DEPARTMENT OF EDUCATION

GRADE 11

CHEMISTRY

MODULE 4

ENERGY AND REACTION RATES

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GRADE 11

CHEMISTRY

MODULE 4

ENERGY AND REACTION RATES

IN THIS MODULE YOU WILL LEARN ABOUT:

11.4.1: FACTORS AFFECTING RATES OF REACTION

11.4.2: ENERGY DIAGRAMS

Acknowledgement

We acknowledge the contributions of all secondary teachers who in one way or another have helped to develop this Course.

Our profound gratitude goes to the former Principal of FODE, Mr. Demas Tongogo for leading FODE team towards this great achievement.

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DIANA TEIT AKIS
PRINCIPAL

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SECRETARY'S MESSAGE

Achieving a better future by individual students and their families, communities or the nation as a whole, depends on the kind of curriculum and the way it is delivered.

This course is a part of the new Flexible, Open and Distance Education curriculum. The learning outcomes are student-centred and allows for them to be demonstrated and assessed.

It maintains the rationale, goals, aims and principles of the national curriculum and identifies the knowledge, skills, attitudes and values that students should achieve.

This is a provision by Flexible, Open and Distance Education as an alternative pathway of formal education.

The course promotes Papua New Guinea values and beliefs which are found in our Constitution, Government Policies and Reports. It is developed in line with the National Education Plan (2005 -2014) and addresses an increase in the number of school leavers affected by the lack of access into secondary and higher educational institutions.

Flexible, Open and Distance Education curriculum is guided by the Department of Education's Mission which is fivefold:

- To facilitate and promote the integral development of every individual
- To develop and encourage an education system that satisfies the requirements of Papua New Guinea and its people
- To establish, preserve and improve standards of education throughout Papua New Guinea
- To make the benefits of such education available as widely as possible to all of the people
- To make the education accessible to the poor and physically, mentally and socially handicapped as well as to those who are educationally disadvantaged.

The college is enhanced through this course to provide alternative and comparable pathways for students and adults to complete their education through a one system, two pathways and same outcomes.

It is our vision that Papua New Guineans' harness all appropriate and affordable technologies to pursue this program.

I commend all the teachers, curriculum writers and instructional designers who have contributed towards the development of this course.

UKE KOMBRA, PhD
Secretary for Education

MODULE 11.4 ENERGY AND REACTION RATES

Introduction

Explosion of fireworks and precipitation are examples of very fast reactions while reactions of metals or carbonates with dilute acids are moderately fast. On the other hand, some reactions like the rusting of iron and the fermentation (conversion of fruit juice into alcohol) of carbohydrates to form wine are slow reactions.

How do we measure the speed or the rate of a chemical reaction in more accurate terms? In what ways is the speed of a reaction affected by external factors such as temperature, pressure, and the nature and quantity of the products?

Is it good to have a fast reaction in the manufacture of chemicals in the industry? If the speed of a chemical reaction is high, does it mean that we can obtain more products from the chemical reaction? As you will learn later, this may not be so! The production of foodstuff and the packaging of fruits and vegetables, a high reaction speed is destructive to the dealers, as it means that the food, fruits, and vegetables will ripen faster and spoil more easily.

In module 11.3, you have learnt that some reactions are faster than others. In this module, you will develop an understanding about why reactions have different rates. You will also look at:

a. Collision theory
b. Factors that affect the rate of reaction
c. Measuring the rate of reaction
d. Exothermic, endothermic reactions and energy diagrams
e. Reversible reactions
f. Bond energy

Learning Outcomes

After going through this module, you are expected to:

- investigate the factors that influence the rate of reaction.
- measure volume, mass or temperature to determine the rate of reaction.
- draw energy diagrams of exothermic and endothermic reactions.
- calculate bond energy.

Time Frame

Suggested allotment time: **7 weeks**

If you set an average of 3 hours per day, you should be able to complete the module comfortably by the end of the assigned week.

Try to do all the learning activities and compare your answers with the ones provided at the end of the module. If you do not get a particular exercise right in the first attempt, you should not get discouraged but instead, go back and attempt it again. If you still do not get it right after several attempts then you should seek help from your friend or even your tutor.

DO NOT LEAVE ANY QUESTION UN-ANSWERED.

Terminologies

Before you get into the thick of things, let us make sure you know some of the terminologies that are used throughout this module.

Activation energy	A minimum energy that is required to start the reaction**.**
Catalyst	Is a substance which increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.
Collision theory	Is a model that describes how the rate of a chemical reaction is determined by the collisions between reacting particles.
Endothermic reaction	Is a reaction in which heat energy is absorbed from the surroundings.
Exothermic reaction	Is a reaction in which heat energy is given out to the surroundings.
Rate	Is a measure of how fast or slow something is. It is a measure of the change that happens in a single unit of time.
Reversible reaction	Is a reaction which can go backwards (reverse) or forward. of the change that happens in a single module of time.

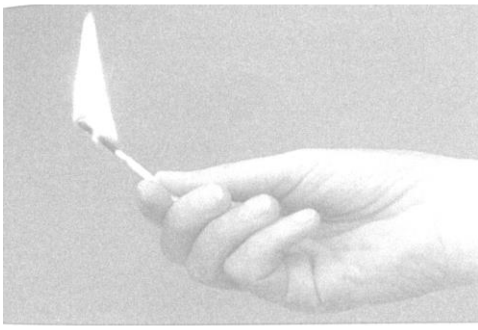
11.4.1 Factors Affecting Rates of Reaction

What is rate?

Rate is a measure of how fast or slow something is. It is a measure of the change that happens in a single module of time. Any suitable unit of time can be used: a second, a minute and an hour.

Rate is a measure of how fast or slow something is. It is a measure of the change that happens in a single unit of time.

Some reactions are fast. Others are slow as mentioned at the start of this module.Can you think of your own example of a reaction, which happens very quickly? Fast reactions, like dynamite exploding, start and finish within a fraction of a second. Slow reactions, like concrete setting, may take days, weeks, or even years to finish. Can you think of another slow reaction?



This combustion lasts for a few seconds.

The chemicals an the base of a party poppers react in fraction of second.

Dynamite exploding

The petrol pump can pump at a rate of 50 litres

A plane has just flown 2000 kilometres in 1 hour. It flew at a rate of 2000 per hour.

The machine can print newspapers at a rate of 10 copies per second.

From the examples given you can see that **rate** is a measure of the change that happens in a single module of time.

What is rate of reaction?

The rate of a reaction tells us how quickly a chemical reaction happens. It is important for people in industry to know how fast a reaction goes. They have to know exactly how much they can make each hour, day, or week for their products. In a shampoo factory, the rate might be 100 bottles per minute.

The rate of reaction tells us how quickly a chemical reaction happens.

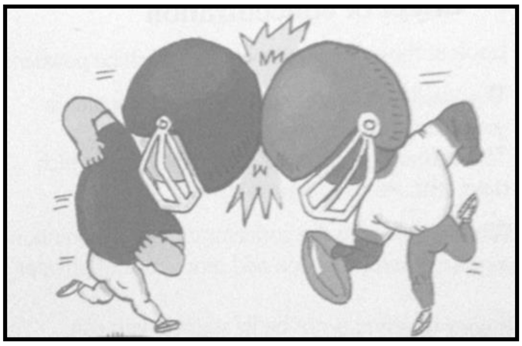
We cannot work out the rate of reaction from its chemical equation. Equations only tell us how product we can get. They do not say how quickly it is made. We can only find the rate by actually doing experiments.

During a reaction, we can measure how much reactant is used up in a certain time. On the other hand, we might choose to measure how much product is formed in a certain time.

Different factors affect the rate of reactions. In order to have a better understanding in each factor affecting the reaction rate, you need to know first the term collision theory.

What is collision theory?

All the factors that affect the speed of chemical reactions are explained in terms of the kinetic particle or collision theory.



Particles must collide before they can react.

Collision theory is a model that describes how the rate of a chemical reaction is determined by the collisions between reacting particles. This theory suggests that a chemical reaction only occurs when two reacting particles (atoms or molecules) collide with each other.

These results in old bonds being broken and new bonds formed between atoms in new molecules. If reacting particles do not collide or impact with enough energy, then the reaction will not proceed any further.

Collision theory is a model that describes how the rate of a chemical reaction is determined by the collisions between reacting particles. The more collisions between particles in a given time, the faster the reaction is.

For the particles to be bonded together, the force of the collision must be great enough to overcome the initial repulsive forces of the atoms. This minimum energy that is required to start the reaction is called the **'activation energy'** (the energy barrier that reacting particles must overcome).

Factors Affecting the Rate of Reaction

a. Temperature
b. Concentration of reactants
c. Surface area(particle size) of the reactants
d. Pressure of gas reactants
e. The presence of catalyst

Temperature

An increase in the temperature of a reacting mixture will increase the kinetic energy of reacting particles, leading to more collisions with greater energy. This will result in a greater number of collisions achieving activation energy in a shorter space of time, and a faster reaction.

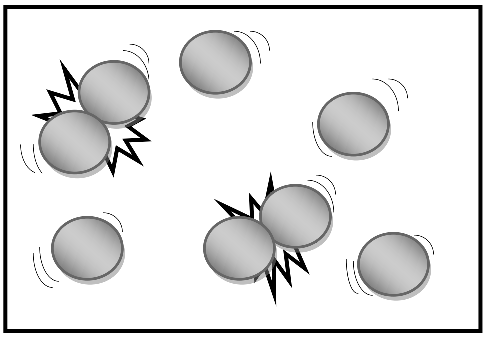
So, raising the temperature makes particles collide more often in a certain time and makes it more likely that collisions result in a reaction. As we increase the temperature, we increase the rate of reaction.

Increase temperature → **Increase kinetic energy** → **Increases collisions** → **Increases the rate of reactions**

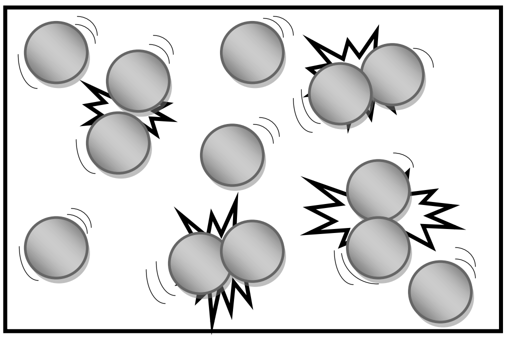
Lowering the temperature of a reacting mixture will reduce the energy of reacting particles, resulting in fewer collisions with less energy and therefore a slower reaction.

Decrease temperature → **Decrease kinetic energy** → **Decreases collisions** → **Decreases the rate of reactions**

The diagram below shows that at higher temperature there will be more colliding particles.



Reaction at 30°C

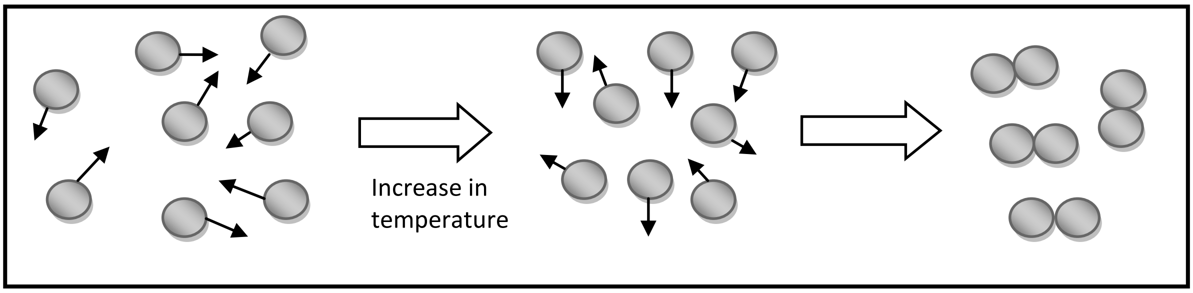


Reaction at 40°C

For example, to make a solute dissolve faster in water, you often heat it up. So, to make a piece of zinc metal react faster with dilute sulphuric acid, you will heat the acid with zinc. At home, your mother would use a stronger fire or even a pressure cooker to cook food faster. The temperature in a pressure cooker can reach 120°C, instead of 100°C under normal atmospheric pressure.

Temperature is very important in influencing the speed of reactions. In general, the speed of a reaction is doubled for every 10°C rise in temperature. For example, if a certain reaction is completed in 30 minutes; at 50°C, it would take only 7.5 minutes, and so on.

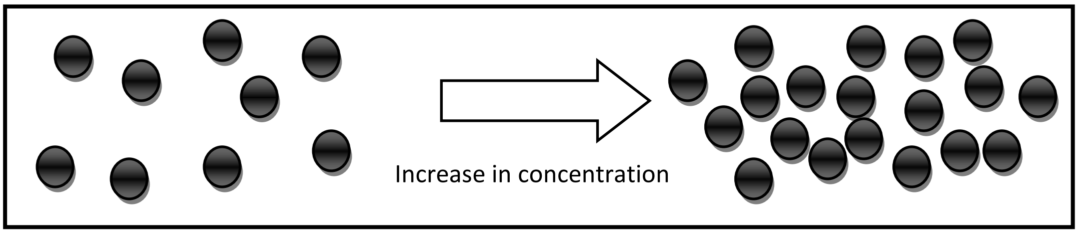
In general, when the particles of two reactants are moving at high speeds, the chances of them colliding with one another are higher, so there is an increase in the number of collisions resulting in the formation of the products.



Particles move at a lower speed Particles move at a higher speed More successful collisions

Concentration
Increasing the concentration of reacting particles will cause a greater number of collisions. With these extra collisions, more reacting particles will achieve activation energy and the reaction will be faster. Lowering the concentration will result in fewer collisions and the reaction will be slower.

When reactions involve gases, increasing the pressure inside the reacting chamber will have the same effect as increasing the concentration.



Reactant particles are less crowded Reactant particles are more crowded

Increase concentration → **Increase number of collisions** → **Increases the rate of reactions**

Decrease concentration →	**Decrease number of collisions** →	**Decreases the rate of reactions**

In general, an increase in concentration of reactants will increase the speed of reaction because as the concentration of the reactant(s) increases, more reacting particles are being crowded into the same volume. This increases the number of collisions between the reacting particles, and so, the speed of the reaction increases.

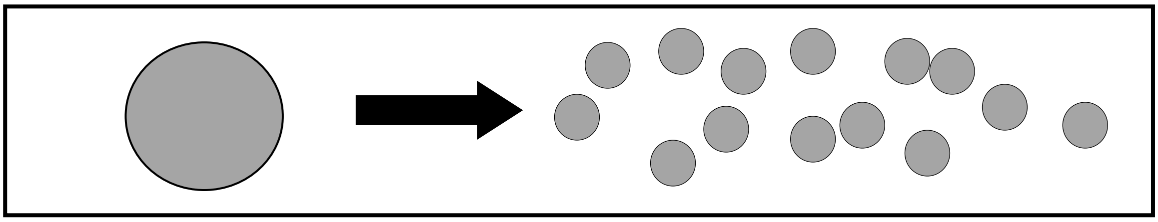
As we increase the concentration, we increase the rate of reaction.

Surface area

We all know that meat and vegetables can be cooked more quickly by cutting them into small pieces. This is because the smaller the size of the particles, the faster the reaction.
By increasing the surface area of any solids involved in a chemical reaction, you can increase the speed of the reaction.

Increase surface area →	**Increase number of collisions** →	**Increases the rate of reactions**

When a solid substance is broken into smaller pieces it provides a larger surface area over which collisions can occur. This results in more collisions between the reacting particles, more particles with sufficient energy and a faster reaction.



A large particle has smaller surface area   Smaller particles have a larger total surface area

As mentioned earlier, the speed of reaction increase with decreasing surface area (increasing particle size) of a solid reactant and so the reaction is faster because there are more particles that are exposed to react.

However, if a big lump of solid reactant (decrease particle size) is used, then it will greatly affect the speed of reaction because there is only one particle size that is allowed to react with other reactant and so, the reaction will become slower.

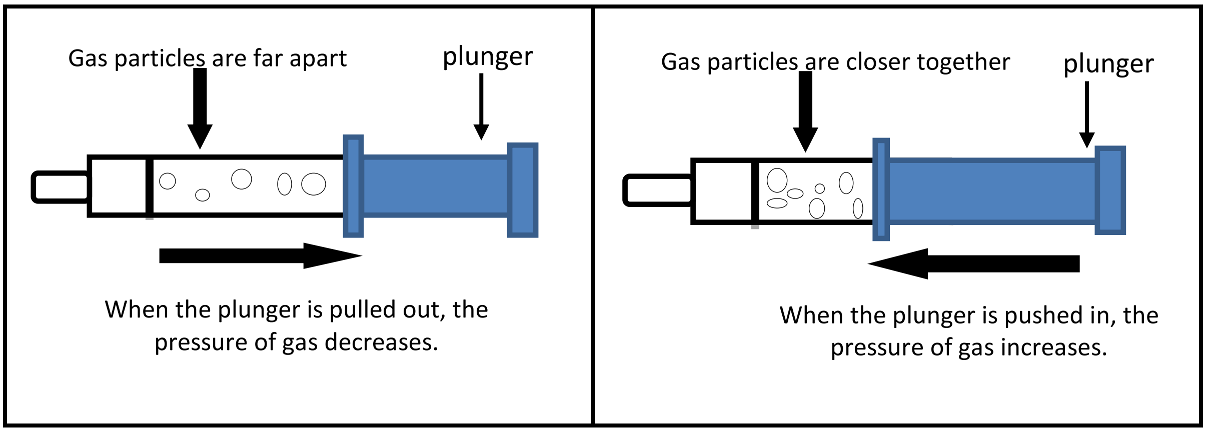
Decrease surface area	➡	**Decrease number of collisions**	➡	**Decreases the rate of reactions**

As we increase the surface area, we increase the rate of reaction.

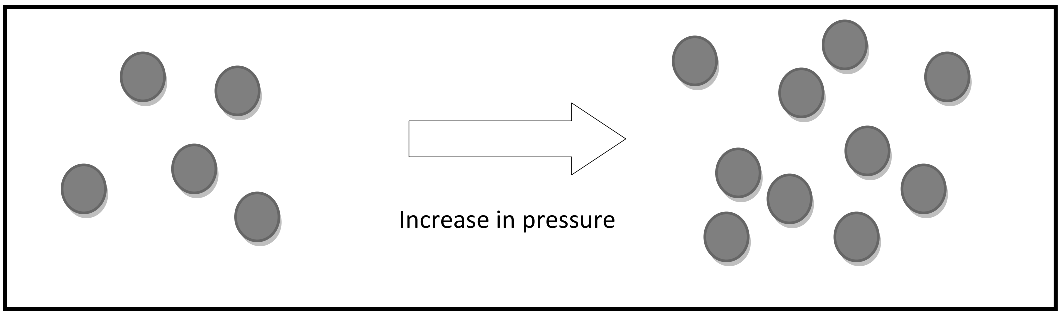
Pressure

Pressure has very little effect on the speeds of reactions in solids and liquids. When reactions involve gases, increasing the pressure inside the reacting chamber will have the same effect as increasing the concentration. Lowering the pressure will slow the reaction due to fewer collisions.

Look at the syringe below: If the end is sealed, how can you increase the pressure of the gases inside the syringe? By pressing the plunger in, you now have the same number of gas particles in a smaller volume. In other words, you have increased the concentration of the gas.



There are more collisions in a given time when you increased the pressure of gases.



Reactant particles are less crowded. Reactant particles are more crowded.

In reactions between gases, increasing the pressure increases the rate of reaction.

Catalyst

A catalyst is a substance which increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.

Characteristics of a catalyst

- A catalyst lowers the activation energy (the energy needed to start up a reaction).
- Only a small amount of catalyst is needed to speed up the reaction.
- A catalyst is selective in its action. This means that one catalyst cannot act on or speed up all types of reactions. Different catalyst speed up different reactions.
- A catalyst is not used up during the reaction. The same amount of catalyst is present at the beginning and at the end of the reaction.
- Impurities can prevent catalysts from working. We say that the catalyst is poisoned or inactivated by impurities.
- The physical appearance of catalyst may change at the end of the reaction, but its chemical properties remain unchanged.
- A catalyst increases the speed and not the yield of a chemical reaction. The same amount of products is formed whether a catalyst is used or not.

The energy needed to start up a reaction is called the activation energy.

A catalyst lowers the activation energy.

Ways of Measuring the Rate of Reaction

The speed of reaction can be determined in three ways:

- measuring the time taken for a reaction to complete,
- measuring the amount of product formed against time and
- measuring the amount of reactant used up or remaining against time.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 1 **30 minutes**

Answer the following questions:

1. Define:
   (i) Rate ____
   (ii) Catalyst ____

---

2. Below are some examples of rate of reaction. Identify each example as SLOW and FAST reaction.

   (i) burning of wood __________
   (ii) rusting of iron nail __________
   (iii) ripening of tomatoes __________
   (iv) silver nitrate and sodium chloride reaction __________
   (v) fireworks exploding __________
   (vi) a newly painted wall __________
   (vii) compost decomposing __________

---

3. What does collision theory say?
   ____
   ____
   ____

---

4. List the factors affecting the rate of reaction.

   (i) ____
   (ii) ____
   (iii) ____
   (iv) ____
   (v) ____

---

5. The minimum energy to start up a reaction is called __________ energy.

6. The catalyst lowers the __________ energy and the reaction will become faster.

7. Write the word INCREASES or DECREASES to describe the rate of reactions in the following situations:

(i) Increasing the temperature ___________
(ii) Increasing the pressure of gases ___________
(iii) Decreasing the concentration of a solution ___________
(iv) Using powdered metal in reaction with a solution ___________

Thank you for completing your learning activity 1. Check your work. Answers are at the end of this module.

Measuring the Rate or Speed of a Reaction

Measuring Speed of Reaction from Changes in Volume

Let us look at an example. The reaction between a reactive metal and a dilute acid is considered fast.

For example, magnesium reacts with dilute hydrochloric acid according to the equation:

magnesium	+	hydrochloric acid	$\rightarrow$	magnesium chloride	+	hydrogen gas
$Mg_{(s)}$	+	$2HCl_{(aq)}$	$\rightarrow$	$MgCl_{2(aq)}$	+	$H_{2(g)}$

As the reaction proceeds, the total volume of hydrogen gas produced increases. The speed of the reaction can therefore be determined by collecting and measuring the volume of hydrogen produced at regular intervals as shown in the experiment on the next page:

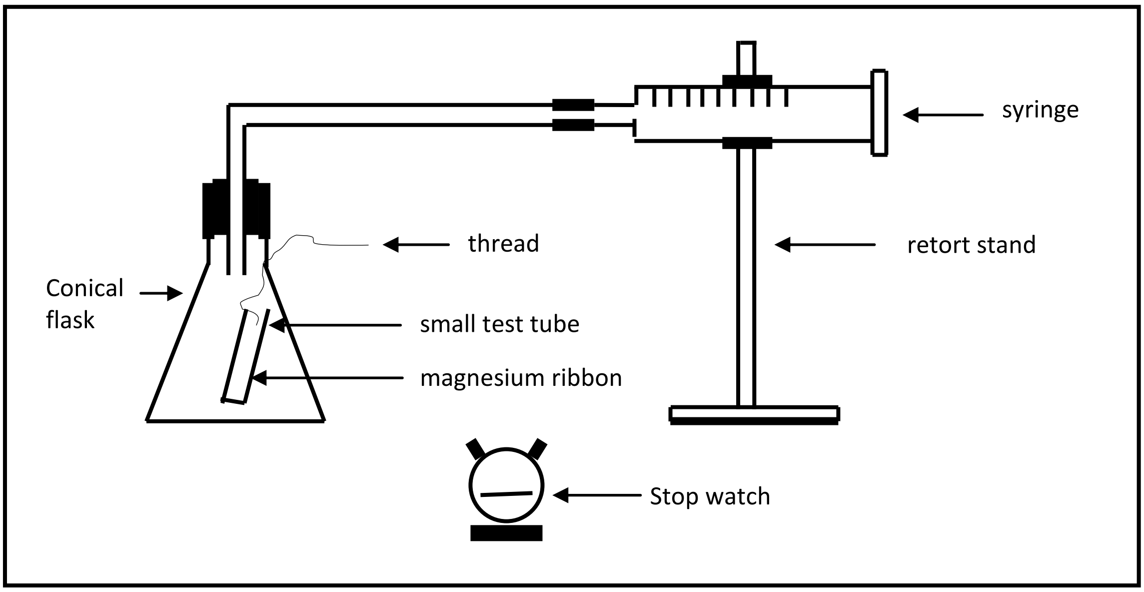
Experiment 1: Measuring the Volume of Gas Evolved

Aim:

To study the speed of reaction between dilute hydrochloric acid and magnesium.

Procedure:

1. The apparatus is set up as shown below. The layer of oxide on the magnesium ribbon is removed using a piece of sandpaper. This ensures that magnesium reacts with the dilute hydrochloric acid. The magnesium ribbon is then put in a small test tube.

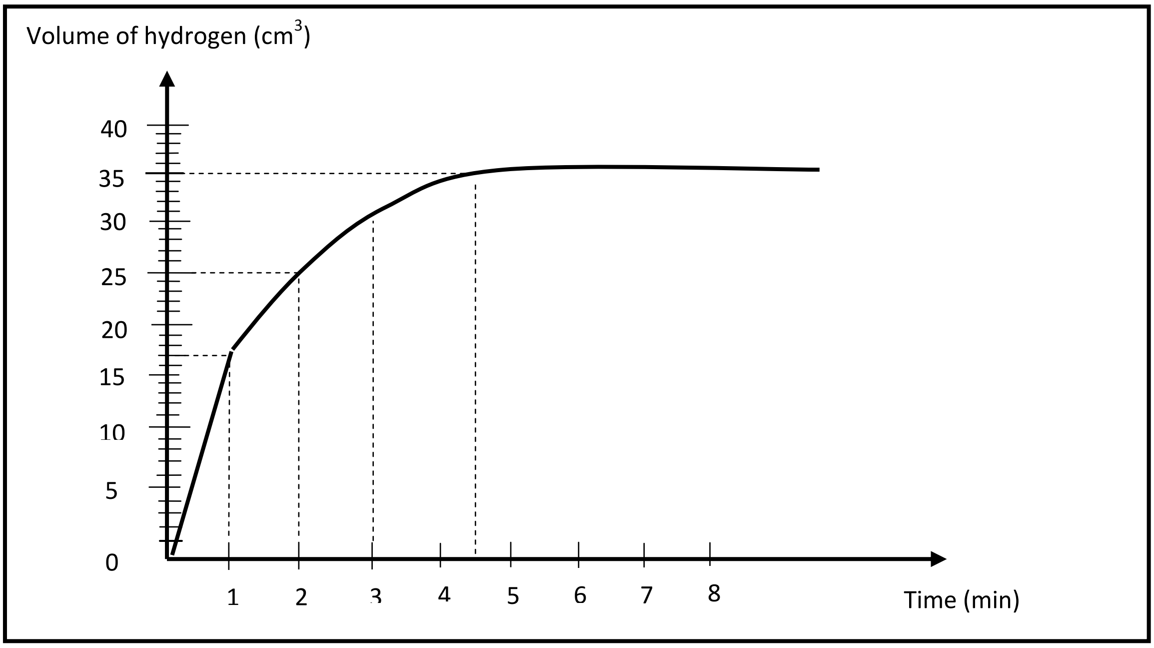


An experiment to study the speed of reaction by measuring the volume of gas released.

2. The conical flask is shaken to mix the magnesium ribbon and acid. The stopwatch is started at the same time.

3. The volume of hydrogen collected in the gas syringe is recorded every half minute. From the results of the experiment, a graph of the volume of hydrogen produced is plotted against time as shown on the next page.

A GRAPH SHOWING THE VOLUME OF HYDROGEN PRODUCED AT DIFFERENT TIME INTERVALS



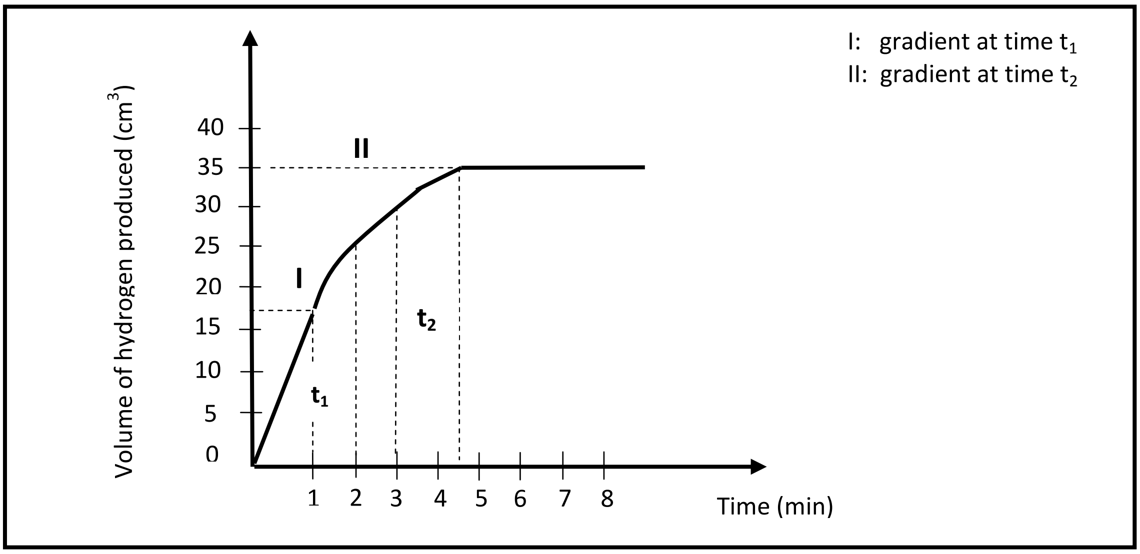
- In the first minute, about $17cm^3$ of hydrogen is produced.
- Between the first and second minute, only about $8cm^3$ (or $25cm^3$–$17cm^3$) of hydrogen gas is produced.
- At the end of 4.5 minute, the reaction stops. In total 35
- cm^3 of hydrogen is produced.

How can we find the speed of reaction at a particular time from a graph?

Generally, for any reaction, if the change in the mass or volume of the reactants or products is plotted against time, the speed of the reaction is indicated by the gradient of the graph.

In the graph below, the gradient at t_1 is the speed of reaction at that time.

THE GRADIENT OF THE GRAPH AT DIFFERENT TIMES



How can we estimate the change of speed of reaction from a graph?
The shape of the graph tells us whether the speed of reaction changes or remains the same as time passes. The steeper the gradient, the faster the speed of reaction is. For example, in the graph above shows that the gradient at t_2 (gradient II) is less steep than that at t_1 (gradient I). As the reaction proceeds, the curve becomes less steep. This means that the speed of reaction is decreasing.

Measuring speed of reaction from changes in mass
The speed of reaction can also be found by measuring the changes in mass of a reaction mixture. This method works best for reactions which produce gases such as carbon dioxide. For example, the speed of reaction between calcium carbonate and hydrochloric acid can be studied this way as shown in the next experiment:

The steeper the gradient is, the faster the speed of reaction.

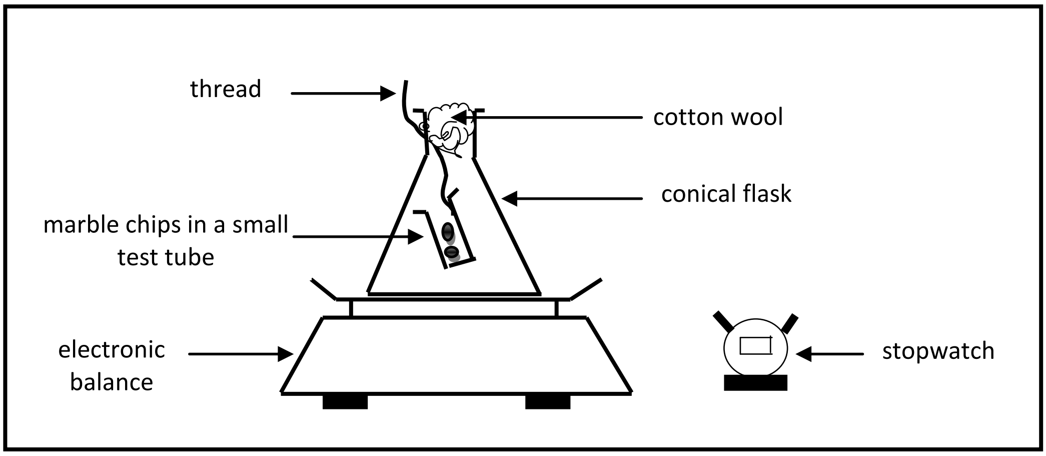
Experiment 2: Measuring the Mass at Different Time Intervals

Aim:

To study the speed of reaction between calcium carbonate and dilute hydrochloric acid

Procedure:

1. The apparatus is set up as shown below. The cotton wool in the mouth of the conical flask is used to prevent acid sprays (to stop acid from splashing out as the reaction takes place).

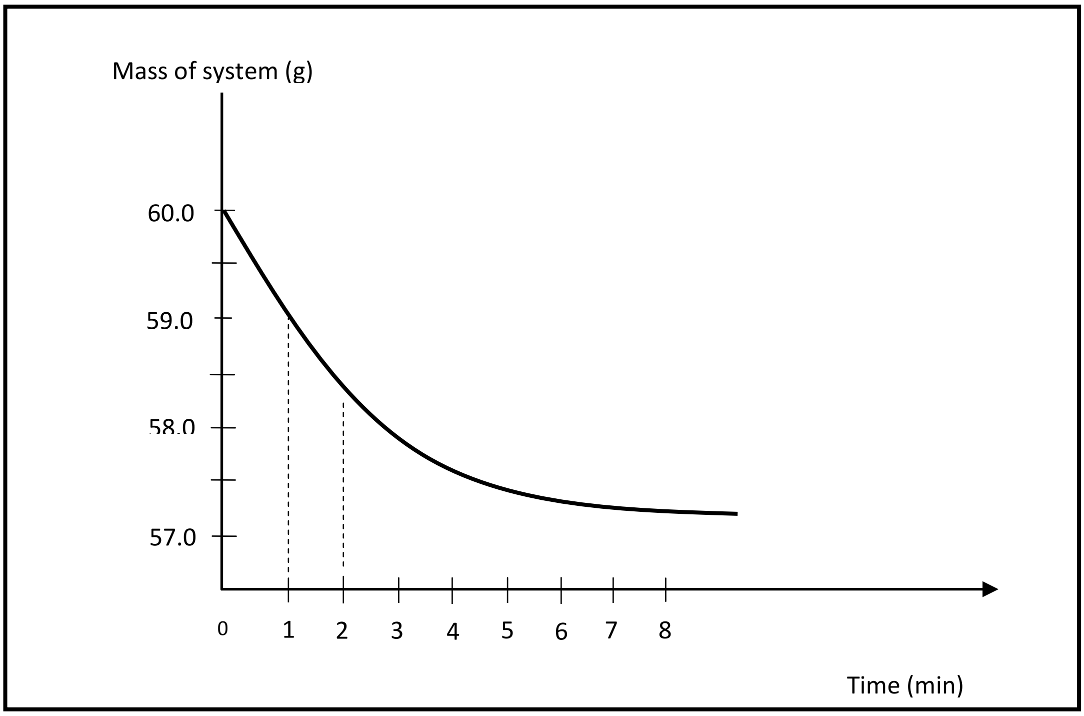


An experiment to study the speed of reaction by measuring the mass at different intervals.

2. The mass of the system is recorded. This includes the mass of the marble chips (calcium carbonate), dilute hydrochloric acid, conical flask, small test tube, string, and cotton wool.
3. The conical flask is shaken to mix the marble chips and acid. The stopwatch is immediately started.
4. The mass of the system is recorded at 1 minute intervals.

A sample set of results from experiment 2 is used to plot a graph as shown below:

A GRAPH SHOWING THE MASS OF THE SYSTEM AT DIFFERENT TIME



As you can see, there was an increase in mass. What happens to the mass of marble chips as shown in the graph?

- During the first minute ,
  decrease in mass = 60.0 g - 59.0 g
  = 1.0 g

  Average speed of reaction in the first minute = decrease in mass/time taken
  = 1.0/1
  = 1.0 g/min

- During the second minute,
  decrease in mass = 59.0 g - 58.4 g
  = 0.6 g

  Average speed of reaction in the second minute = decrease in mass/time taken
  = 0.6/1
  = 0.6 g/min

- The gradient decreases with time. The reaction slows down as it proceeds. The reaction stops after 7.5 minutes.

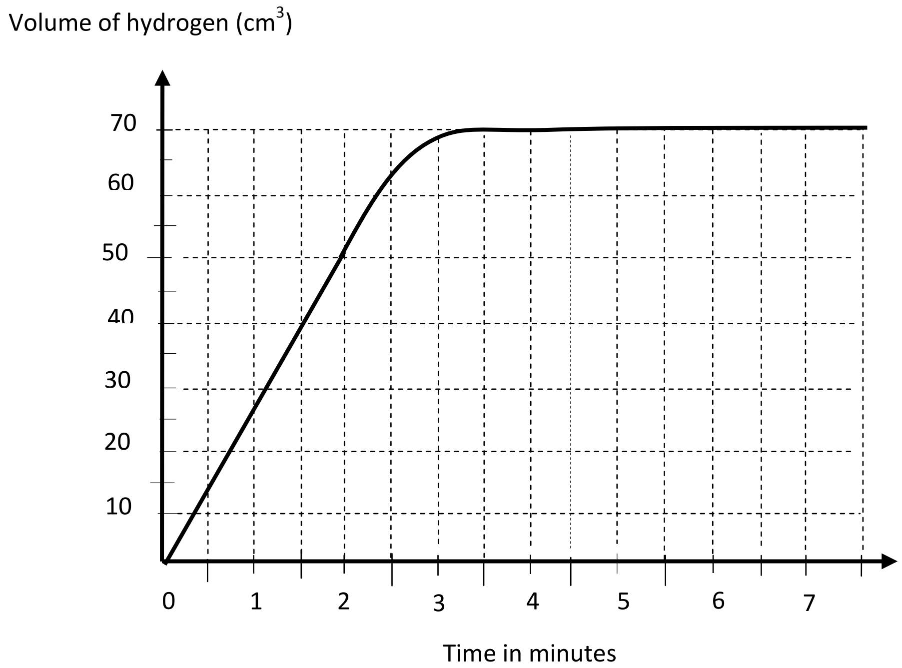
Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 2 **30 minutes**

Answer the following questions:

Refer to the graph showing the reaction between magnesium and dilute hydrochloric acid to answer **Questions 1 to 4**.

A GRAPH SHOWING THE REACTION BETWEEN MAGNESIUM AND DILUTE HYDROCHLORIC ACID



1. a. What is the volume of hydrogen produced in the first minute? __________
   b. What is the rate in the first minute? __________
   c. What is the volume of hydrogen gas produced in the second minute? __________
   d. What is the rate in the second minute? __________

---

2. a. From the graph given, how can you tell that the reaction is over? __________
   b. How much hydrogen is produced in 3.5 minutes? __________

---

3. a. In what time does the reaction is fastest? __________
   b. In what time does the reaction is slowing down? __________

4. a. What is the average rate of reaction above? __________
   b. The faster the reaction, the __________ the curve.

Thank you for completing your learning activity 2. Check your work. Answers are at the end of this module.

The Effect of Concentration

When a reaction involves solutions, we can alter the speed of reaction by changing the concentration of the reactant. To study the effect of concentration on the speed of reaction, let us look at the reaction between magnesium ribbons and dilute hydrochloric acid.

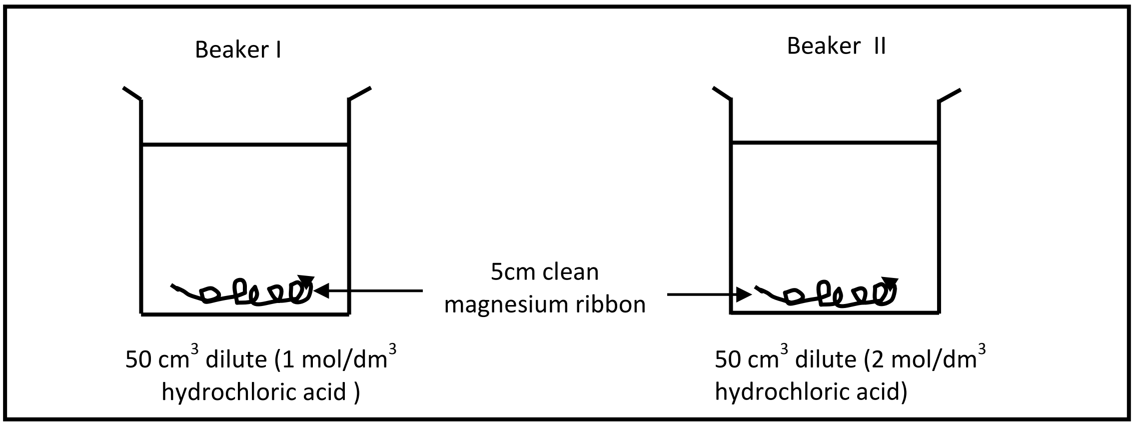
Experiment 3: The Effect of Concentration on the Speed of Reaction

Aim:

To study the effect of concentration on the speed of the reaction between magnesium and hydrochloric acid.

Procedure:

1. The apparatus is set up as shown below:



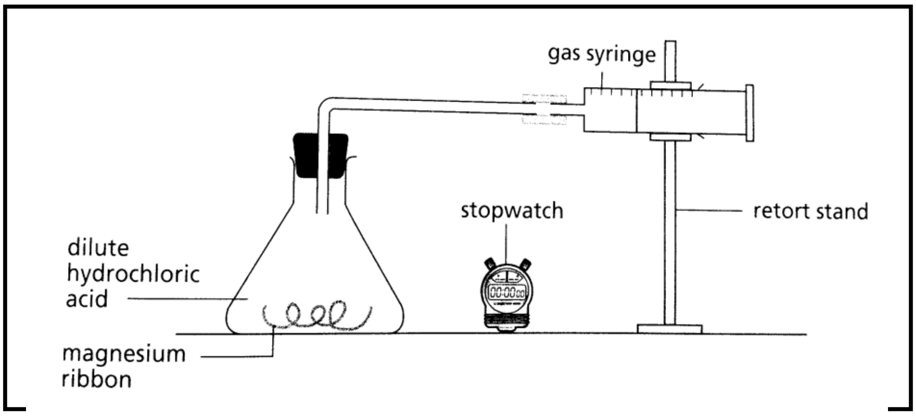
An experiment to study the effect of concentration on the speed of reaction.

2. The time taken for each piece of magnesium ribbon to dissolve completely is recorded. The results of this experiment are shown in Table 1.

Beaker	I	II
Time taken for magnesium to dissolve (g)	The reaction took 39 seconds to complete.	The reaction took 38 seconds to complete.

Results of Experiment 3

In the above experiment, the acid in beaker II is twice as concentrated as the acid in beaker II. The time taken for magnesium to react completely in beaker II is shorter. This means that magnesium reacts faster in more concentrated acid. Thus, we can conclude that a reaction proceeds faster when the concentration of a reactant is increased.
We can study the reaction between magnesium and hydrochloric acid in more detail using the apparatus shown below:



An experiment to study the effect of concentration on the speed of reaction by measuring the volume of gas released.

The experimental details for the reaction between magnesium and hydrochloric acid are given in Table 2.

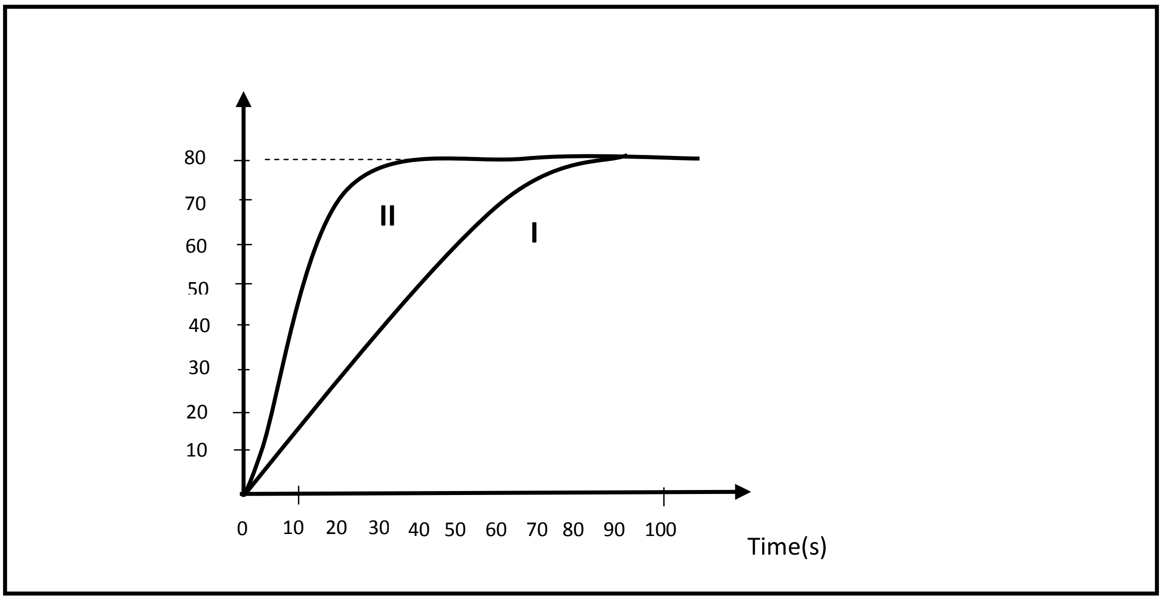
Experiment	**I**	**II**
Volume of hydrochloric acid (cm^3)	50.0	50.0
Concentration of hydrochloric acid (mol/dm^3)	1.0	2.0
Mass of magnesium (g)	0.1	0.1

Experimental conditions to investigate the effect of concentration on the speed of reaction.

For both investigations, the volume of hydrogen produced is recorded at regular time intervals.

The results of the two investigations are plotted on the same axes, with syringe readings on the y – axis and the time taken in the x- axis as shown on the graph next page.

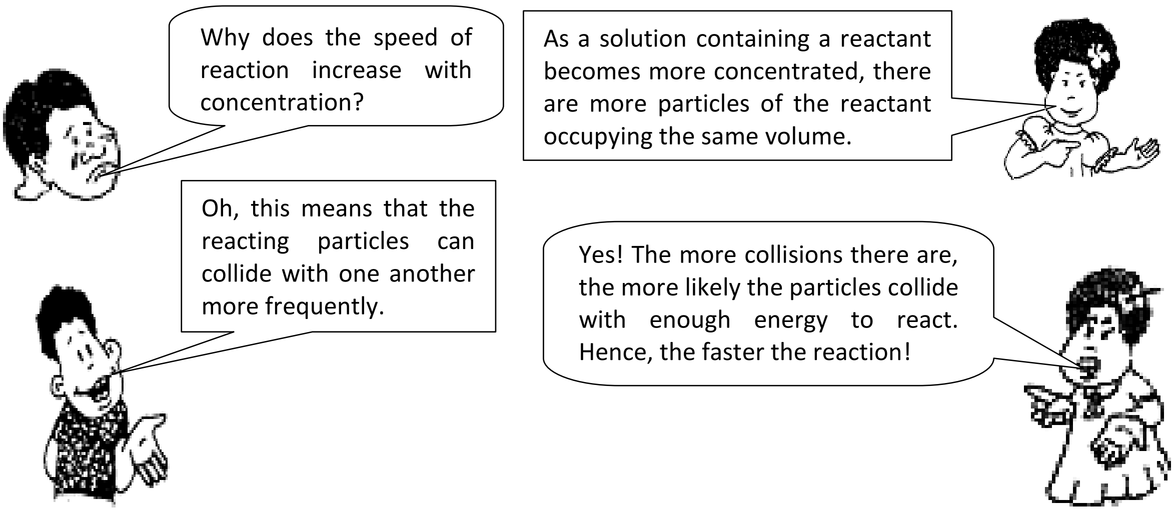
A GRAPH SHOWING THE EFFECT OF CONCENTRATION IN THE SPEED OF REACTION



1. Since the same mass of magnesium is used in each experiment, the same volume of hydrogen is produced.

2. Plot II is steeper than plot I at the start of the investigations. This means that the reaction is faster in investigation II. Since the concentration of the acid is higher in investigation II, it can be concluded that the speed of reaction increases when the concentration of the reactant increases.

3. In both investigations, the reaction slows down and the graph becomes less steep over time. This is because the concentration of hydrochloric acid and the amount of magnesium decrease as the reaction proceeds.

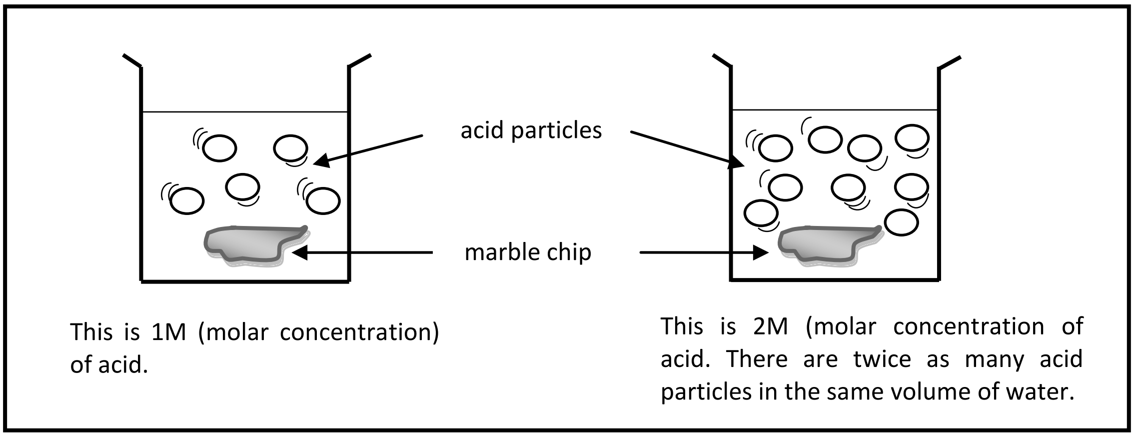


Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 3 **40 minutes**

Answer the following questions:

1. Look at the diagram to explain the effect of concentration in a reaction and answer the following questions below:



The effect of concentration between an acid and marble chip.

a. The marble chip in the diagram is also called calcium carbonate. What is the chemical formula of marble chip? __________

b. Which one is more concentrated, a 1M or 2M solution? __________

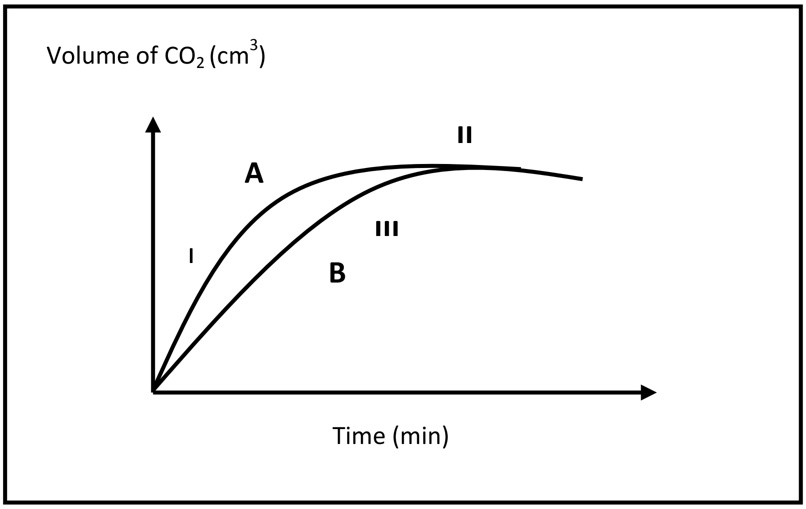
c. The acid in the beaker is a dilute hydrochloric acid. Write a balanced equation for its reaction with marble chip.

____

d. Which beaker has faster reaction, the one with a concentration of 1M or the beaker with a concentration of 2M and why?

____
____
____
____

2. A graph below is a sketch of curve on the use of two different concentrations, A and B in two different experiments.



A graph showing the use of two different concentrations in a reaction.

a. Which number represents a reaction is over? ___________
b. Which number represents the reaction is fastest? ___________
c. Which letter represents a less concentrated solution? ___________
d. Which number represents a reaction that slows down? ___________
e. Which letter represents the use of a more concentrated solution? ___________

Thank you for completing your learning activity 3. Check your work. Answers are at the end of this module.

The Effect of Temperature

Milk will turn sour very quickly if it is exposed to the air at room temperature, but it will keep fresh for several days in a refrigerator. The souring of milk is a decomposition reaction caused by bacteria.

A chemical reaction can be made to proceed faster or slower by increasing or decreasing the temperature of the reactants.

For example, when dilute hydrochloric acid is added to sodium thiosulphate solution, a fine precipitate of sulphur slowly forms and the solution becomes cloudy. The reaction can be represented as:

Sodium thiosulphate + hydrogen chloride $\longrightarrow$ sodium chloride + water + sulphur dioxide + sulphur

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

The speed of this reaction can be calculated as the rate of sulphur precipitation.

In the next experiment below, the time is recorded for the same amount of sulphur to precipitate. The shorter the time taken for sulphur to precipitate, the higher the speed of reaction.

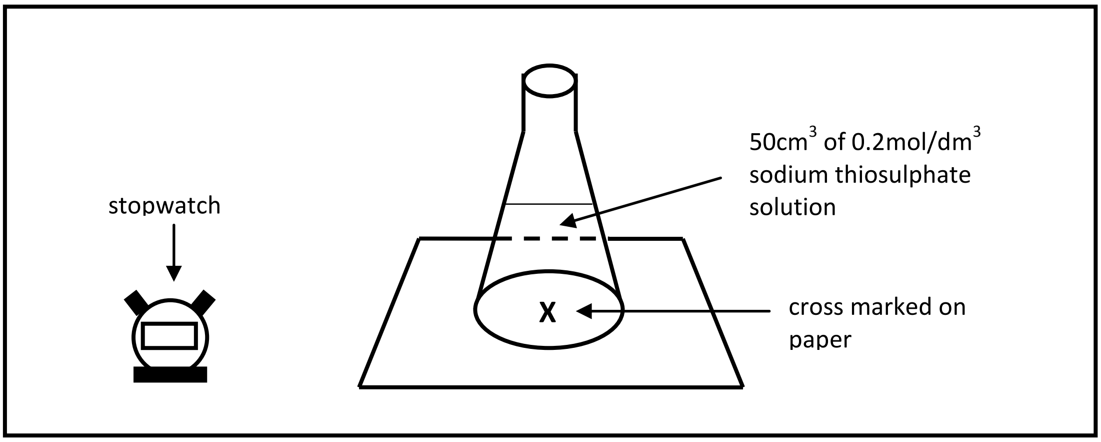
Experiment 4: The Effect of Temperature on the Speed of Reaction

Aim:

To study the effect of temperature on the speed of reaction between sodium thiosulphate solution and dilute hydrochloric acid.

Procedure:

1. The apparatus is set up as shown below:



An experiment to study the speed of reaction between hydrochloric acid and sodium thiosulphate.

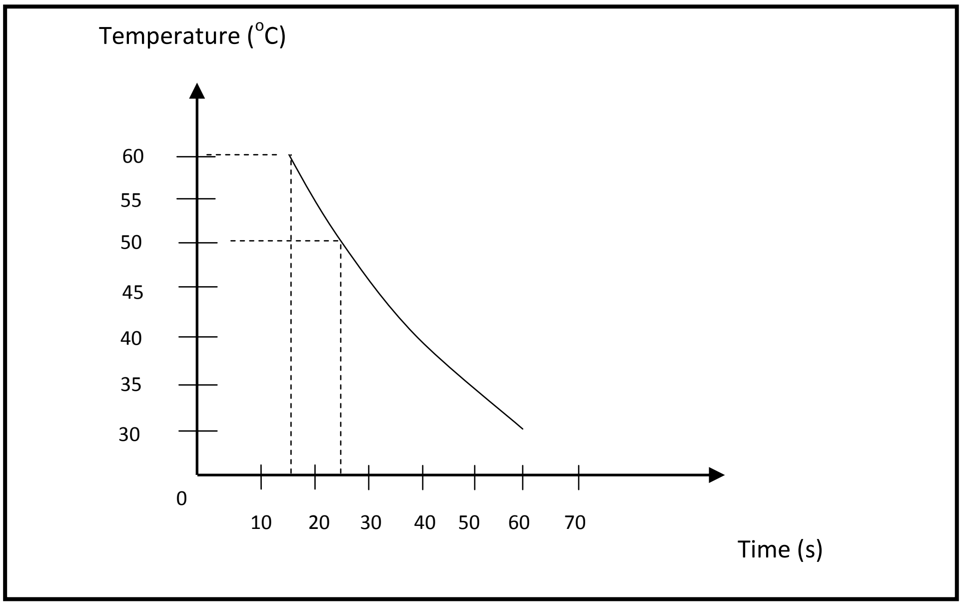
2. $50cm^3$ of dilute hydrochloric acid is then quickly poured into the sodium thiosulphate solution and the stopwatch is started immediately.

3. The mixture is swirled once and the stopwatch is stopped at the moment the cross disappears from view. The time taken is recorded.

4. The experiment is repeated three times using fresh solutions of the same reactants, but with sodium thiosulfate solution heated to higher temperatures.

The graph on the next page shows the results obtained in **experiment 4**. In this experiment, the same amount of sulphur is precipitated at each temperature. This means that the less time it takes for the cross to disappear from view, the faster the reaction is.

A GRAPH SHOWING THE TIME TAKEN FOR THE CROSS TO DISAPPEAR AT DIFFERENT TEMPERATURES

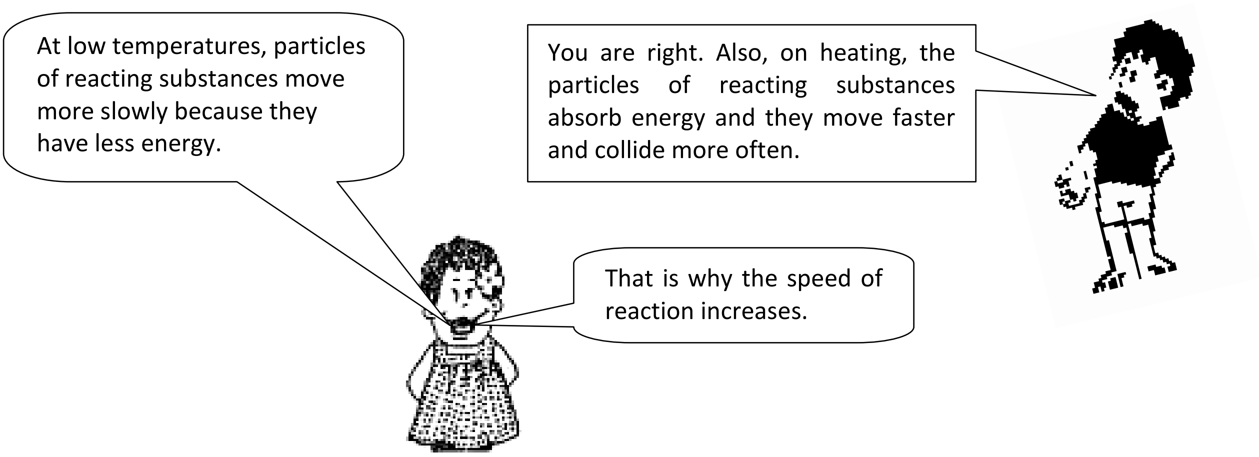


Look carefully at the graph above:

1. The higher the temperature, the shorter the time taken for the cross to disappear from view. This means that the higher the temperature, the faster the speed of reaction.

2. Observe the gradient of the graph at various temperatures. The speed of reaction increases rapidly as the temperature increases.

The graph shows clearly that the speed of reaction is faster at higher temperatures.

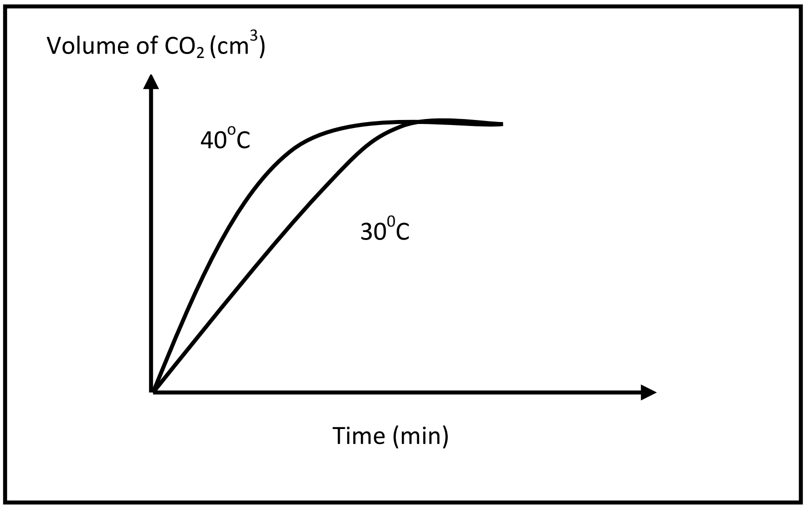


Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 4 **40 minutes**

Answer the following questions:

1. Refer to the graph as shown below to answer **Questions 1 and 2**.



A graph showing the use of different temperatures in a reaction.

1. a. Which reaction is faster, at 30°C or 40 °C? ___________
   b. How can you tell that the reaction is faster? ___________

---

2. a. Can you get more products when you do the same experiment higher than 40 °C?
   Give a reason for your answer.
   ____

   b. If you do the same experiment at 50°C, where will you put the curve in the above graph? Sketch a curve using the same graph above.

---

3. Fill in the blanks.

   a. At higher temperatures, the particles are moving ___________.
   b. Raising the temperatures make the particles ____________ more.
   c. Increasing the temperature will ___________ the rate of reaction.

Thank you for completing your learning activity 4. Check your work. Answers are at the end of this module.

The Effect of Surface Area

Many chemical reactions involve solids. Solids come in different sizes or we say they have different particle sizes. How is the speed of reaction affected by particle size? To find out, we can study the reaction between marble chips (calcium carbonate) and dilute hydrochloric acid in Experiment 5 below:

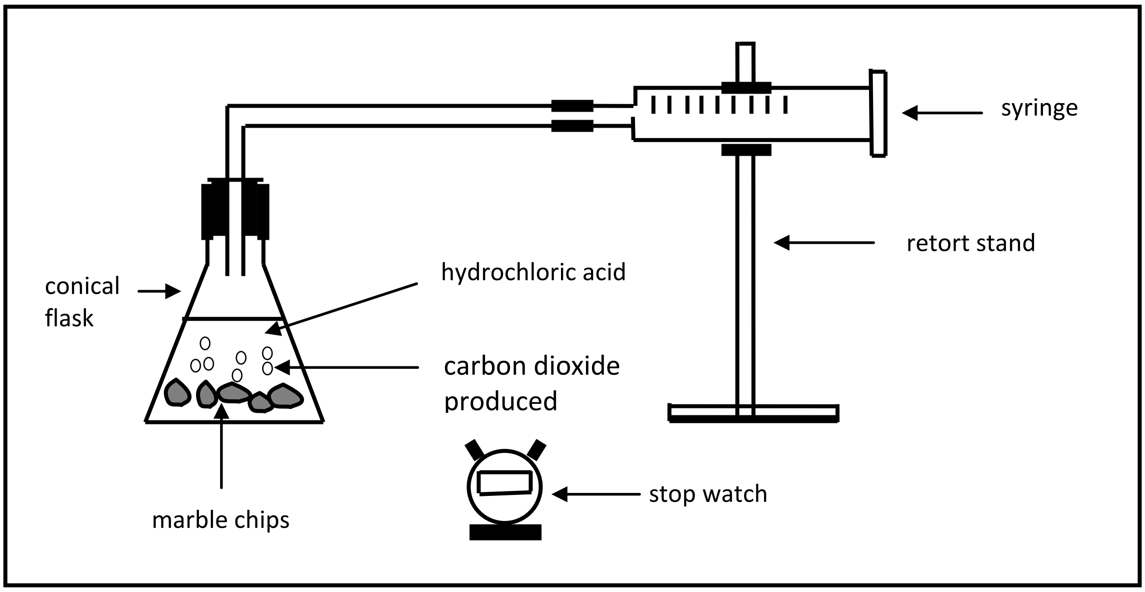
Experiment 5: The Effect of Particle Size on the Speed of Reaction

Aim:

To study the effect of particle size on the speed of reaction.

Procedure:

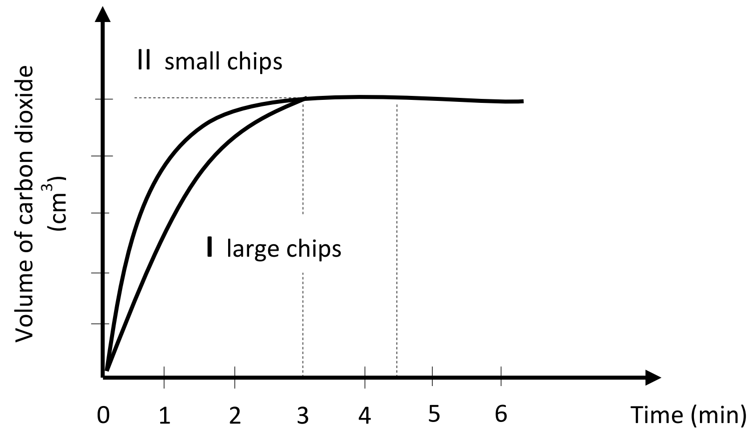
1. The experiment is set up as shown below. The volume of gas produced is recorded at 1 minute intervals for investigation I.



An experiment to study the effects of particle size on the speed of reaction.

2. The experiment is repeated for **investigation II**, with marble chips that have been crushed into much smaller pieces.

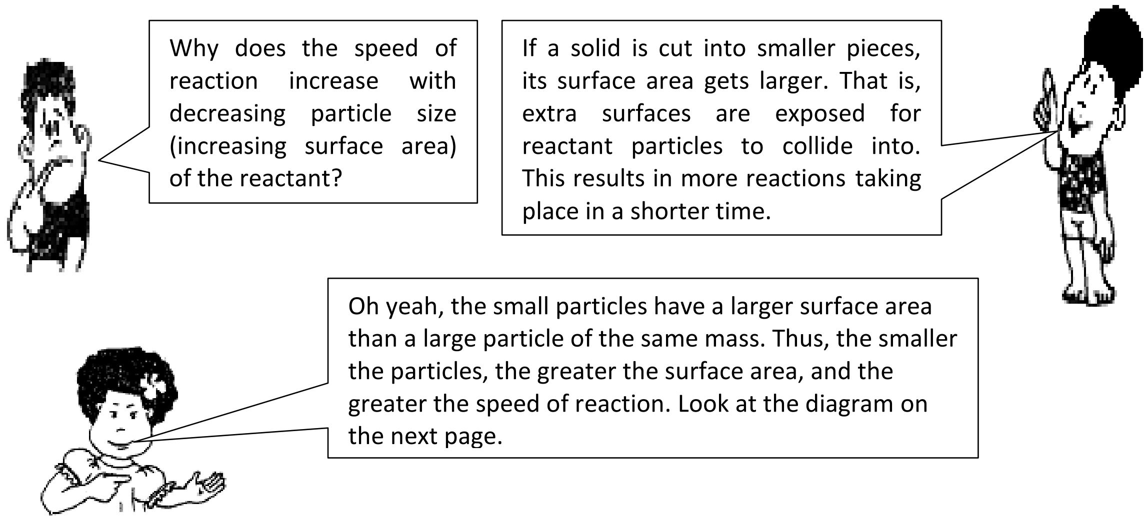
The results of the experiment are shown in the graph below:



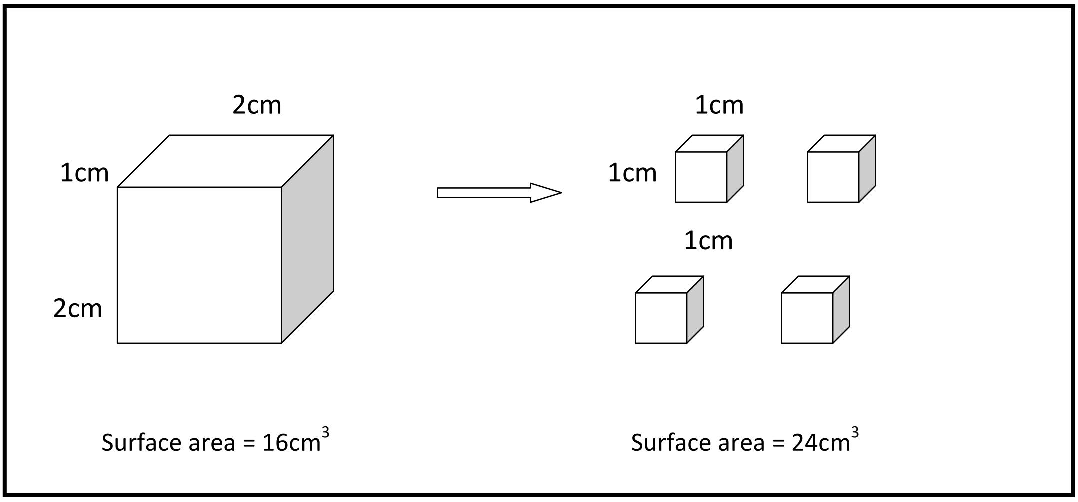
A graph showing the effect of surface area.

- Both graphs level off at the same value. This is because the same amounts of marble chips and hydrochloric acid are used for both the investigations.

- Graph II is steeper than graph I. This shows that the reaction is faster with smaller chips (smaller particle size). This is confirmed by the fact that with the smaller chips, the reaction is completed after 3 minutes. With the larger chips, the reaction stops only after 4.5 minutes.



The smaller the particles, the greater the surface area, and the greater the speed of reaction.



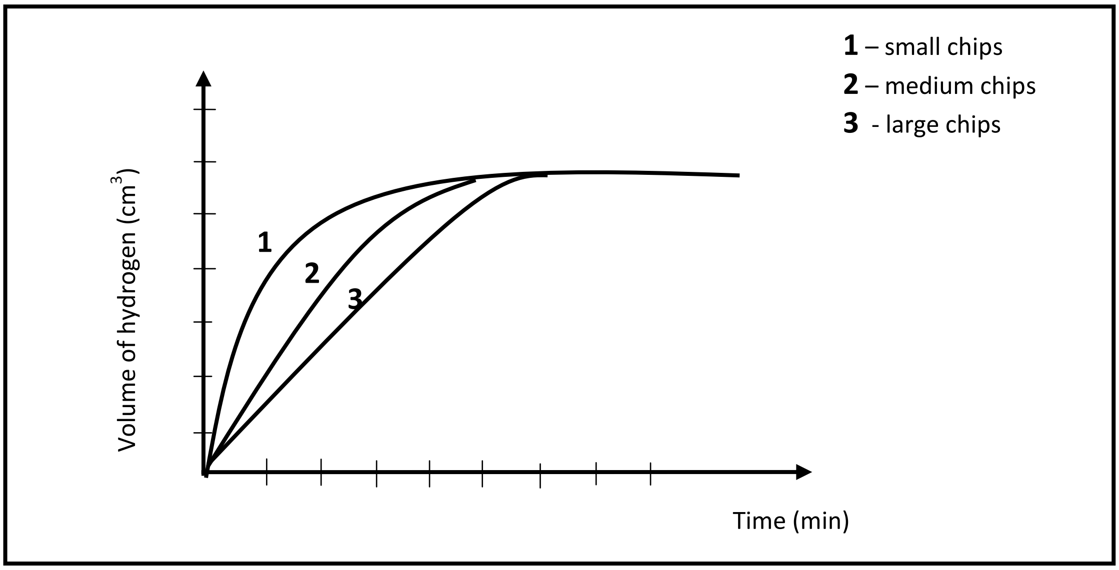
Smaller particles of a solid reactant have a larger surface area available for reaction to occur.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 5 **30 minutes**

Answer the following questions:

1. Three separate investigations were carried out by a group of students to find the reaction between magnesium (same masses) and hydrochloric acid (same concentration). The result of investigations is shown below:



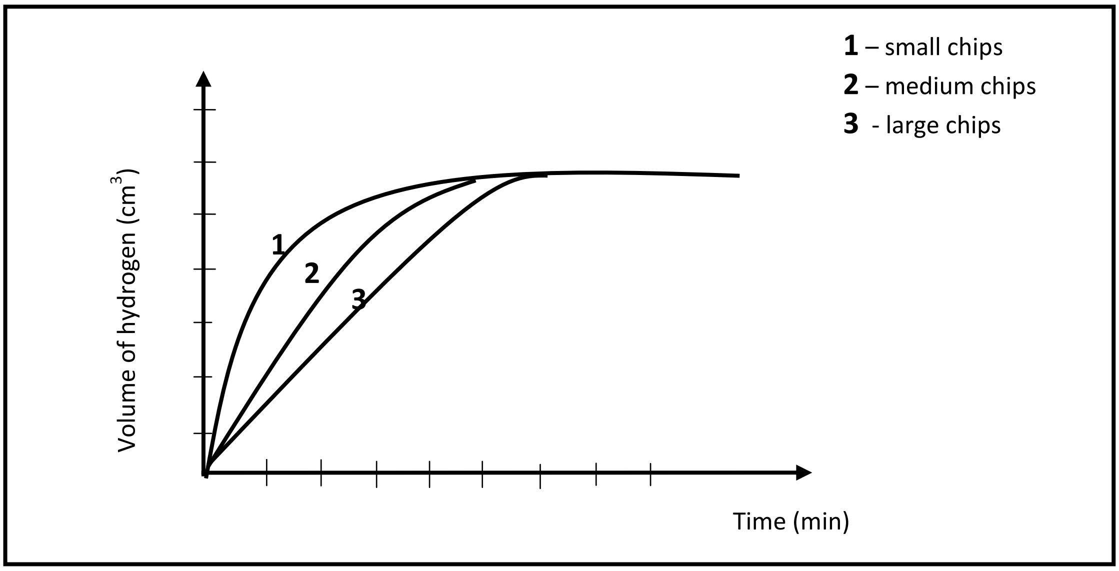
a. Which size of magnesium has the greatest surface area? ___________

b. Which size of magnesium reacted fastest? ___________
Explain why. ____

c. Which size of magnesium reacted the slowest? ___________
Explain why. ____

---

2. The students doing the experiments also tried reacting the same mass of powdered magnesium with the acid. Sketch your answer on the same graph below and label this number 4.



Thank you for completing your learning activity 5. Check your work. Answers are at the end of this module.

The Effect of Catalyst

How does the catalyst, manganese(IV) oxide, affect the decomposition of hydrogen peroxide?

Under normal conditions, hydrogen peroxide decomposes very slowly to give water and oxygen. Let us see what happens when a very small amount of manganese (IV) oxide is added to hydrogen peroxide.

$$\text{Hydrogen peroxide} \xrightarrow{\text{Manganese(IV)oxide}} \text{Water} + \text{Oxygen}$$

$$2H_2O_{2\,(aq)} \xrightarrow{\text{Manganese(IV)oxide}} 2H_2O_{(l)} + O_{2(g)}$$

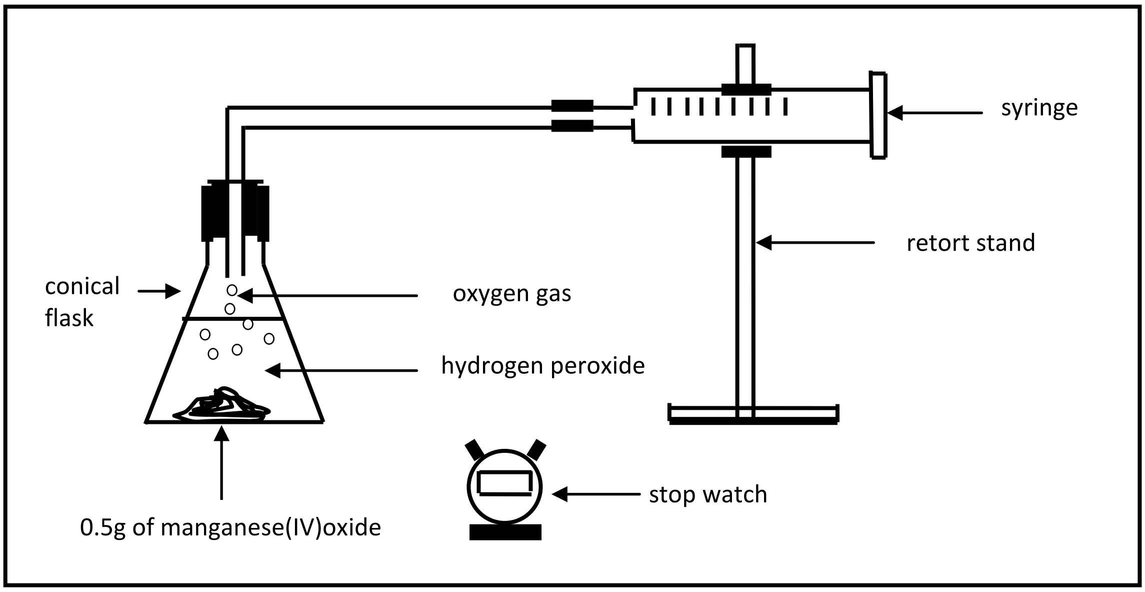
Experiment 6: The Effectiveness of Catalyst

Aim:

To study the effect of manganese(IV) oxide on the speed of decomposition of hydrogen peroxide.

Procedure:

1. A flask containing $50cm^3$ of hydrogen peroxide solution are set up as shown in the diagram below.



An experiment to study the effects of catalyst on the speed of reaction.

2. 0.5g of black manganese(IV) oxide (MnO_2) is added to the solution in the flask and the reaction takes place.
3. Observations are then recorded.

The following observations are:

(i) In the flask manganese(IV) oxide is added to hydrogen peroxide, bubbles of oxygen gas are quickly produced.

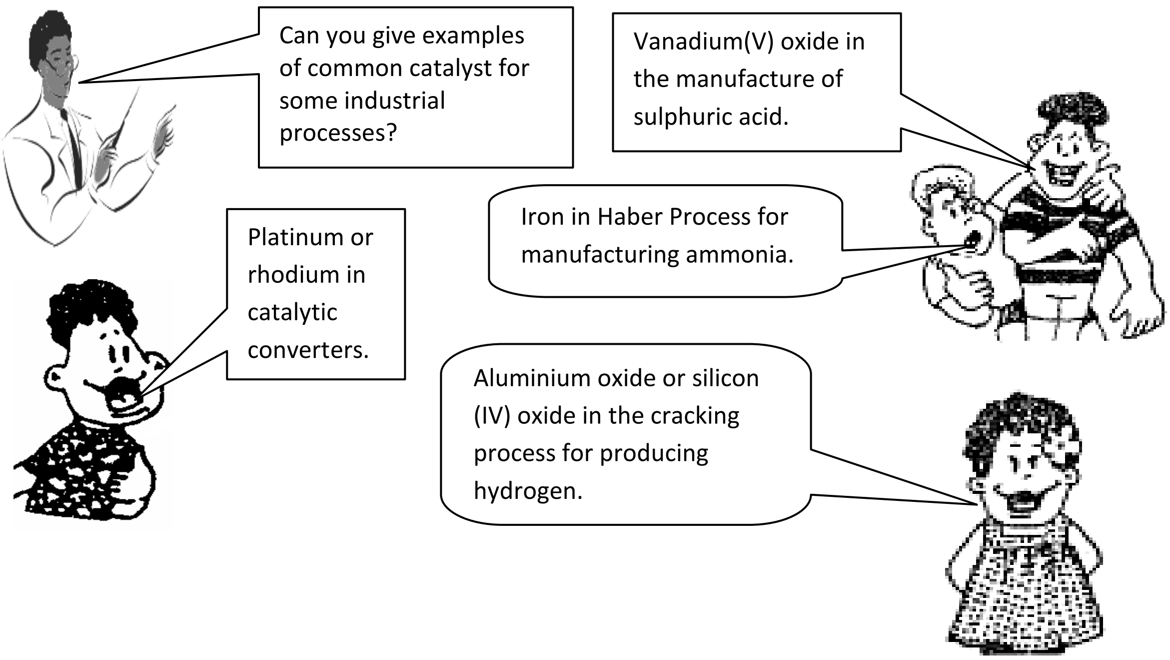
(ii) A small amount of oxygen gas is transferred to a test tube and a glowing splint is placed at the mouth of the test tube. The glowing splint burns brightly which means that the gas is really an oxygen gas.

(iii) When the reaction was over, it was found out that the mass of manganese(IV) oxide remained the same.

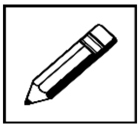
This experiment shows that manganese(IV) oxide speeds up the decomposition of hydrogen peroxide to produce oxygen gas. It is called a **catalyst** for the reaction. Manganese(IV) oxide is not the only catalyst that is used to speed up the decomposition of hydrogen peroxide.

Other suitable catalysts that can be used for the decomposition of hydrogen peroxide are copper(II) oxide and iron(III) hydroxide.

A catalyst is a substance which speeds up the rate of a chemical reaction. At the end of the reaction, the catalyst is chemically unchanged.



Now, check what you have just learnt by trying out the learning activity below!



Learning Activity 6

20 minutes

Answer the following questions:

1. List the ways of measuring the rate of reaction.
   a. ____________________
   b. ____________________
   c. ____________________

---

2. Hydrogen peroxide decomposes to form water and oxygen gas. The reaction is catalyzed by a transition metal oxide.
   a. What is the name and formula of the catalyst used in the reaction?
      Name ____________________

Formula ____

b. Write a balanced equation for the decomposition of hydrogen peroxide.

____

____

3. Name the two other catalysts that can be used in the decomposition of hydrogen peroxide apart from manganese(IV) oxide.

a. ____

b. ____

Thank you for completing your learning activity 6. Check your work. Answers are at the end of this module.

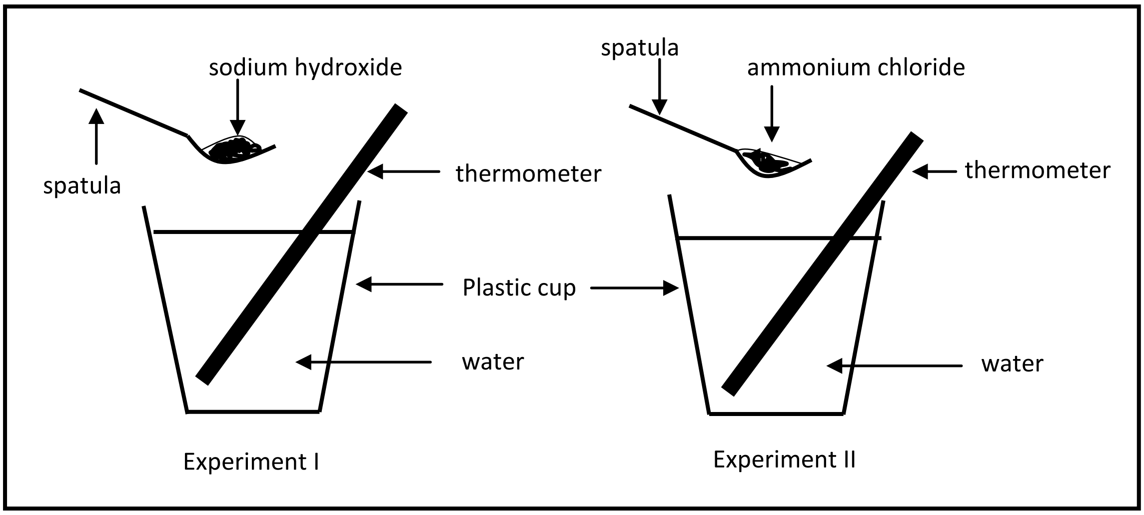
11.4.2 Energy Diagrams

Exothermic and Endothermic Reactions

Energy can be changed from one form to another. Energy changes occur in chemical reactions and even in some physical processes.

In the two experiments shown below, the temperature changes are measured when a solid is dissolved in water.

In experiment I, solid sodium hydroxide is added to water. The mixture is carefully stirred to dissolve the solids. The temperature of the water is recorded before and after adding sodium hydroxide. In experiment II, the experiment is repeated using ammonium chloride crystals.



Experiment I and II show the measurement of temperature change during a reaction.

The results of the experiments are recorded in table below.

Experiment	Solid	Initial temperature of water (^{o}C)	Temperature after adding solid (^{o}C)
I	Sodium hydroxide, $NaOH$	28	34
II	Ammonium chloride, NH_4Cl	28	22

Results of experiments I and II

What conclusions about energy changes can be made from these two experiments?
In experiment I, heat energy was given out when the solid dissolved in water. The temperature of the solution increases. The solution gets hotter. This is an exothermic reaction.

Experiment I shows the changes in temperature when an exothermic reaction occurs. Initially, the temperature of the reaction mixture rises until the highest temperature is reached. When the reaction is completed, the temperature of the reaction mixture falls until it reaches room temperature.

What are the characteristics of exothermic reactions?

- Heat is given out and is transferred from the chemicals to the surroundings.
- The temperature of the reaction mixture rises. The container feels hot.

Examples of exothermic reactions include:

- the combustion of fuels
- the rusting of iron
- the corrosion of metals
- the reaction between acid and an alkali (neutralization)
- respiration
- physical processes like condensation, freezing and dissolving of acids in water

Experiment II shows the changes in temperature when an endothermic reaction occurs. Initially, the temperature of the reaction mixture falls until the lowest temperature is reached. Heat energy is absorbed from the surrounding when the solid dissolved in water. The temperature of the solution decreases and gets colder. This is an endothermic reaction.
When the reaction is completed, the temperature of the reaction mixture rises until it reaches room temperature.

What are the characteristics of endothermic reactions?

- Heat energy is absorbed and is transferred from the surroundings to the reactants.
- The temperature of the reaction mixture falls. The container feels cold.

Examples of endothermic reactions include:

- photosynthesis
- the action of light on silver bromide in photographic film
- thermal decomposition
- physical processes like evaporation, melting, and the dissolving of some ionic compounds in water such as ammonium chloride, potassium nitrate and copper(II) sulphate crystals

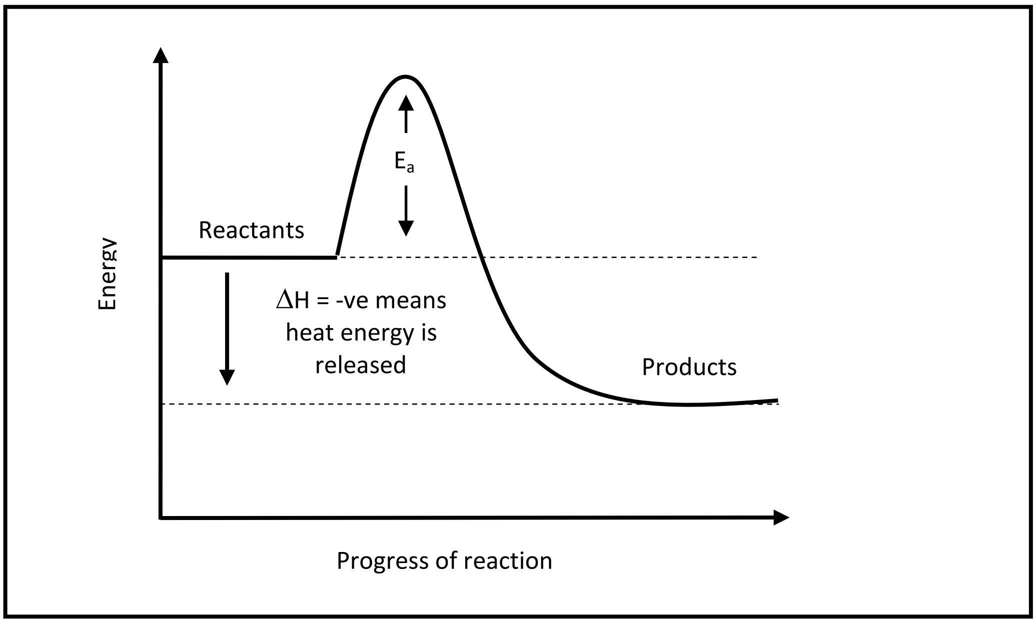
Heat of Reaction

In order for you to understand the meaning of heat of reaction, you need to study the basic energy level diagrams of exothermic and endothermic reactions.

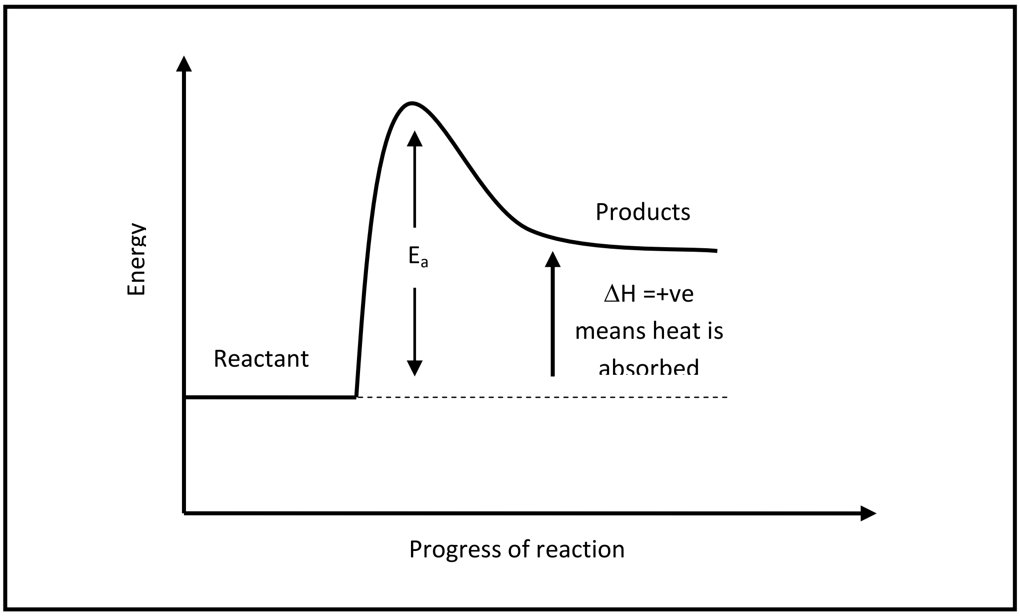
In the diagrams, the following parts are always seen such as:

- activation energy(E_a)
- heat of reaction (ΔH)

The **activation energy (E_a)** is the energy required to start up a reaction while the **heat of reaction** is the amount of heat energy released or absorbed during a chemical reaction. For exothermic reaction, the **heat of reaction (ΔH)** is **negative** and in endothermic reaction, it is **positive**. The energy level diagrams of exothermic and endothermic reactions are shown below.



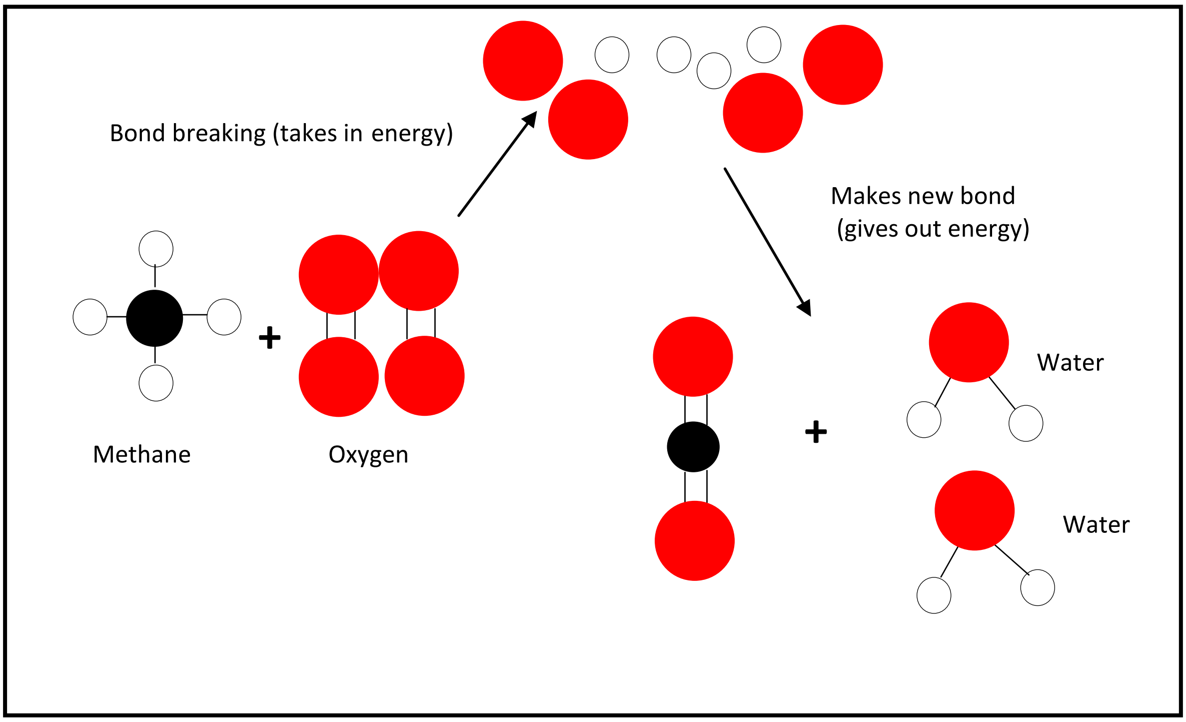
Energy level diagram for an exothermic reaction.



Energy level diagram for an endothermic reaction.

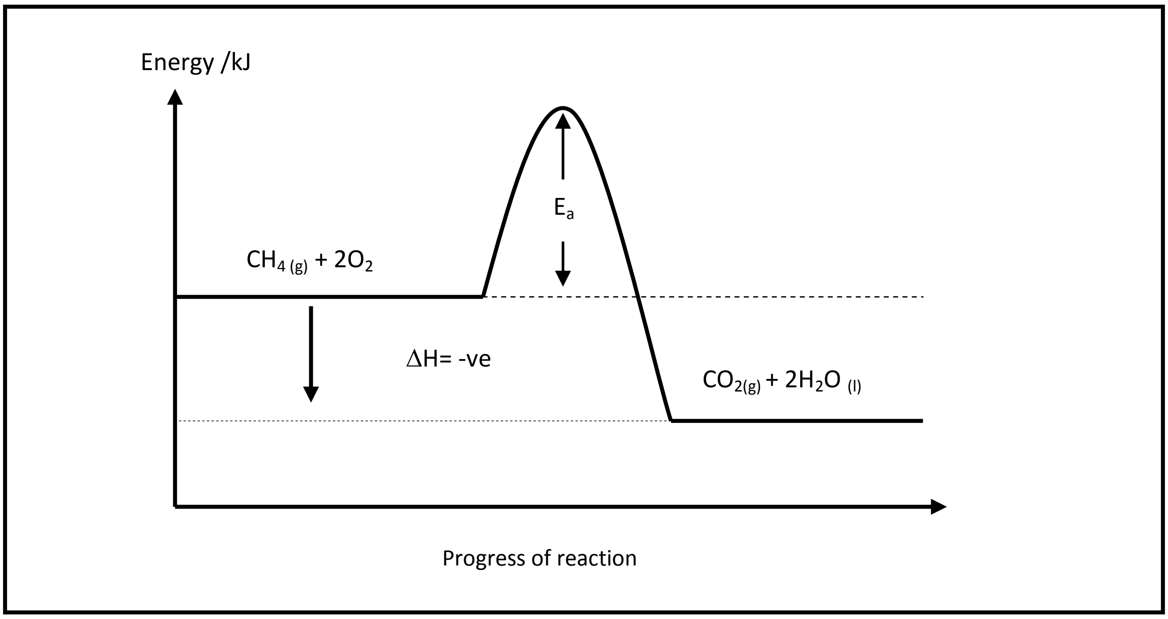
Chemical bonds are forces of attraction between the atoms, ions or molecules in a substance. To break these bonds, energy must be supplied. When bonds are created, energy is given out. In a chemical reaction, bonds are broken, and new bonds are made.

The diagram below shows the breaking and making of bonds when methane burns with oxygen forming carbon dioxide and water.



Breaking and forming of bonds during the combustion of methane.

The combustion of methane is an exothermic reaction as shown in the energy level diagram below:

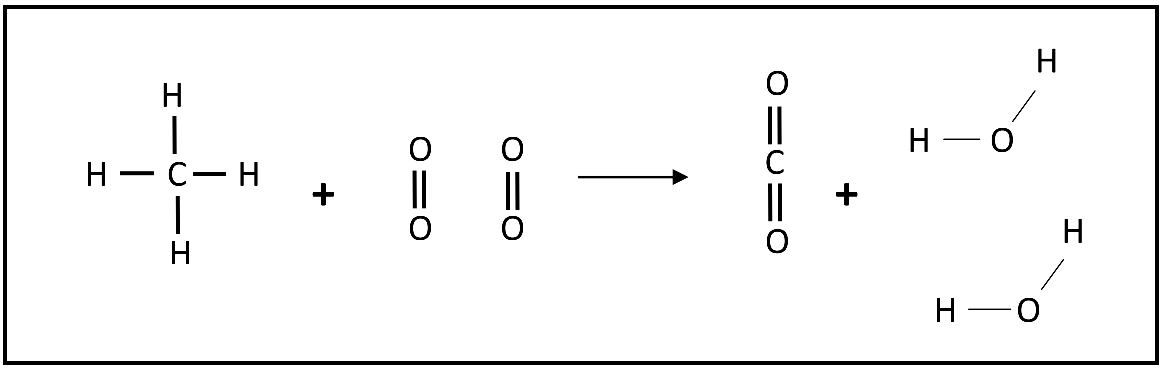


Energy level diagram for the combustion of methane.

The equation for the reaction is,

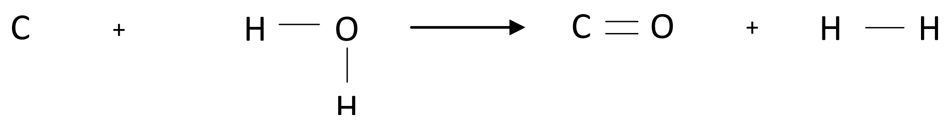
Methane	+	Oxygen	$\rightarrow$	carbon dioxide	+	Water
$CH_{4(g)}$	+	$2O_{2(g)}$	$\rightarrow$	$CO_{2(g)}$	+	$2H_2O_{(l)}$

The energy given out when the new bonds are made is greater than the energy taken in to break the old bonds. This reaction is **exothermic**.



Bonds broken and made when methane burns.

In exothermic reaction, the energy given out when the new bonds are made is greater than the energy taken in to break the old bonds.



The bonds in water must be broken. Energy must be taken in.

As the bonds in carbon monoxide and hydrogen gas are made, energy is given

The breaking and making of bonds when carbon (coke) reacts with steam to form carbon monoxide and hydrogen.

The equation for the reaction is,

$$C_{(s)} \quad + \quad H_2O_{(l)} \longrightarrow CO_{(g)} \quad + \quad H_{2(g)}$$

In this reaction the energy taken in to break the old bonds is greater than the energy given out when the new bonds form. The reaction is **endothermic**.

In endothermic reaction, the energy taken in to break the old bonds is greater than the energy given out when the new bonds form.

Heat Changes in a Reaction

The amount of energy involved in a reaction is known as **the heat change (heat of reaction) or enthalphy change of the reaction**. It is measured in **kilojoules (kJ)** and represented by the symbol **ΔH**. Δ is the Greek letter **delta**, which means **change**. **H** means **energy content**.

For an exothermic reaction, **ΔH is negative**. This is because the chemicals have released energy to the surroundings.

For an endothermic reaction, **ΔH is positive** the chemicals have absorbed energy from the surroundings.

For example:

- When one mole of carbon is burnt in excess oxygen, 349kJ of heat is produced. This is an exothermic reaction. The heat of reaction is ΔH is – 349 kJ. The equation for the reaction is,

$$\text{Carbon} \quad + \quad \text{Oxygen} \quad \rightarrow \quad \text{Carbon dioxide}$$
$$C_{(s)} \quad + \quad O_{2(g)} \quad \rightarrow \quad CO_{2(g)} \qquad \Delta H = -349 \text{ kJ}$$

- When one mole of hydrogen reacts with one mole of iodine, 52kJ of heat is absorbed from the surroundings. This is an endothermic reaction. The heat of reaction is ΔH is +52kJ. The equation for the reaction is,

Hydrogen + Iodine $\rightarrow$ Hydrogen iodide

$H_{2(g)} + I_{2(g)} \rightarrow 2HI_{(g)}$ $\Delta H = +52$ kJ

The difference between the energy levels of the products and reactants is equal to the amount of energy given out by the reaction and is expressed using the formula,

$$\Delta H = H_{products} - H_{reactants}$$

The table below shows the comparison of exothermic and endothermic reactions:

Exothermic reaction	Endothermic reaction
• gives out heat to the surroundings • causes an increase in temperature • has a negative ΔH • has products that have lower energy than the reactants	• takes in heat from the surroundings • cause a decrease in temperature • has a positive ΔH • has products that have higher energy than the reactants

---

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 7

40 minutes

Answer the following questions:

1. Define the following reactions:

 a. Exothermic

 ____

 ____

 b. Endothermic

 ____

 ____

---

2. Give three (3) examples each of:

 a) exothermic reaction which are only physical processes.

 (i) ____

(ii) ____
(iii) ____

b) endothermic reactions which are only physical processes.
(i) ____
(ii) ____
(iii) ____

---

3. Give the characteristics of an exothermic reaction.
(i) ____
(ii) ____

---

4. In exothermic reaction, the energy __________ when the new bonds are made is greater than the energy __________ to break the old bonds.

---

5. In endothermic reaction, the energy __________ to break the bonds are greater than the energy __________ when the new bonds form. ____________________________

---

6. When hydrogen gas reacts with bromine gas, hydrogen bromide is produced as shown in the equation below:

Hydrogen + Bromine $\longrightarrow$ Hydrogen bromide
$H_{2(g)} + Br_{2(g)} \longrightarrow 2HBr_{(g)}$

a. Using the box below, show by means of diagrams how the bonds break between hydrogen gas and bromine gas and how hydrogen bromide is made.

b. Is the breaking of bonds, an endothermic reaction? __________
c. Is the making of bonds, an exothermic reaction? __________

Thank you for completing your learning activity 7. Check your work. Answers are at the end of this module.

Reversible Reactions

Many chemical reactions can proceed in one direction only. They cannot be reversed.
For example, potassium hydroxide reacts with dilute sulphuric acid to form potassium sulphate and water.

Potassium hydroxide + Sulphuric acid → Potassium sulphate + Water

$$2KOH_{(aq)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$$

There are some chemical reactions that can be reversed. For example, when vapour of concentrated ammonia solution comes into contact with vapour of concentrated hydrochloric acid, white fumes of ammonium chloride are formed.

Ammonia + Hydrogen chloride → Ammonium chloride

$$NH_{3(g)} + HCl_{(aq)} \rightarrow NH_4Cl_{(s)}$$

Solid ammonium chloride, in turn, can be decomposed on heating to form ammonia and hydrogen chloride gases. The equation for this reaction is

Ammonium chloride → Ammonia + Hydrogen chloride

$$NH_4Cl_{(s)} \rightarrow NH_{3(g)} + HCl_{(aq)}$$

Since the reaction can go in either direction, it is a reversible reaction. A double arrow, $\rightleftharpoons$, is used to indicate a reversible reaction. The equation for the reversible reaction of ammonia and hydrogen chloride should be written as

Ammonia + Hydrogen chloride $\rightleftharpoons$ Ammonium chloride

$$NH_{3(g)} + HCl_{(aq)} \rightleftharpoons NH_4Cl_{(s)}$$

There are many other examples of reversible reactions.

Example 1. The reaction between nitrogen and hydrogen to form ammonia. It is written as

Nitrogen + Hydrogen $\rightleftharpoons$ Ammonia

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

To avoid confusion, chemists always called

- the reaction from left to right; forward reaction.
- the reaction from right to left; reverse reaction.

Reversible reaction
Nitrogen + Hydrogen $\rightleftharpoons$ Ammonia
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

Forward reaction	Reverse reaction
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$	$2NH_{3(g)} \rightleftharpoons N_{2(g)} + 3H_{2(g)}$

The rate at which the nitrogen and hydrogen react to produce ammonia is equal to the rate at which the ammonia decomposes. This situation is called **chemical equilibrium**. Because the processes continue to happen, the equilibrium is said to be dynamic.

Equilibrium (balance) is reached as ammonia dissociates into nitrogen and hydrogen. Nitrogen is an unreactive gas, and the position of equilibrium is very far towards the left-hand side. **Fritz Haber** developed this reaction to manufacture ammonia. He had to try a number of ways to get the equilibrium to move towards the right- hand side and give a more yield of ammonia. He took note of **Le Chatelier's Principle**. This states that: When conditions are changed, a system in a state of equilibrium adjusts itself in such a way as to minimize the effects of the change.

Using the above reversible equation, when the reaction goes from left to right (forward reaction) there is decrease in the number of moles of gas. If the reaction went to completion, the volume of ammonia would be only half that of the mixture of nitrogen and hydrogen used. So, if the pressure on the mixture is increased, the system can absorb the increase in pressure by reducing its volume, that is, by reacting to form ammonia. The reaction is exothermic because heat is given out; going from left to right.

From the reverse reaction above, the dissociation (separation) of ammonia is endothermic. If the temperature is raised, the system adjusts to absorb heat. Running the process at high temperature reduces the percentage conversion of the elements into ammonia. If the reactants are at very low temperature, the system takes a long time to come to equilibrium. In practice, a high pressure of about 200 atmospheres and a moderate temperature of about 450^{o}C are used. A catalyst (iron or iron(III)oxide) is used to increase the speed of reaction.

Equilibrium occurs when the forward reaction rate is equal the reverse reaction. Reversibility is shown by the double arrow.

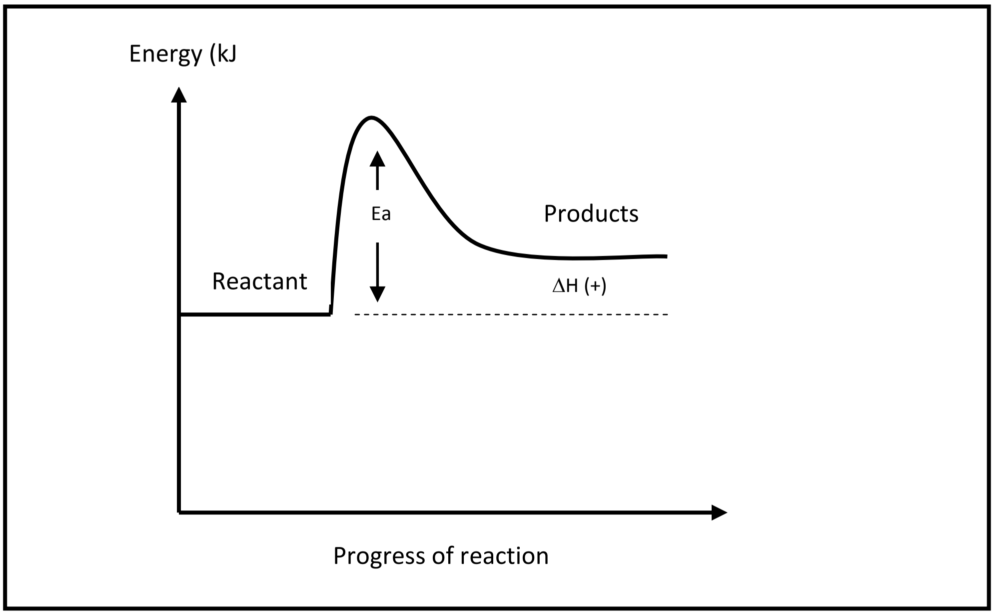
Example 2. The reaction between hydrogen and iodine to form hydrogen iodide. It is written as

Hydrogen + Iodine $\rightleftharpoons$ Hydrogen iodide

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

Consider the forward reaction:

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)} \quad +47.4\ kJ$$

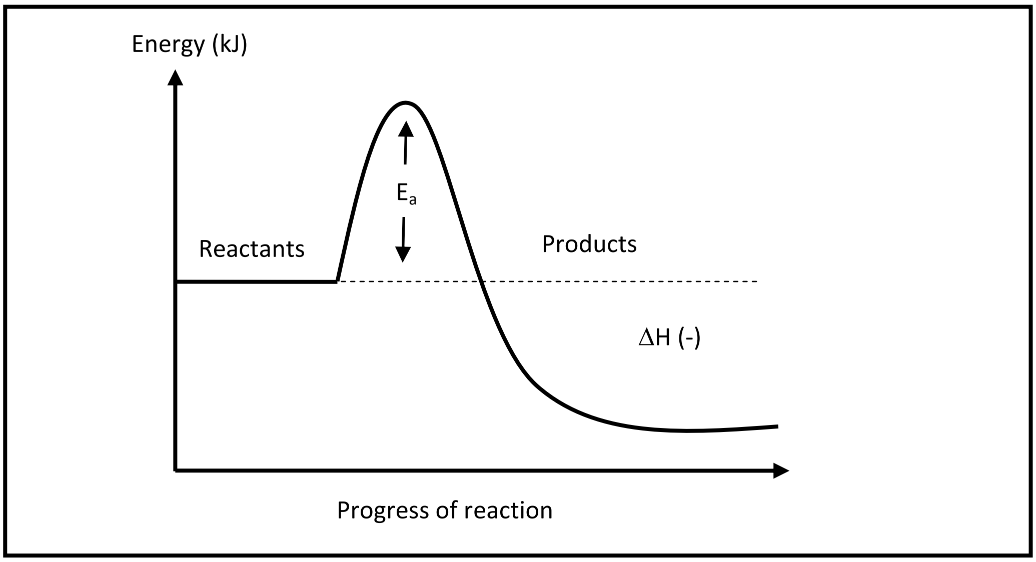


Energy change in a forward reaction.

Since ΔH is positive, this is an **endothermic reaction.**

Consider the reverse reaction:

$$2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)} \quad -47.4\ kJ$$



Energy change in reverse reaction.

Since ΔH is negative, this is an **exothermic reaction**.

Two factors that can alter equilibrium are:

a. A change in concentration of a reactant or product.
b. A change in temperature.

Factor change	Effect on equilibrium
• increase in concentration of reactants	• Shifts to decrease the concentration of the reactants
• Increase in the concentration of the products	• Shifts to decrease the concentration of the products
• For gases, increase in pressure	• Shifts to decrease the total number of molecules
• For gases, increase in volume	• Shifts to decrease the total number of molecules
• Increase in temperature	• Shifts to absorption of heat (endothermic reaction favoured)
• Decrease in temperature	• Shifts to absorption of heat (exothermic reaction favoured)

Example 3. In the reaction nitrogen dioxide forming dinitrogen tetroxide the following will happen

Nitrogen dioxide $\rightleftharpoons$ Dinitrogen tetroxide

$$2NO_{2\,(g)} \rightleftharpoons N_2O_{4\,(g)}$$

- If a product, $N_2O_{4(g)}$ is removed, more NO_2 would react to form new product, so the equilibrium would shift to the right.
- If a reactant, $NO_{2(g)}$ is added, it would again shift to the right as more $NO_{2(g)}$ would react.
- In the above reaction, adding $N_2O_{4(g)}$ would shift the equilibrium to the left as the system tries to reduce the amount of the product.
- In the above reaction, removing $NO_{2(g)}$ would result in a shift to the left as more reactant is produced.
- For gaseous reactions, an increase in pressure (or decrease in volume) would cause an increase in concentration, and a decrease in pressure (an increase in volume) would cause a decrease in concentration, only if the number of molecules is different on one side of the equation from the other.

For example:

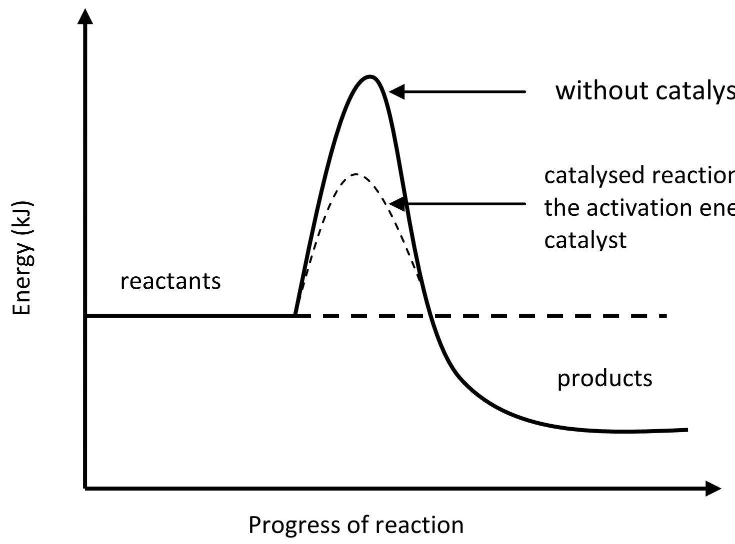
$$2NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$$

2 molecules 1 molecule

Increase in pressure will drive the reaction to the right to reduce the total volume.

Catalysts and equilibrium
Catalysts are used to speed up reactions, especially in industry. Catalysts cannot change the position of the equilibrium in a reaction. They only **reduce the activation energy**. Catalysts do not change the concentration of substances at equilibrium.

The graphs show the effect of catalyst on the activation energy of reaction.



Effect of catalyst on the activation energy of reaction.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 8 **60 minutes**

Answer the following questions:

1. Heating calcium carbonate (marble chips) until it decomposes to form calcium oxide (lime) and carbon dioxide gas is a reversible reaction as shown in the equation below:

   Calcium carbonate $\rightleftharpoons$ Calcium oxide + Carbon dioxide

a. Write a complete balanced equation for the forward reaction.
____

b. Write a balanced equation for the reverse reaction.

____

2. When heating is done in an open air and carbon dioxide has escaped, what will happen to the equilibrium? Select the letter of your answer.
The equilibrium will shift
   a. to the left to produce more calcium carbonate .
   b. to the right to produce more carbon dioxide gas.
   c. to the right and left to produce more calcium oxide.
   d. both right and left to maintain the equilibrium.

3. List two factors that will alter (destroy) the equilibrium.
   a. ____
   b. ____

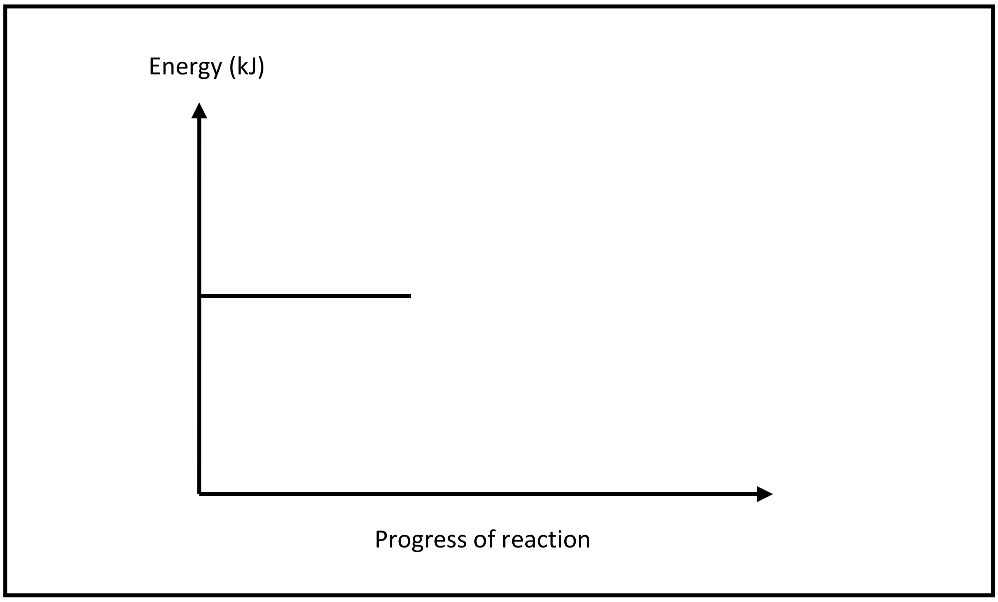
---

4. **Refer to the equation below to answer Questions a and b.**

Nitrogen + Hydrogen $\rightleftharpoons$ Ammonia

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

a. Draw an energy change for the forward reaction.



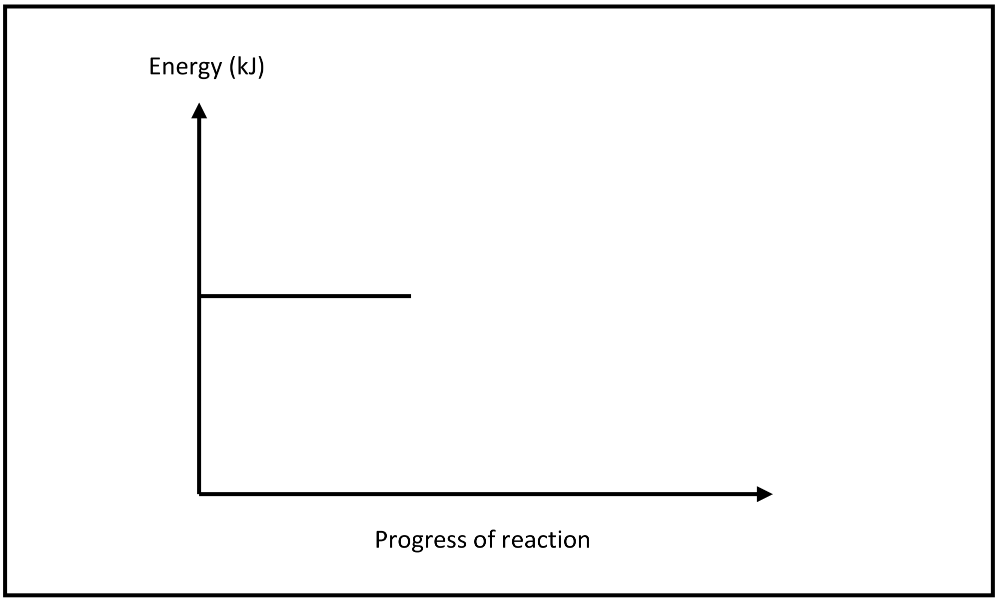
b. In your graph label the activation energy (E_a) and the heat of reaction (ΔH).

5. **Refer to the equation below to answer Questions a and b.**

$$\text{Ammonia} \rightleftharpoons \text{Nitrogen} + \text{Hydrogen}$$

$$2NH_{3(g)} \rightleftharpoons N_{2(g)} + 3H_{2(g)}$$

a. Draw an energy change for the backward (reverse) reaction.



b. In your graph, label the activation energy (E_a) and the heat of reaction (ΔH).

Thank you for completing your learning activity 8. Check your work. Answers are at the end of this module.

Bond Energy

Using bond energy values

Chemists have drawn up tables which tell exactly how much energy it takes to break different chemical bonds. To break 1 mole C – C bonds require 348 kJ. We can say that the bond energy of the C – C bond is 348 kJ. The table on the next page shows the bond energies of some bonds in kJ/mol. To break a chemical bond energy must be supplied so all the bond energies are positive. There are no bonds that fly apart by magic and release energy.

We can use energy values to calculate a value for the heat of reaction.

Bond	Energy values(kJ/mol)
H – H	436
O-O and O = O	496
C – C	348
C = C	612
C – H	412
C – O	360
C = O	743
H – O	463
Br – Br	224
C – Br	276
H - Br	366

Bond energy values

Example 1. Use bond energies to calculate ΔH for the reaction,

Methane + Oxygen ⟶ Carbon dioxide + Water

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_20_{(l)}$$

You must first show the bonds that are broken and the bonds that are made.

```
     H
     |
H —— C —— H     +    2 O = O    →    O = C = O    +    2 H —— O —— H
     |
     H
```

Next, list the bonds that are broken and the bonds that are made and their bond energies.

Bonds broken are:
4 (C – H) bonds; energy = 4 x 412 = 1648 kJ/mol
2 (O =O) bonds; energy = 2 x 496 = 992 kJ/mol
Total energy required = + 2640 kJ/mol

Bonds made are:
2 (C = O) bonds; energy = - 2 x 743 = - 1486 kJ/mol
4 (H – O) bonds; energy = - 4 x 463 = - 1852 kJ/mol
Total energy required = - 3338 kJ/mol

Heat of reaction = Energy required to break old bonds + Energy given out when new bonds are made
= + 2640 - 3338
= - 698 kJ/mol

Example 2. Calculate the heat of reaction (enthalpy change) for the equation below:

Hydrogen	+	Bromine	$\longrightarrow$	Hydrogen bromide
$H_{2(g)}$	+	$Br_{2(g)}$	$\longrightarrow$	$2HBr_{(g)}$

You are given the following bond energies:

H_2	ΔH	=	+ 436 kJ/mol
Br_2	ΔH	=	+ 224 kJ/mol
H + Br	ΔH	=	- 366 kJ/mol

Solution:

Heat absorbed to break the bonds in H_2 and Br_2 = (436 + 224) = 660 kJ/mol
Heat given out in forming 2 H-Br = 2 x 366 kJ/mol = 732kJ/mol

Therefore, heat of reaction is ΔH = Heat absorbed - Heat given out
= 660 – 732 = -72 kJ/mol

Since ΔH is negative, the reaction is **exothermic**.

Now, check what you have just learnt by trying out the learning activity below!

Learning Activity 9 **60 minutes**

Answer the following questions:

1. Write a balanced equation for the reaction between hydrogen gas and oxygen gas to form liquid water. States are required.

____

---

2. Find the heat of reaction using the energy values given below:

Bond	Energy values
H – H	436
O – O	496
H – O	463

Add the energies of the bonds that are broken as positive and the energies of the bonds made as negative.

a) Bonds broken

b) Bonds made

c) Heat of reaction

3. Calculate the heat of reaction for the incomplete combustion of methane producing carbon monoxide and water. Use the energy values on page 51.

Methane	+	Oxygen	$\longrightarrow$	Carbon monoxide	+	Water
$2CH_{4(g)}$	+	$3O_{2(g)}$	$\longrightarrow$	$2CO_{(g)}$	+	$4H_2O_{(l)}$

a) Bonds broken

b) Bonds made

c) Heat of reaction

Thank you for completing your learning activity 9. Check your work. Answers are at the end of this module.

REVISE WELL USING THE MAIN POINTS ON THE NEXT PAGE.

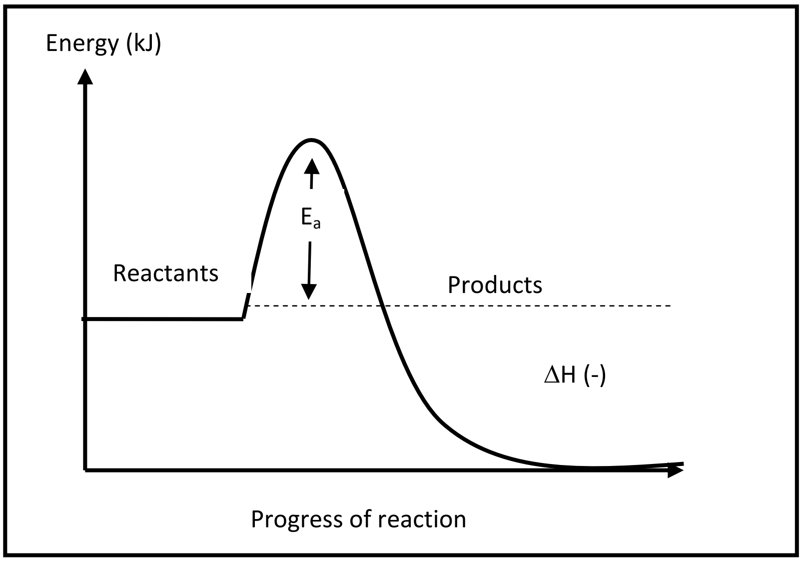
SUMMARY

You will now revise this module before doing Assessment 6. Here are the main points to help you revise. Refer to the module topic if you need more information.

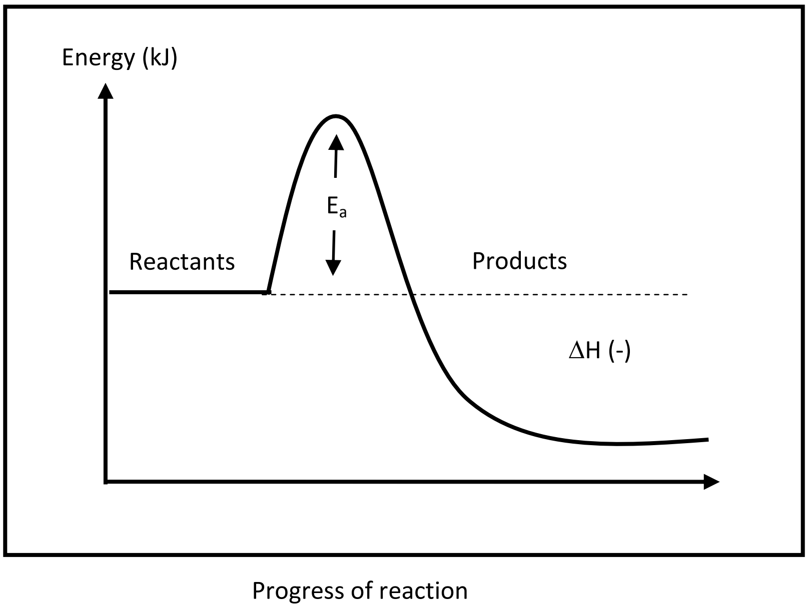
- **Reactions** takes place at different speeds: some reactions are very fast, like explosions; some are slow like rusting of iron, and some are moderate like reaction of zinc with dilute acids.
- **Five Factors Affecting the Rate of Reaction** are
    - **temperature** - the higher the temperature, the faster the reaction.
    - **concentration of reactants** – the higher the concentration, the faster the speed of reaction.
    - **surface area or the particle size of the reactants** – the smaller the particle size (or the larger the surface area) the faster is the speed of reaction.
    - **presence of catalyst** – catalyst speeds the rate of reactions, and
    - **pressure** – applicable only to gases: the higher the pressure, the faster is the
    - speed of reaction.
- **Rate** is a measure of how fast or slow something is or is a measure of the change that happens in a single module of time.
- **Rate of reaction** tells us how quickly a chemical reaction happens.
- The **speed of reaction** can be determined in three ways:
    - measuring the time taken for a reaction to complete,
    - measuring the amount of product formed against time and
    - measuring the amount of reactant used up or remaining against time.
- **Collision Theory** is a model that describes how the rate of a chemical reaction is determined by the collisions between reacting particles. This theory suggests that a chemical reaction only occurs when two reacting particles (atoms or molecules) collide with each other. They must collide with sufficient energy to break existing bonds and make new bonds to form products.
- **Catalyst** is a substance which affects the speed of a chemical reaction, but remains unchanged at the end of the reaction.
- **Characteristics of a catalyst**
    - A catalyst lowers the activation energy (the energy needed to start up a reaction).
    - Only a small amount of catalyst is needed to speed up the reaction.
    - A catalyst is selective in its action. This means one catalyst cannot act on or speed up all types of reactions. Different catalyst speed up different reactions.
    - A catalyst is not used up during the reaction.
    - Impurities can prevent catalysts from working.
    - The physical appearance of catalyst may change at the end of the reaction, but its chemical properties remain unchanged.
    - A catalysts increase the speed and not the yield of a chemical reaction. The same amount of products is formed whether a catalyst is used or not.

- At **higher temperature**, the speed of reaction is faster because the reacting particles have more kinetic energy and they collide more often and with greater force resulting in more successful collisions.
- At **higher concentrations**, more particles are crowded into the same volume. This results in more collisions between the reacting particles and increases the rate of reaction.

- An **exothermic reaction** is a reaction in which heat energy is given out to the surroundings.
- An **endothermic reaction** is a reaction in which heat energy is absorbed from the surroundings.
- In an **exothermic reaction**, the **products** formed are at **lower energy level** than that of the reactants and the excess energy is given out to the surroundings.



- In an **endothermic reaction**, the **products** formed are at **higher energy level** than that of the reactants and the energy is absorbed by the reactants from the surroundings.



- The amount of heat energy given out or absorbed during a chemical reaction is called the **enthalphy change** or **heat of reaction**, given as ΔH. In **exothermic reaction** the heat of reaction is **negative** while **endothermic reaction** is **positive**.
- Heat energy is absorbed for breaking bonds; therefore, **bond breaking** is **endothermic.**
- Heat energy is released when bonds are formed; therefore, **bond making** is **exothermic**.
- **Reversible reaction** is a reaction which can go backwards (reverse) or forward. Reactants form products, but in different conditions, the products can also react together to form the reactants again. If nothing is allowed to escape, a reversible reaction can reach a position of **equilibrium**. The amount of reactants and the products stay the same.

For example, the reaction between hydrogen gas and nitrogen gas to produce ammonia is a reversible reaction.

Reversible reaction
Nitrogen gas + Hydrogen gas $\rightleftharpoons$ Ammonia
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

Forward reaction	Reverse reaction
$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$	$2NH_{3(g)} \rightleftharpoons N_{2(g)} + 3H_{2(g)}$

- At the **point of equilibrium**, the rate of the forward reaction and backward reaction is the same.
- The **position of equilibrium** can be altered by changing the conditions. The position of equilibrium always shifts to cancel out the change introduced.
- In a reversible reaction, chemists always call the reaction from left to right is called **forward reaction** and the reaction from right to left is called **reverse** or **backward reaction**.
- Catalysts are used to speed up reactions especially in industry. Catalysts cannot change the position of the equilibrium in a reaction. They only reduce the activation energy.

- Chemists have drawn up tables which tell how much energy it takes to break different chemical bonds as follows:

Bond	Energy values
H – H	436
O – O and O = O	496
C – C	348
C = C	612
C – H	412
C – O	360
C = O	743
H – O	463
Br – Br	224
C – Br	276
H - Br	366

- **Example of Bond Energy Calculations:**

Use bond energies to calculate ΔH for the reaction,

Methane	+	Oxygen	→	Carbon dioxide	+	Water
$CH_{4(g)}$	+	$2O_{2(g)}$	→	$CO_{2(g)}$	+	$2H_20_{(l)}$

You must first show the bonds that are broken and the bonds that are made.

$$H-\underset{\displaystyle H}{\overset{\displaystyle H}{\underset{|}{\overset{|}{C}}}}-H \quad + \quad 2\ O=O \quad \rightarrow \quad O=C=O \quad + \quad 2\ H-O-H$$

Next, list the bonds that are broken and the bonds that are made and their bond energies.

- **Bonds broken** are:
  4 (C = H) bonds; energy = 4 x 412 = 1648 kJ/mol
  4 (O =O) bonds; energy = 2 x 496 = 992 kJ/mol
  Total energy required = + 2640 kJ/mol

- **Bonds made** are:
  2 (C – O) bonds; energy = - 2 x 743 = - 1486 kJ/mol
  2 (H – O) bonds; energy = - 4 x 463 = - 1852 kJ/mol
  Total energy required = - 3338 kJ/mol

NOW YOU MUST COMPLETE ASSESSMENT 4 AND RETURN IT TO THE PROVINCIAL CENTRE CO-ORDINATOR.

ANSWERS TO LEARNING ACTIVITIES 1- 9

Learning Activity 1

1. (i) The rate of reaction tells us how quickly a chemical reaction happens.
   (ii) Catalyst is a substance which speeds up the rate of reaction.

2. (i) FAST
   (ii) SLOW
   (iii) SLOW
   (iv) FAST
   (v) FAST
   (vi) SLOW
   (vii) SLOW

3. Collision Theory states that a chemical reaction is determined by the collisions between reacting particles. This theory suggests that a chemical reaction only occurs when two reacting particles, atoms or molecules, collide with each other.

4. The five factors affecting the rate of reactions are:
   (i) temperature
   (ii) concentration
   (iii) surface area
   (iv) pressure of gases
   (v) catalyst

5. activation

6. activation

7. (i) increases
   (ii) increases
   (iii) decreases
   (iv) increases

Learning Activity 2

1. a. $26cm^3$ accept $25cm^3$
   b. $26cm^3/min$ accept $25cm^3/min$
   c. $50cm^3 - 26cm^3 = 24cm^3$ accept $25cm^3$
   d. $24cm^3/min$ accept $2cm^3/min$

2. a. The reaction is over when the curve of the graph becomes flat.
   b. 7 cm^3

3. a. The reaction is fastest at 1 minute.
   b. The reaction is slowing down at 2.5 minutes.

4. a. The average rate of reaction is 70/3.5 = $20.0cm^3/min$.
   b. steeper

Learning Activity 3

1. a. $CaCO_3$
   b. 2M solution
   c. $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$
   d. The beaker with 2M concentration has a faster reaction because there are more reacting particles colliding in a given volume of solution.

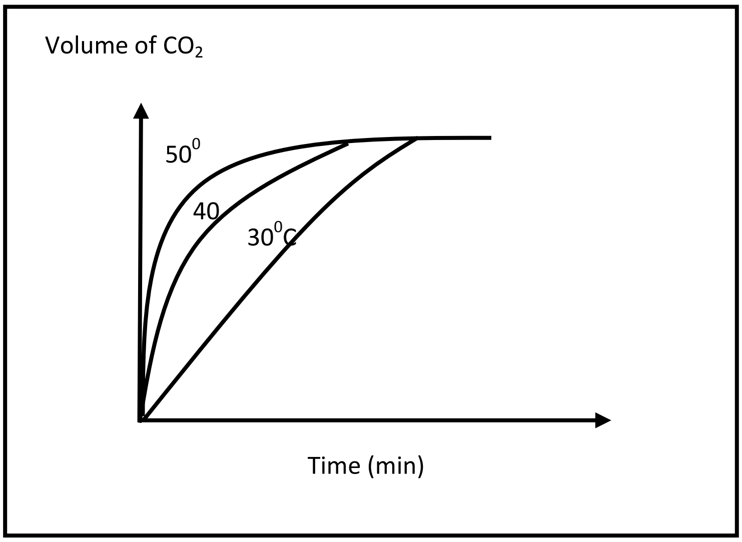
2. a. II
   b. I
   c. B
   d. III
   e. A

Learning Activity 4

1. a. 40^oC
   b. The reaction is faster when the curve of the graph is steep.

2. a. No, because the temperature only affects the speed of the reaction and not the amount of the products.

b.



3. a. faster
   b. collide
   c. increase

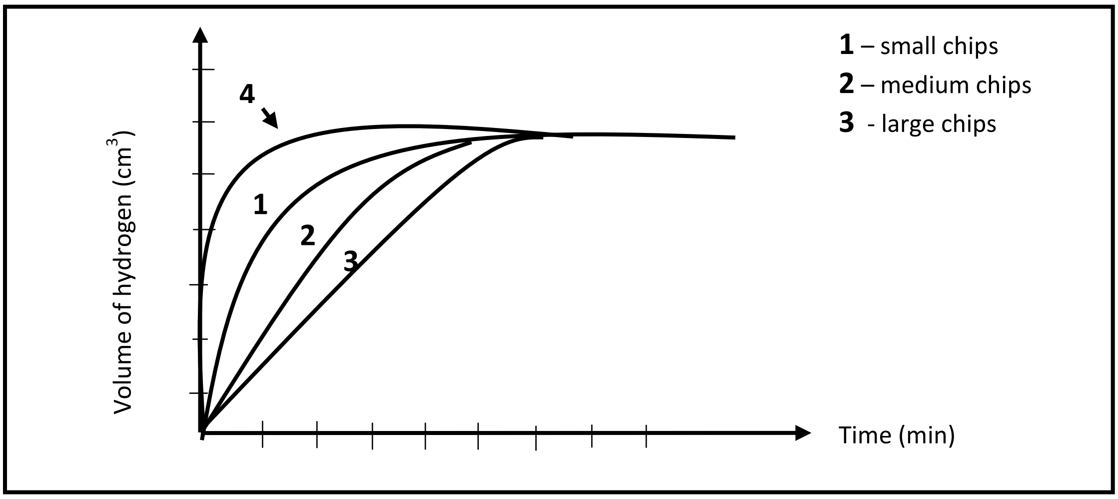
Learning Activity 5

1. a. 1 small chip

   b. 1 small chip, because there are more colliding particles.

   c. 3 large chips, because the surface areas are small for the reacting particles to collide and makes the reaction the slowest compared to small chips and medium chips.

2.



Learning Activity 6

1. a. Measuring the time taken for a reaction to complete.
   b. Measuring the amount of product formed against time.
   c. Measuring the amount of reactant used up against time.

2. a. Manganese(IV) oxide, MnO_2
   b.

$$2H_2O_{2(aq)} \xrightarrow{\text{Manganese (IV) oxide}} 2H_2O_{(l)} + O_{2(g)}$$

3. a. copper(II)oxide
   b. iron(III)hydroxide

Learning Activity 7

1. a. Exothermic reactions are reactions which give out heat energy to the surroundings.
   b. Endothermic reactions are reactions which absorb heat energy from the surroundings.

2. a. Examples of exothermic reactions are:
   - condensation
   - freezing
   - dissolving of acids in water

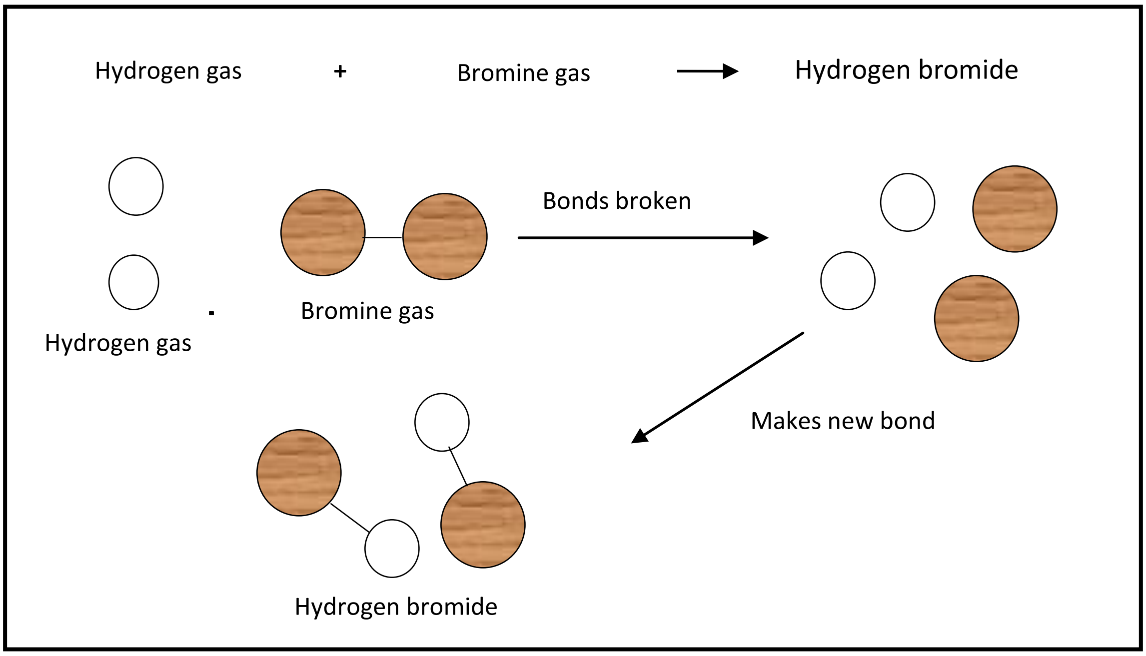
   b. Examples of endothermic reactions are:
   - evaporation
   - melting
   - dissolving of some ionic compounds like ammonium chloride and potassium nitrate.

3. (i) Heat is given out and is transferred from the chemicals to the surroundings.
   (ii) The temperature of the reaction mixture rises. The container feels hot.

4. given out
   taken in

5. takes in
   Given out

6.



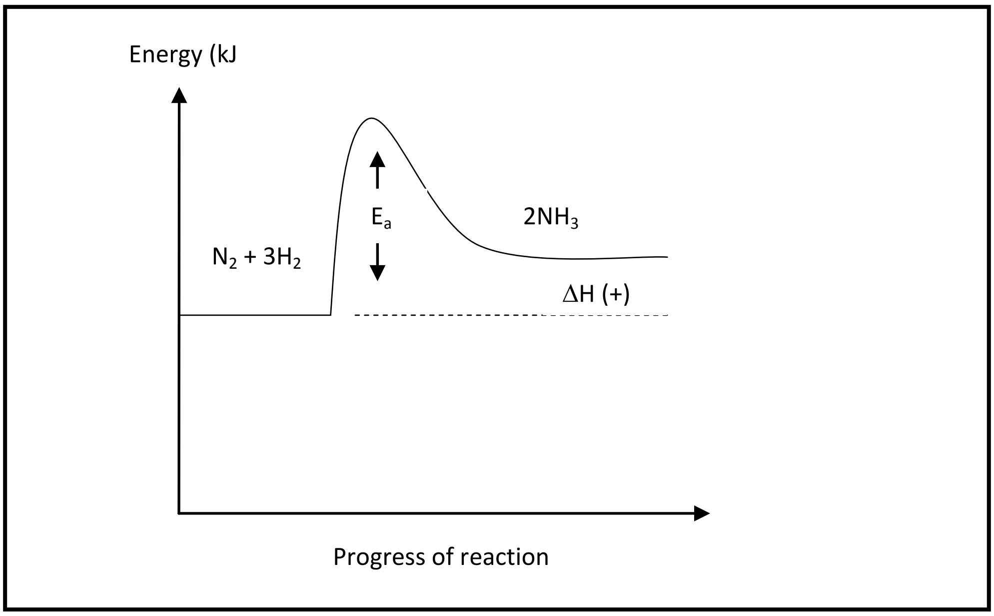
Learning Activity 8

1. a. Forward reaction: $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$
   b. Reverse reaction: $CaO_{(s)} + CO_{2(g)} \rightleftharpoons CaCO_{3(s)}$

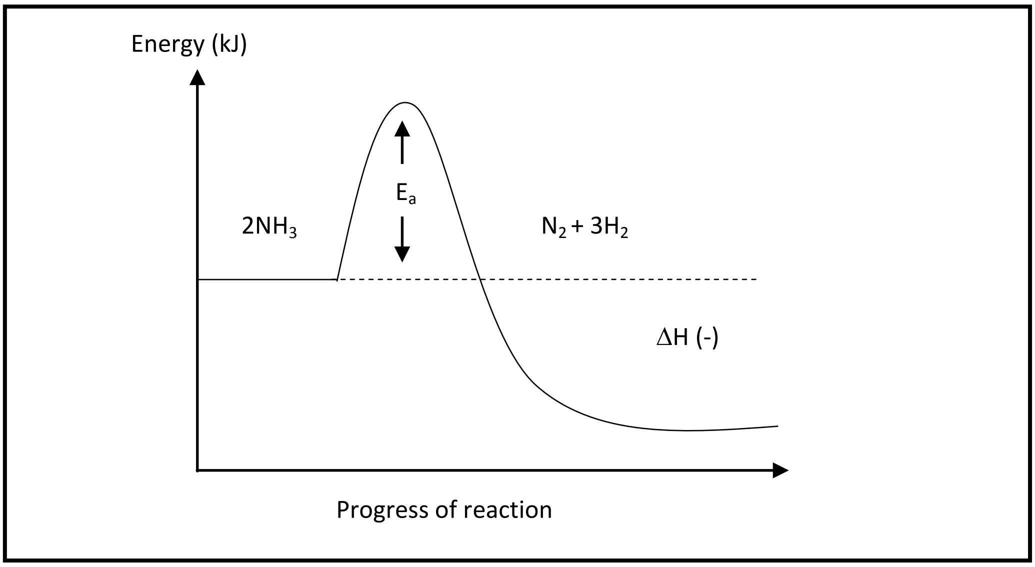
2. a

3. a. A change in concentration of a reactant or product.
   b. A change in temperature.

4. Forward reaction:



5. Backward reaction:



Learning Activity 9

1. $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$

2.

Bonds broken:

H — H
H — H + O = O

Bonds made: H – O – H   H – O – H

Heat of reaction:

Bonds broken are:

2 (H – H) bonds	=	2 x 436	=	+872 kJ/mol
1 (O = O) bond	=	1 x 496	=	+496 kJ/mol
Total energy required	=	+1368 kJ/mol		

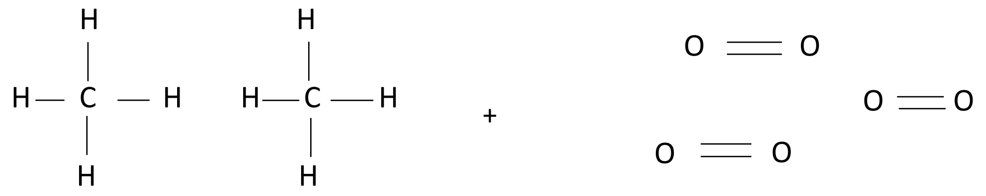
Bonds made are:

4 (H – O) bonds	=	4 x 463	=	-1852 kJ/mol
Total energy required	=	-1852 kJ/mol		

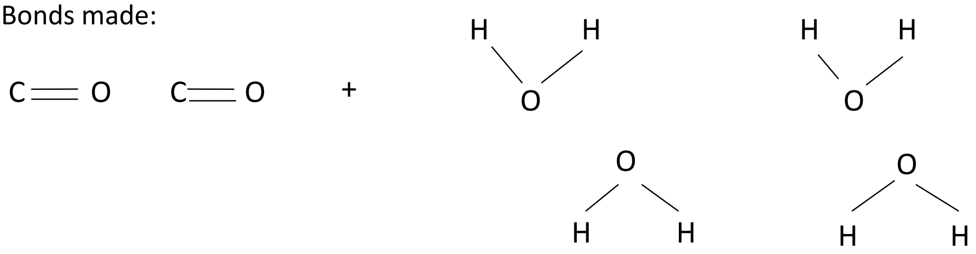
Heat of reaction	=	Energy needed to Break old bonds	+	Energy given out to make new bonds
	=	+1368 + – 1852		
	=	-484kJ/mol		

3.

Bonds broken:



Bonds made:



Heat of reaction:

Bonds broken are:

8 (C - H) bonds	=	8 X 412	=	+3296 kJ/mol
3 (O = O) bond	=	3 X 496	=	+1488 kJ/mol
Total energy required	=	+4784 kJ/mol		

Bonds made are:

8 (H – O) bonds	=	8 X 463	=	-3704kJ/mol
2 (C = O) bonds	=	2 X 743	=	-1486 kJ/mol
Total energy required	=	-5190 kJ/mol		

Heat of reaction	=	Energy needed to Break old bonds	+	Energy given out to make new bonds
	=	+4784 + - 5190		
	=	-406kJ/mol		

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FODE PROVINCIAL CENTRES CONTACTS

PC NO.	FODE PROVINCIAL CENTRE	ADDRESS	PHONE/FAX	CUG PHONE (COORDINATOR)	CUG PHONE (SENIOR CLERK)
1	ALOTAU	P. O. Box 822, Alotau	6411343/6419195	72228130	72229051
2	BUKA	P. O. Box 154, Buka	9739838	72228108	72229073
3	CENTRAL	C/- FODE HQ	3419228	72228110	72229050
4	DARU	P. O. Box 68, Daru	6459033	72228146	72229047
5	GOROKA	P. O. Box 990, Goroka	5322085/5322321	72228116	72229054
6	HELA	P. O. Box 63, Tari	73197115	72228141	72229083
7	JIWAKA	c/- FODE Hagen		72228143	72229085
8	KAVIENG	P. O. Box 284, Kavieng	9842183	72228136	72229069
9	KEREMA	P. O. Box 86, Kerema	6481303	72228124	72229049
10	KIMBE	P. O. Box 328, Kimbe	9835110	72228150	72229065
11	KUNDIAWA	P. O. Box 95, Kundiawa	5351612	72228144	72229056
12	LAE	P. O. Box 4969, Lae	4725508/4721162	72228132	72229064
13	MADANG	P. O. Box 2071, Madang	4222418	72228126	72229063
14	MANUS	P. O. Box 41, Lorengau	9709251	72228128	72229080
15	MENDI	P. O. Box 237, Mendi	5491264/72895095	72228142	72229053
16	MT HAGEN	P. O. Box 418, Mt. Hagen	5421194/5423332	72228148	72229057
17	NCD	C/- FODE HQ	3230299 ext 26	72228134	72229081
18	POPONDETTA	P. O. Box 71, Popondetta	6297160/6297678	72228138	72229052
19	RABAUL	P. O. Box 83, Kokopo	9400314	72228118	72229067
20	VANIMO	P. O. Box 38, Vanimo	4571175/4571438	72228140	72229060
21	WABAG	P. O. Box 259, Wabag	5471114	72228120	72229082
22	WEWAK	P. O. Box 583, Wewak	4562231/4561114	72228122	72229062

FODE SUBJECTS AND COURSE PROGRAMMES

GRADE LEVELS	SUBJECTS/COURSES
Grades 7 and 8	1. English
	2. Mathematics
	3. Personal Development
	4. Social Science
	5. Science
	6. Making a Living
Grades 9 and 10	1. English
	2. Mathematics
	3. Personal Development
	4. Science
	5. Social Science
	6. Business Studies
	7. Design and Technology- Computing
Grades 11 and 12	1. English – Applied English/Language& Literature
	2. Mathematics – General / Advance
	3. Science – Biology/Chemistry/Physics
	4. Social Science – History/Geography/Economics
	5. Personal Development
	6. Business Studies
	7. Information & Communication Technology

REMEMBER:

- For Grades 7 and 8, you are required to do all six (6) subjects.
- For Grades 9 and 10, you must complete five (5) subjects and one (1) optional to be certified. Business Studies and Design & Technology – Computing are optional.
- For Grades 11 and 12, you are required to complete seven (7) out of thirteen (13) subjects to be certified.

Your Provincial Coordinator or Supervisor will give you more information regarding each subject and course.

Notes: You must seek advice from your Provincial Coordinator regarding the recommended courses in each stream. Options should be discussed carefully before choosing the stream when enrolling into Grade 11. FODE will certify for the successful completion of seven subjects in Grade 12.

GRADES 11 & 12 COURSE PROGRAMMES			
No	**Science**	**Humanities**	**Business**
1	Applied English	Language & Literature	Language & Literature/Applied English
2	General / Advance Mathematics	General / Advance Mathematics	General / Advance Mathematics
3	Personal Development	Personal Development	Personal Development
4	Biology	Biology/Physics/Chemistry	Biology/Physics/Chemistry
5	Chemistry/ Physics	Geography	Economics/Geography/History
6	Geography/History/Economics	History / Economics	Business Studies
7	ICT	ICT	ICT

CERTIFICATE IN MATRICULATION STUDIES		
No	**Compulsory Courses**	**Optional Courses**
1	English 1	**Science Stream:** Biology, Chemistry and Physics
2	English 2	**Social Science Stream:** Geography, Intro to Economics and Asia and the Modern World
3	Mathematics 1	
4	Mathematics 2	
5	History of Science & Technology	

REMEMBER:

You must successfully complete 8 courses: 5 compulsory and 3 optional.