 and O_2 formed. Backward reaction forms more molecules.	More SO_3 formed Forward reaction forms fewer molecules.
$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$	More NO and H_2O formed backward reaction accompanied by decrease in volume.	More NH_3 and O_2 formed. Forward reaction accompanied by increase in the number of molecules.
$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$	No effect. No change in number of molecules.	No effect because number of molecules remain the same.

Effect of change of temperature on equilibrium.

A change in temperature alters the position of a chemical equilibrium.

A decrease in temperature favours the reaction which liberates heat (exothermic reaction). Subsequently, an increase in temperature favors a reaction which proceeds with absorption of heat (endothermic reaction).

The effect of change in temperature on equilibrium can be summarized by considering the reaction A + B $\rightleftharpoons$ C +D $\Delta H = -ve$

Effect of temperature on equilibrium

Reaction	**Effect of change in temperature on equilibrium**	
	Increase	*Decrease*
Exothermic (ΔH -ve)	Shifts to the left and more of A and B would be formed.	Shifts to the right and more of C and D would be formed.
Endothermic (ΔH +ve)	Shifts to the right and more of C and D would be formed.	Shifts to the left and more of A and B would be formed.

For example, Copper turnings react with concentrated nitric(V) acid to produce brown nitrogen(IV) oxide gas ($NO_2(g)$).

$$Cu(s) + 4HNO_3(l) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

When the flask is cooled, the colour of the gas inside changes from brown to yellow. This is because on cooling, the brown nitrogen(IV) oxide (NO_2) dimerises (i.e., two molecules of the gas combine to form one molecule) to form dinitrogen tetra oxide (N_2O_4). At room temperature and pressure an equilibrium between the two gases is established.

$$\underset{\textit{Pale yellow}}{N_2O_4(g)} \underset{cool}{\overset{heat}{\rightleftharpoons}} \underset{\textit{Dar-brown}}{2NO_2(g)}$$

When the mixture is heated, the dinitrogen tetraoxide molecules break up to form nitrogen (IV) oxide molecules. The mixture turns a darker brown colour. The forward reaction in the equation is endothermic since rising the temperature will favour the process that absorbs heat. In this case, the equilibrium shifts from left to right and that is why the mixture becomes dark-brown.

A decrease in temperature favours the reaction which liberates heat. Since backward reaction is exothermic, the colour of the mixture becomes pale-yellow.

Catalysts and Equilibrium.

A catalyst has no effect on the position of equilibrium, it only alters the rate at which the equilibrium is attained.

Some Industrial Applications of Chemical Equilibrium

The conditions required to obtain the highest yield of the products in the shortest time at minimum cost are referred to as optimum conditions.

The position of equilibrium is normally changed by removal of one of the products or through variation of pressure and/or temperature.

The Haber Process

The reaction between nitrogen and hydrogen in the manufacture of ammonia is a reversible reactions. It provides a good example of the application of the equilibrium principle.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \; \Delta H = -92 \text{ kJ mol}^{-1}$$

The forward reaction is exothermic and therefore an increase in temperature favours a shift in the equilibrium from right to left, i.e., ammonia breaks down to form nitrogen and hydrogen. This implies that the yield of ammonia is less at high temperatures. Lowering the temperature will favour production of more ammonia. However, the rate at which ammonia is produced will be too slow and therefore uneconomical. An optimum temperature of about 450°C is normally used.

In addition, the forward reaction is accompanied by a decrease in the number of molecules, i.e., volume. Increase in pressure therefore, favours the process of reaction that will tend to lower the pressure. Therefore high pressure favours the production of more ammonia. High pressure also increases the rate of reaction. However the cost of producing and maintaining high pressure in a system is quite enormous. Therefore an optimum pressure of 200 atmospheres is normally applied.

If some of the ammonia is removed from the system at equilibrium, more of the nitrogen and hydrogen will react to produce ammonia to replace that which is removed. Normally ammonia is removed as soon as it is produced so that more of the nitrogen and hydrogen can continue reacting.

The catalyst used for this process is iron. Though platinum is a better catalyst, it is very expensive and easily poisoned by impurities.

The Contact Process

In the manufacture of sulphuric(VI) acid, the main reaction taking place between sulphur (IV) oxide and oxygen to produce sulphur(VI) oxide is also a reversible reaction:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -197 \text{ kJ mol}^{-1}$$

The forward reaction is exothermic therefore an increase in temperature favours the backward reaction. Therefore, at high temperature less sulphur(VI) oxide will be formed. Low temperatures will favour the forward reaction but the rate will be slow. For maximum yield, an optimum temperature of 450°C is used.

The forward reaction is accompanied by a decrease in the number of molecules from 3 to 2. Therefore an increase in pressure will favour production of more sulphur(IV) oxide. The optimum pressure used is atmospheric pressure. This gives a percentage conversion of sulphur(IV) oxide to sulphur(VI) oxide of about 96%.

Vanadium(V) oxide is used as a catalyst to increase the rate of the reaction.

1. 2006 Q 9 P1

At 20 °C, NO_2 and N_2O_4 gases exist in equilibrium as shown in the equation below

$$2\,NO_2(g) \rightleftharpoons N_2O_4(g); \quad \Delta H = -ve$$

Brown *pale yellow*

State and explain the observation that would be made when;

(a) A syringe containing the mixture at 20 °C is immersed in ice-cold water; (1 mark)

(b) The volume of the gaseous mixture in a syringe is reduced. (1 mark)

2. 2006 Q 7c-g P2

The table below shows the volumes of nitrogen dioxide gas produced when different volume of 1M nitric acid were each reacted with 2.07 g of lead at room temperature.

Volume of 1 M nitric acid (cm^3)	Volume of nitrogen dioxide gas (cm^3)
5	60
15	180
25	300
35	420
45	480
55	480

(a) Give a reason why nitric acid is not used to prepare hydrogen gas. (1 mark)

(b) Explain how the rate of the reaction between lead and nitric acid would be affected if the temperature of the reaction mixture was raised. (2 marks)

(c) On the grid provided below, plot a graph of the volume of the gas produced (Vertical axis) against volume of acid. (3 marks)

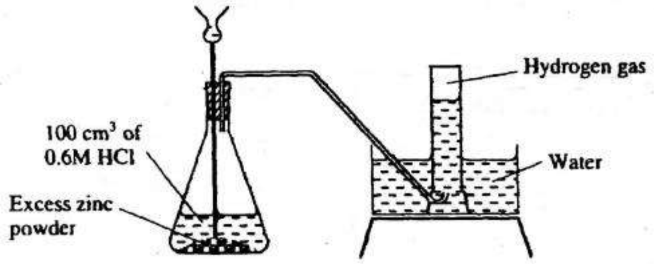
(d) Using the graph, determine the volume of:
   (i) Nitrogen dioxide produced when $30cm^3$ of 1 M nitric acid were reacted with 2.07 g of lead (1 mark)
   (ii) 1M nitric acid which would react completely with 2.07g of lead. (1 mark)

(e) Using the answer in d(i) above, determine:
   (i) The volume of 1M nitric acid that would react completely with one mole of lead (Pb=207) (2 marks)
   (ii) The volume of nitrogen dioxide gas produced when one mole of lead reacts with excess 1 M nitric room temperature. (1 mark)

(f) Calculate the number of moles of:
   (i) 1M nitric acid that reacted with one mole of lead (1 mark)
   (ii) Nitrogen dioxide produced when one mole of lead were reacted with excess nitric acid. (Molar gas volume of $2400cm^3$) (1 mark)

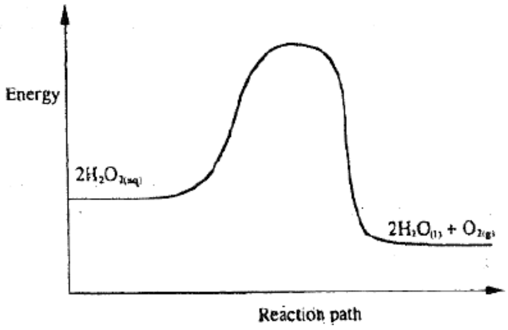
(g) Using the answers obtained in (f) (i) and (ii) above, write the equation for the reaction between lead and nitric acid given that one mole of lead nitrate and two moles of water were also produced. (1 mark)

3. 2007 Q 25 P1
The diagram below shows a student's set-up for the preparation and collection of hydrogen gas.



(a) How would the final volume of hydrogen gas produced be affected if $80cm^3$ of 0.75 M hydrochloric acid was used? (1 mark)
(b) Give a reason why helium is increasingly being preferred to hydrogen in weather balloons. (1 mark)

4. 2007 Q 27
The diagram below is a sketch of the graph of the non-catalysed decomposition of hydrogen peroxide. (1 mark)



On the same axis, sketch the graph for the decomposition of hydrogen peroxide when manganese (IV) oxide is added. (2 marks)

5. 2007 Q 4 P2, 2016 P2 Q3.

(a) Methanol is manufactured from carbon (IV) oxide and hydrogen gas according to the equation:

$$CO_2\ (g) + 3H_2\ (g) \rightleftharpoons CH_3OH\ (g) + H_2O\ (g)$$

The reaction is carried out in the presence of a chromium catalyst at 700K and 30kPa. Under these conditions, equilibrium is reacted when 2% of the carbon (IV) oxide is converted to methanol.

(i) How does the rate of the forward reaction compare with that of the reverse reaction when 2% of the carbon (IV) oxide is converted to methanol? (1 mark)

(ii) Explain how each of the following would affect the yield of methanol:

I. Reduction (2 marks)

II. Using a more efficient catalyst (2 marks)

(iii) If the reaction is carried out at 500K and 30kPa, the percentage of carbon (IV) oxide converted to methanol is higher than 2%

I. What is the sign of ΔH for the reaction? Give a reason (2 marks)

II. Explain why in practice the reaction is carried out at 700K but NOT at 500K

(1mark)

(b) Hydrogen peroxide decomposes according to the following equation:

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

In an experiment, the rate of decomposition of hydrogen peroxide was found to be 6.0×10^{-8} mol $dm^{-3}s^{-1}$.

(i) Calculate the number of moles per dm^3 of hydrogen peroxide that had decomposed within the first 2 minutes.
(2 marks)

(ii) In another experiment, the rate of decomposition was found to be 1.8×10^{-7} mol $dm^{-3}s^{-1}$. The difference in two rates could have been caused by addition of a catalyst. State, giving reasons, one other factor that may have caused the difference in two rates of decomposition.
(2 marks)

6. 2008 Q 23

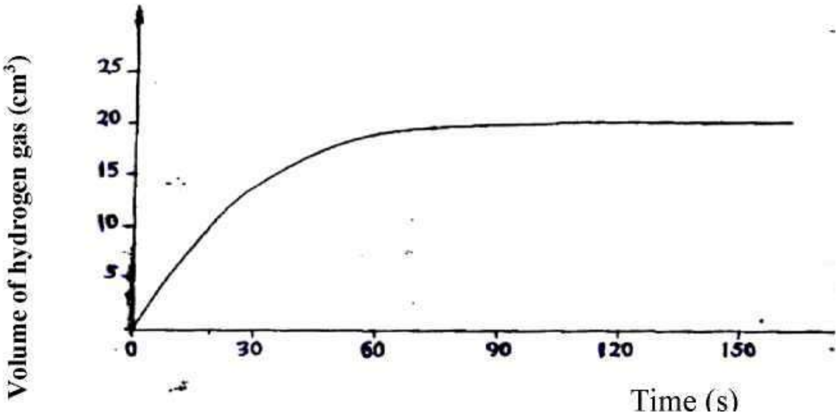
In a closed system, aqueous iron (III) chloride reacts with sulphide gas as shown in the equation below.

$$2FeCl_3\ (aq) + H_2S\ (g) \rightleftharpoons 2FeCl_2\ (aq) + 2HCl\ (aq) + S(s)$$

State and explain the observation that would be made if dilute hydrochloric acid is added to the system at equilibrium. (2 marks)

7. 2008 Q 29

A certain mass of a metal E_1 reacted with excess dilute hydrochloric acid at 25 °C. The volume of hydrogen gas liberated was measured after every 30 seconds. The results were presented as shown in the graph below.



(a) Name one piece of apparatus that may have been used to measure the volume of gas liberated. (1 mark)

(b) (i)On the same axis, sketch the curve that would be obtained if the experiment was repeated at 35 °C. (1 mark)

(ii) Explain the shape of your curve in b(i) above. (1 mark)

8. 2008 Q4 P2

(a) (i) State the Le chatelier's principle. (1mark)

(ii) Carbon (II) oxide gas reacts with steam according to the equation;

$$CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$$

What would be the effect of increasing the pressure of the system at equilibrium? Explain. (2marks)

(b) The table below gives the volumes of oxygen gas produced at different times when hydrogen peroxide decomposed in the presence of a catalyst.

Time (Sec)	0	10	20	30	40	50	60
Volume of oxygen (cm^3)	0	66	98	110	119	120	120

(i) Name the catalyst used for this reaction (1mark)

(ii) On the grid provided, draw the graph of volume of oxygen gas produced (vertical axis) against time. (3marks)

(iii) Using the graph, determine the rate of decomposition of hydrogen peroxide after 24 seconds. (2marks)

(iv) Give a reason why the total volume of oxygen gas produced after 50 seconds remains constant. (1mark)

9. 2009 Q 27

The following reaction is in equilibrium in a closed container.

$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

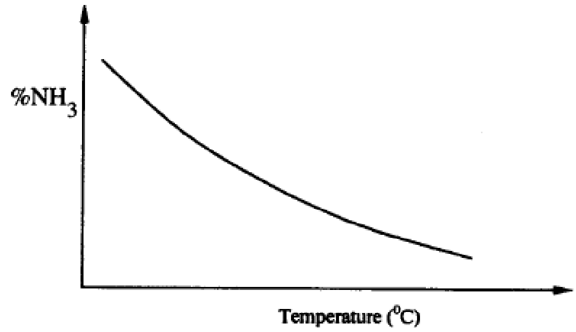
State giving reasons how an increase in pressure would affect the amount of hydrogen. (2 marks)

10. 2010 Q 22 P1

Nitrogen and hydrogen react to form ammonia gas as shown in the following equation:

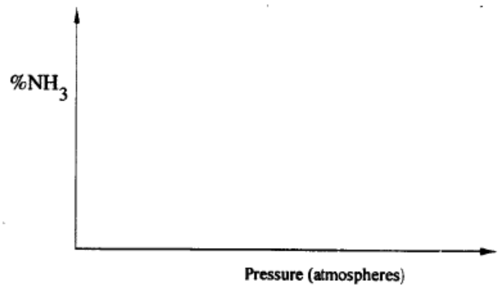
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; ΔH is negative

(a) The figure below shows how the percentage of ammonia gas in the equilibrium mixture change with temperature.



Explain why the percentage of ammonia gas change as shown in the figure. (2 marks)

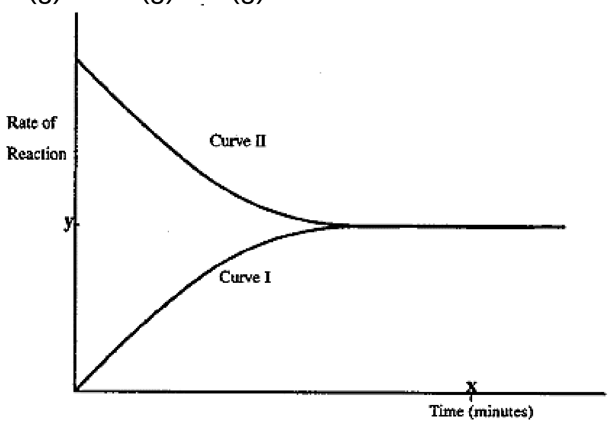
(b) On the axes below, sketch a graph showing how the percentage of ammonia gas in equilibrium mixture changes with pressure. (1 mark)



11. 2010 Q 7 P2

(a) The figure below shows how the rate of the following reaction varies with the time.

$A (g) + B(g) \rightleftharpoons 2C(g) + D(g)$



(i) Which of the two curves represent the rate of the reverse reaction? Give a reason (2 marks)

(ii) What is the significance of point **X** and **Y** on the figure? (2 marks)

(b) State and explain the effect of an increase in pressure on the rates of the following reactions.
(2 marks)
I. $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$
II. $CH_3OH(l) + CH_3COOH(l) \longrightarrow CH_3COOCH_3(l) + H_2O(l)$

(c) In an experiment to study the rate of reaction between barium carbonate and dilute hydrochloric acid; 1.97g of barium carbonate were reacted with excess 2M hydrochloric acid. The equation for the reaction is

$$BaCO_3(s) + 2HCl(aq) \longrightarrow BaCl_2(aq) + CO_2(g) + H_2O(l)$$

The data in the table was obtained

Time in seconds	0	30	60	90	120	150	180	210	240
Volume of gas (cm^3)	0	80	135	175	210	230	240	240	240

(i) On the grid provided, plot a graph of volume of gas produced (vertical axis) against time
(3 marks)

(ii) From the graph, determine the rate of the reaction at:
(I) 15 seconds (1 mark)
(II) 120 seconds (1 mark)
(III) Give a reason for the difference between the two values.
(1 mark)

12. 2011 Q 3
(a) Ethanol can be manufactured from ethane and steam as steam as shown in the equation below:

$$C_2H_4\ (g) + H_2O\ (g) \rightleftharpoons CH_3CH_2OH\ (g)$$

Temperature and pressure will affect the position of equilibrium of the above reaction. Name the other factor that will affect the position of equilibrium of the above reaction. (1 mark)

(b) The data in the table was recorded when one mole of ethane was reacted with excess steam. The amount of ethanol in the equilibrium mixture was recorded under different conditions of temperature and pressure. Use the data to answer the questions that follow.

Temperature (°C)	**Pressure (Atm)**	**Amount of ethanol at equilibrium (moles)**
300	50	0.40
300	60	0.46
300	70	0.55
250	50	0.42
350	50	0.38

(i) State whether the reaction between ethane and steam is exothermic or endothermic. Explain your answer.
(3 marks)

(ii) State and explain one advantage and one disadvantage of using extremely high pressure in this reaction.
I. Advantage (2 marks)
II. Disadvantage (2 marks)

(c) In an experiment to determine the rate of reaction between calcium carbonate and dilute hydrochloric acid, 2g of calcium carbonate were reacted with excess 2M hydrochloric acid.The volume of carbon (IV) oxide evolved was recorded at regular intervals of one minute for six minutes. The results are as shown in the table below.

Time (minutes)	1	2	3	4	5
Volume of Carbon (IV)oxide (cm^3)	170	296	405	465	480

(i) Plot a graph of time in minutes on the horizontal axis against volume of carbon (IV) oxide on the vertical axis. (3 marks)

(ii) Determine the rate of reaction at 4 minutes. (2 marks)

13. 2012 Q13 P1

A dynamic equilibrium is established when hydrogen and carbon (IV) oxide react as shown below:

$$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$$

What is the effect of adding powdered iron catalyst on the position of the equilibrium? Give a reason. (2 marks)

14. 2012 Q4 P2

The factors which affect the rate of reaction between lead carbonate and dilute nitric (V) acid were investigated by carrying out three experiments:

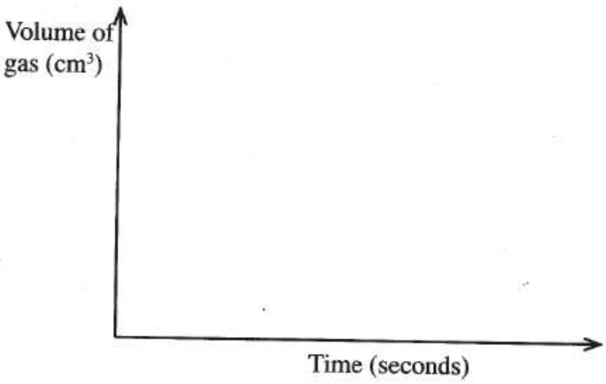
Experiment	**Lead carbonate**	**Concentration of nitric (V) acid**
1	Lumps	4M
2	Powdered	4M
3	Lumps	2M

(a) Other than concentration, name the factor that was investigated in the experiments. (1 mark)

(b) For each experiment, the same volume of acid (excess) and mass of lead carbonate were used and the volume of gas liberated measured with time.

(i) Draw a set up that can be used to investigate the rate of reaction for one of the experiments. (3 marks)

(ii) On the grid provided, sketch the curves obtained when the volume of gas produced was plotted against time for each of the experiments and label each as 1, 2, or 3. (4 marks)



(iii) Write an equation for the reaction that took place. (1 mark)

(c) If the experiments were carried out using dilute hydrochloric acid in place of dilute nitric (V) acid, the reaction would start, slow down and eventually stop. Explain these observations.
(2 marks)

(d) A solution of bromine gas in water is an example of a chemical reaction in a state of balance. The reaction involved is represented by the equation below.

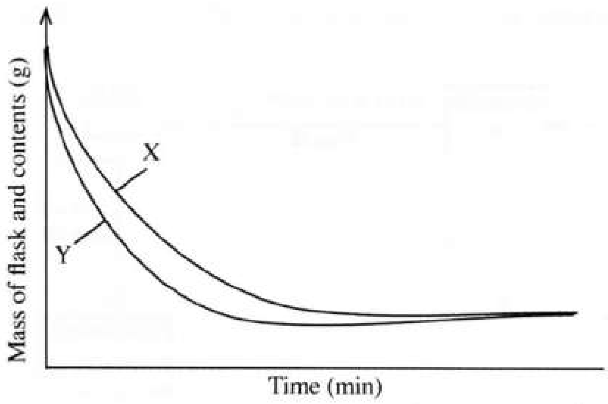
$$Br_2(g) + H_2O(l) \rightleftharpoons 2H^+(aq) + Br^-(aq) + OBr^-(aq)$$

Yellow/orange *Colourless*

State and explain the observations made when hydrochloric acid is added to the mixture.
(2 marks)

15. 2013 Q16 P1

(a) The curves below represent the change in mass when equal masses of powered zinc and zinc granules were reacted with excess 2M hydrochloric acid. Study them and answer the question below.



Which curve represents the reaction with zinc granules? Explain your answer. (2 marks)

16. 2014 Q 17 P1

In an experiment on rates of reaction, potassium carbonate was reacted with dilute sulphuric (VI) acid.

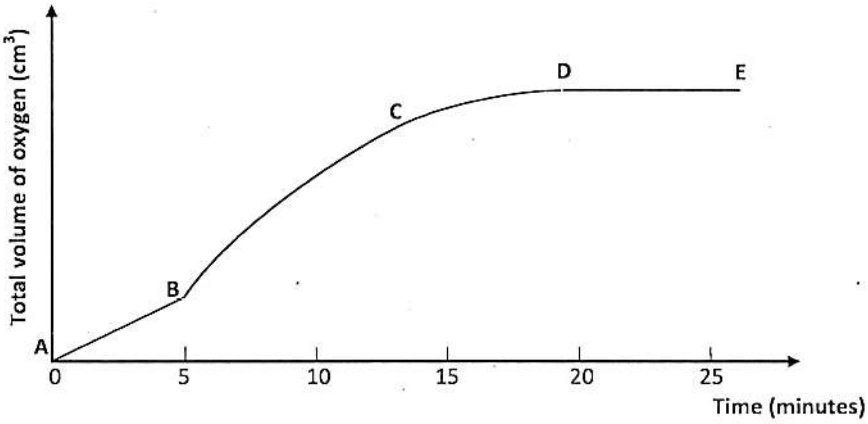
(a) What would be the effect of an increase in the concentration of the acid on the rate of the reaction? (1 mark)

(b) Explain why the rate of reaction is found to increase with temperature. (2 marks)

17. 2014 Q 4 P2

(a) Other than temperature, state two factors that determine the rate of a chemical reaction. (2 marks)

(b) A solution of hydrogen peroxide was allowed to decompose and the oxygen gas given off collected. After 5 minutes, substance G was added to the solution of hydrogen peroxide. The total volume of oxygen evolved was plotted against time as shown in the graph below



(i) Describe the procedure of determining the rate of the reaction at minute 12. (3 marks)

(ii) How does the production of oxygen in region **AB** compare with that in region **BC**? Explain (2 marks)

(iii) Write an equation to show the decomposition of hydrogen peroxide. (1 mark)

(c) Sulphur (IV) oxide react with oxygen to form sulphur (VI) oxide as shown in the equation below

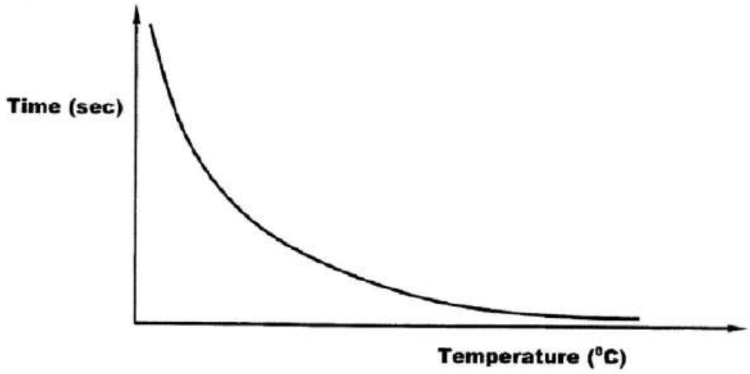
$$2SO_2(g) + O_2(g) \quad 2SO_3(g): \Delta H = -192kJ$$

(i) Explain the effect on the yield of SO_3 of lowering the temperature of this reaction. (2 marks)

(ii) Name one catalyst used for the reaction. (1 mark)

18. 2015 Q12 P1

The curve shown below shows the variation of time against temperature for the reaction between sodium thiosulphate and hydrochloric acid.



(a) Write the equation for the reaction between sodium thiosulphate and dilute hydrochloric acid. (1 mark)
(b) Explain the shape of the curve (2 marks)

19. 2015 Q6 P2
(a) Other than concentration, state two factors that determine the rate of a reaction. (2 marks)

(b) In an experiment to determine the rate of reaction, excess lambs of calcium carbonate were added to 2 M hydrochloric acid. The mass of calcium carbonate left was recorded after every 30 seconds. The results are shown in the table below.

Time (seconds)	0	30	60	90	120	150	180	210
Mass of calcium carbonate left (g)	2.00	1.60	1.30	1.00	0.85	0.8	0.8	0.8

(i) Write the equation for the reaction that took place. (1 mark)
(ii) On the grid provided, plot a graph of mass of calcium carbonate vertical axis against time. (3 marks)
(iii) Determine the rate of reaction at the 105th second. (3 marks)

(c) Why does the curve level off after some time? (1 mark)

(d) On the same grid, sketch a curve for the same reaction using 4 M hydrochloric acid and label the curve **R**. (2 marks)

20. 2016 Q19 P1
In an experiment on rates of reaction, potassium carbonate was reacted with dilute sulphuric (VI) acid.
(a) What would be the effect of an increase in the concentration of the acid on the rate of the reaction? (1 mark)
(b) Explain why the rate of reaction is found to increase with temperature. (1mark)

21. 2017 P1 Q8.
(a) State one characteristic of a reaction where equilibrium has been attained. (1 mark)
(b) The following equation is in a state of equilibrium: C D Use it to sketch a graphical

representation of concentration against time in seconds for the equilibrium. (2 marks)

22. 2018 P1 Q 13.
In the Haber process, nitrogen reacts with hydrogen according to the following equation.

$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$; $\Delta H = -92\ kJmol^{-1}$

(a) What would be the effect of adding a catalyst on the position of the equilibrium? (1 mark)

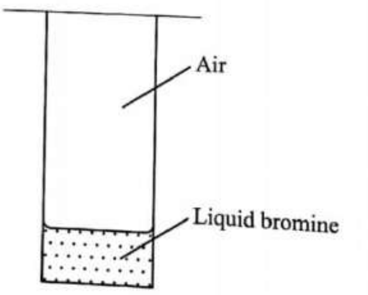
(b) Explain why it is not advisable to use temperatures higher than 773 K in the Haber process. (2 marks)

23. 2019 P1 Q2.
Describe how an increase in concentration increases the rate of a reaction. (2 mark)

24. 2019 P2 Q7.
(a) What is meant by rate of reaction? (1 mark)

(b) In the space provided, sketch the diagram of a set-up that can be used to determine the rate of reaction between manganese (IV) oxide and hydrogen peroxide. (3 marks)

(c) A student placed a small amount of liquid bromine at the bottom of a sealed gas jar of air as shown in Figure 4.



(i) Describe what will be observed: (1 mark)
I. After two minutes..............................
II. After 30 minutes ....

(ii) Use the Kinetic theory to explain the observations: (2 marks)
I. After two minutes..............................
II. After 30 minutes ....