Target

in

Chemistry 3

Senior Secondary

Anderson Mpinganjira
Sydney Namaumbo
Gertrude Malema-Mwakikunga

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Bookland International,
Kenyatta Drive, Chitawira,
Private Bag 140,
Blantyre

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Experimental Techniques

Success Criteria

By the end of this unit you should be able to:

1. Identify waste products from chemical reactions
2. Describe safe ways of disposing chemical wastes in the laboratory
3. Design scientific investigations
4. Carry out scientific investigations to determine the purity of a substance
5. Describe simple tests for water, ions and gases

Key words:

In this unit you will find these key terms and concepts:

chemical waste disposal, waste management, incineration, pure substance, experimental design, chromatography, titration, filtration, anion, cation

Ensure that you understand and learn how to apply them both for your academic and real life situations.

During chemical reactions, *useful* or *desirable products* (*materials*) are produced. Examples of products obtained through chemical reactions in industries include cement, refined oils, pure metals, paints and fertilisers.

However, many types of chemical activities generate hazardous wastes which are harmful to human health and the environment. These chemical wastes may be found in different physical states such as *gas* (e.g. sulphur dioxide (SO_2), carbon dioxide (CO_2), methane (CH_4)), *liquid* (e.g. *effluent*) or *solid*. Sometimes waste products are generated in form of **heat energy**, which is simply wasted to the environment. In other cases **radioactive wastes** are generated. Radioactive wastes comprise actively decaying or elements of *unstable form*. Such elements emit different types and levels of radiations which may last for different periods of time.

Chemical wastes are harmful to human health and the environment in which they are dumped. Harmful chemicals may lead to suffocation, skin irritation or burns, mutation of cells and many other effects.

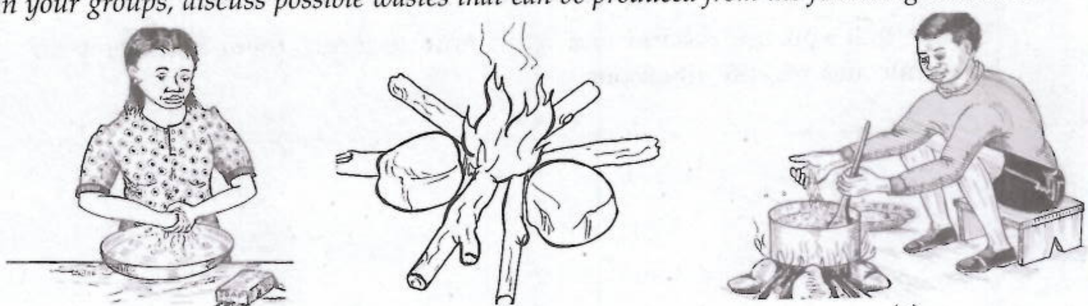
Waste products from chemical reactions

There are many chemical reactions you can carry out in the laboratory. A simple example, is burning or combustion of a candle in air to provide heat for heating up materials. During this combustion, the reaction between oxygen in the air and the candle provides **heat energy**, but also produces carbon dioxide (CO_2), carbon monoxide and soot. These wastes are produced from the **combustion reaction**. If you and your friends use several candles then substantial amounts of carbon dioxide and soot will be produced as wastes. You will now have to think of ways of removing or disposing these waste substances from the laboratory to prevent harmful exposure.

Activity 1

Aim: *Brainstorming types of wastes produced from chemical reactions*

1. *In your groups, discuss possible wastes that can be produced from the following activities.*



(*a*) (*b*) (*c*)

Figure 1.1 waste product from chemical reactions

   a) Washing the body with soap
   b) Burning firewood
   c) Cooking nsima

2. *Discuss any two chemical reactions you carry out in the laboratory and the wastes that can be produced from such reactions*
3. *Share your answers with the class*

Safe ways of disposing chemical wastes in the laboratory

From *Activity 1*, you will notice that various wastes are generated during chemical reactions. Some reactions generate solid wastes, which you can simply collect and throw in the bin or rubbish pit. But some waste materials are generated in liquid or gaseous form. Some of these waste products may be hazardous or harmful. How can you dispose such wastes?

If you are working in the laboratory you need to know how to properly dispose these wastes for you to continue working in a safe place.

Once chemical products have been generated in the laboratory, they must be treated properly to prevent hazardous effects to both human beings and the environment. The following activity will help you to understand how to manage some wastes that are hazardous.

Activity 2

Aim: *Managing chemical wastes*

The following table gives some of the measures to considered when managing chemical wastes in the laboratory. In groups, discuss and complete the table by filling in the possible reason(s) a particular measure must be followed.

Measure	Reason
Labelling all bottles of chemical wastes	
Organic waste bottles must be capped	
Do not store waste in a fume hood where reactions are being carried out	
Do not store acidic and basic wastes in the same cabinet	
Do not mix immiscible solvents in one bottle	

Table 1 Measures to be followed when managing chemical waste

Discussion

From Activity 2, *you learnt some measures that can be considered in order to manage chemical wastes. Once the waste has been generated, it must be collected and stored for proper disposal. Label the bottles properly for collecting the wastes. Labels should indicate date and type of waste collected. The bottles must be stored according to the type or nature of the chemical waste collected, because some chemicals may react and produce more harmful products. For example, when acidic wastes are stored close to organic wastes fire and/or explosion may result when they mix.*

1. Disposing of solid wastes

Solid wastes that are generated in the laboratory must be disposed carefully to avoid their harmful effects. Some solids are naturally reactive, hence if not carefully managed they may react with other solid, liquid as well as gaseous substances around them. Most solid wastes are harmful to the environment, hence the need for proper care when disposing them.

The following are some of the ways you should follow when disposing the waste:

a. Store solid waste in a well sealed container labelled "*hazardous solids*".

b. Reactive solid waste must be stored in separate containers.

c. Very harmful solid waste must be incinerated (*burnt*). This significantly reduces their accumulation.

d. Less harmful solid waste such as paper, plastic, rubber and wood should be placed in waste bins available in the laboratory. These will later be collected and disposed into a rubbish pit or the local authority refuse collection area.

e. Some less harmful waste such as paper and plastic can be recycled into useful products and reduce their accumulation.

2. Disposing of liquid wastes

In the laboratory a lot of liquids are used while a lot of by-product liquid wastes are generated. These wastes could be harmful to both human beings and the environment. Depending on the characteristics of the liquid waste, you should first treat it before disposing to minimise its effects. As a matter of fact, a good laboratory must have a sink used for emptying or disposal of some liquid wastes.

The following liquid wastes must be collected and managed as hazardous wastes and never discharged into the sanitary sewer or sink in any amount:

a. *Raw chemical wastes:* Including unused, pure or concentrated chemicals

b. *Chlorinated hydrocarbons wastes*: Chemical wastes containing chlorine, hydrogen and carbon atoms such as chloromethane

c. *Chlorofluorocarbon (CFC) wastes:* Wastes containing carbon, chlorine and fluorine atom. They are also called freons.

d. *Brominated Hydrocarbon wastes*

e. *Cyanide wastes:* These waste include cyanide, cyanate (OCN-) and thiocyanate (SCN) compounds such as Potassium cyanide, zinc cyanide, hydrogen cyanide, etc.

f. *Heavy metal wastes* such as antimony, mercury, arsenic, cooper, silver, etc.

g. *Corrosive wastes:* All wastes that could cause structural damage to the laboratory, sinks or sewer piping. These wastes usually have *pH* lower than 5 (*strong acids*) and higher than 9 (*strong bases*)

h. *Solvent wastes:* All wastes containing the following solvents in any concentration: acetone, ethylether, benzene, methanol, isobutanol, nitro-benzene, carbon disulphide, etc.

i. *Oil and Grease wastes:* Including vacuum pump oil

j. *Ignitable wastes.* All wastes which are a mixture of ignitable chemicals and other chemicals

Other groups of chemicals that **must not be discharged into the sinks** include: reactive wastes, solid or sticky wastes, untreated waste, hot liquid or vapour wastes, etc.

> **Warning:** *Containers labelled 'Hazardous waste' should be handled with care.*

3. Disposing of gaseous wastes

Most chemical reactions produce gases. Some of these gases are harmful both to human beings and the environment. These gases include: carbon dioxide (CO_2), oxides of nitrogen (NO, NO_2), hydrogen sulphide (H_2S) and oxides of sulphur (SO_2) and cyanide. When conducting experiments that emit gases do not expose youself or your colleague to them. It is advisable to carry out such reactions in a fume hood; which suck fumes out into the sky. If you do not have fume hood in your laboratory, such reactions can be done in a well ventilated room or open space.

Warning: *Avoid inhaling such fumes because they can cause respiratory problems.*

Exercise 1

1. A student burn a piece of cloth on a spirit burner during an activity
   a) What form of waste would be produced from the activity above?
   b) Describe one way you dispose of the waste you have named in (a).
2. Describe a safe way of disposing of:
   a) acidic waste
   b) liquid waste

Scientific investigations

At Junior Secondary level in Chemistry, you learnt that scientific investigations are carried out in order to develop *knowledge*, *skills* and *establish facts*. You learnt that scientific investigations are carried out following an orderly procedure in order to ensure that accurate results are obtained. Please remember that an investigation should follow the following steps:

1. Identifying a problem

The first step in making a scientific inquiry is to identify what the problem is. The problem will need a solution(s). All scientific advances, are a result of the problem that existed before.

Note: A problem should always be in form of a question.

2. Formulating the hypothesis

If you may recall, the *hypothesis* is a predicted answer you give as a solution to the problem you want to investigate.

Note: Always come up with *at least two hypotheses*; at least one positive and the other negative, so that if one fails the opposite one may be true.

3. Selecting variables

Variables, you may recall, are factors that affect the results of an investigation. They are usually derived from the hypotheses. For example, if you want to find out the effect of temperature on the rate of reaction of sodium chloride with sulphuric acid, you would suggest to carry out the reaction at various temperatures to observe some effects. Therefore, "*temperature*" is a variable *which you can manipulated in this investigation*. It is an independent variable, which you will need to change in order to observe the change in the rate of reaction.

Note: A dependent variable is a factor that is affected as a result of change in the independent variable during the investigation.

4. **Controlling the variables**

In order to test each hypothesis you need to *fix* or *hold other variables **constant***, while you change the other. This is known as *controlling the variables*. In the example above, to find out the effect of temperature on the rate of reaction of sodium chloride with sulphuric acid you will need to hold other variables (*such as volumes and concentrations*) constant.

5. **Testing the hypothesis**

At this stage you are actually doing the investigation. You should therefore have identified your *independent*, *dependent* and *controlled variables* with materials and apparatus to use in the investigation.

6. **Recording Data**

During the investigation you will make observations and will therefore need to record the information or data obtained. This data will help you when making conclusions concerning the factors under investigation. The flow of your data will show whether you have identified and controlled the variables correctly. Data can be recorded in form of *numbers*, *sentences*, *patterns*, *pictures of observations* you make.

7. **Presenting and Interpreting Data**

Finally, you will need to present your findings to others. There are various ways in which you may present your data for easy and meaningful interpretation. Two common ways are tables and graphs. These methods will help you present data as neatly and orderly as possible such that you can easily note the relationship between the dependent and independent variables for ease ***interpretation*** of your results.

To carry out an investigation you need to *plan well* so that the investigation is successful. In this section you will practice how to design a scientific investigation. In your design, you will be required to *state the problem*, *hypotheses*, *variables* and *how you would control them*. After that, you will be required to *record your findings* and *report the results*.

Activity 3

Aim: *Setting up an investigation*

Think about this:

While at home with a friend, you decide to prepare the sweetest tea. Your friend, who does not take very hot tea, decides to cool her tea first before adding sugar to it. To your surprise you see that she uses less sugar than you, in your hot tea before it got saturated.

1. *What is the problem you would establish in the situation above? Put your problem in question form.*
2. *Can you formulate hypotheses for the problem that you have identified above.*
3. *What are the variables you will work with in the investigation? Which one should be independent and dependent variable?*
4. *What materials will you need to use in this investigation?*
5. *Now set up the experiment to test your hypotheses and collect the required data.*
6. *Record your data in a table for easy reading.*
7. *Look for patterns in the data and state the effect of temperature on the dissolution of sugar.*

For instance, if ***hot water dissolves more sugar than cold water'*** *is your problem, then the question you would raise is:* ***Does the difference in the temperatures of water affect the amount of sugar dissolved in it?***

You have to develop two opposing hypotheses (H1 and H2), for example:

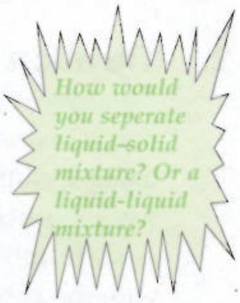
H1: Increasing the temperature of water increases the amount of sugar dissolved in it.
H2: Increasing the temperature of water does not increase the amount of sugar dissolved in it.

The experimental set up will help you find out the effect of temperature on dissolution of sugar. As you carry out the investigation then you will make observations of what happens and record the findings. Examples of materials you will need are: gas burner, water, cups, sugar and stirrer.

Designing Scientific Investigations

In the section above you have been reminded of how you would carry out scientific investigations. You will carry out many investigations and experiments in your chemistry course. In this section, you will learn how to design and carry out experiments involving some *analytical techniques*, used in chemistry. These techniques are meant to help you *acquire skills and knowledge to separate*, *identify* and *determine relative amounts* of one or more components of a substance. The techniques you will use are as follows: *filtration*, *titration*, *distillation*, and *chromatography*.

In all these activities you learnt how to separate a desired component from a substance and determine its relative amount. Components of a substance may be *in form of atoms*, *groups of atoms* or *compounds* which may need to be separated by chemical or physical means. The choice of separation *method may therefore depend on whether we want to separate into atoms* or *compounds* or whether the substance is in solid, liquid or gaseous state. For instance you may have a liquid-solid mixture of sand and water. You can also have a liquid-liquid mixture of ink.



Activity 4

Aim: *Separating Mixtures*

1. *Mention any three mixtures you know.*
2. *For each of the mixtures you have named in question 1, discuss a way you would use to separate one important component for you to use.*

After discussing mixtures and ways of separating them, we can now study some common methods of separating mixtures in chemistry.

1. Filtration

Filtration is a method which is used to separate a liquid-solid mixture. The apparatus you would use for a simple filtration process include a filter funnel fitted with a filter paper and a collecting tube or beaker as shown in *Figure 1*. The substance that remains on the filter paper after filtration is the residue and is in solid state. The liquid which passes through the filter paper and is collected into the collecting vessel is called the filtrate.

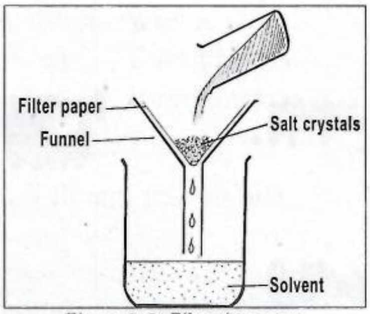


Figure 1: 2: Filtration set up

Activity 5

Aim: *Separating a mixture of salt and sand*

Materials: • *spoon* • *salt* • *sand* • *water* • *beaker*
• *stirrer* • *filter funnel* • *filter paper* • *balance.*

Procedure:

1. *Weigh 100g of sand and 50g of table salt.*
2. *Thoroughly mix the measured amounts of salt and sand in a beaker*
3. *Describe what you would do to separate the mixture*
4. *What materials will you use to separate the mixture*
5. *Then carry out an activity to separate the salt from the sand and recover both substances in their origin state*

Results:

Have you obtained the same masses that you mixed?

Write a laboratory report describing the aim, materials used for the activity, procedure followed, your observations and the results of the activity.

2. Titration

Titration is the gradual addition of a solution to another one. it is used to determine the concentration of a solution using a solution with a known concentration. The solution whose concentration we want to find out is called the ***analyte*** and is often placed in a conical flask. The solution with the known concentration, called the ***titrant***, is placed in a burrette and is released slowly in single droplets into the analyte.

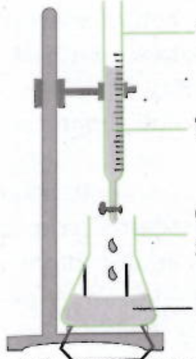


Figure 1:3: Set up of apparatus for titration

3. Chromatography

From your Junior Secondary Chemistry course, you learnt that chromatography is another method of separating and identifying components of a mixture.

What type of mixture do you seperate using chromatography?

In chromatography, the mixture to be identified or separated is called the ***analyte***. It must be in a solution of a solvent which is able to flow up the surface. Such a solvent is called a ***mobile phase***. A small amount of the mixture is put on an absorbent material called the ***stationary phase***, where it can flow with the mobile phase as it dissolves. *Good examples of absorbent materials are **chalk and filter paper**.*

The solvent that flows with the analyte is called the ***eluent***. The principle used is the fact that, *different substances in the mixture dissolve at different rates and get adsorbed to the absorbent material differently*. The components of the mixture will get partitioned between the solvent and stationary phase. *The component which is most soluble and least attracted to* *the absorbent material, moves furthest from the starting point.* It has low retention time compared to the others that are less soluble and get absorbed strongly.

The simplest chromatography is called the *paper* or *chalk chromatography*. Paper chromatography uses paper, while chalk chromatography uses chalk.

Activity 6

Aim: *Separating and comparing the components of black and blue ink by paper chromatography*

Materials:
- *black ink*
- *ethanol*
- *a strip of filter paper*
- *blue ink*
- *stirrer*
- *beaker*

Procedure:

1. *Design an experiment you would carry out to separate and compare the components of black and blue inks.*
2. *Explain your experimental design to your class.*
3. *Carry out the separation and compare the components of the two inks. Which one has more components?*

Results:

How many components do you obtain from: (i) black ink (ii) blue ink?

Write a laboratory report describing your activity. The report must include: the aim, materials used, procedure, observations, discussions and conclusion.

What you did from activity 6 above is to separate ink into its components. So far you have observed that pen ink is a mixture. You have also found out that leaves contain mixture of pigments that can be separated by chromatography.

Exercise 2

1. Explain why it is important to follow the right procedure when carrying out a scientific investigation?
2. What is a hypothesis?
3. Explain what is meant by:
   a) controlling the variables
   b) interpreting data

Determining the Purity of a Substance

Most substances found in nature are mixtures. A pure substance without traces of impurities can be obtained through *extraction* or *separation* using such techniques as *distillation* and *chromatography*. Obtaining a pure substance could be a long and expensive process, hence useful substances, that have their purity slightly less than 100%, can still be used as they are.

How would you know that a substance is pure? You use the properties of a substance to determine their purity and identity. The following are some of the properties you can use:

1. Melting and boiling points of a substance
2. Density of a substance

3. Colour of a substance
4. Ordour of a substance
5. Vapour pressure of a substance
6. Electrical and thermal conductivity of a substance
7. Number and types of components a substance contains

Some pure substances are elements while others are compounds. In this section, you will learn how to identify pure substances using their melting and boiling points. You will also learn how to determine the purity of a substance using paper chromatography.

1. Using melting and boiling points

A pure substance has uniform melting and boiling points. While a substance with impurities has fluctuating melting and boiling points. They become lower than normal since the binding forces in it are disrupted. These processes can also start over a range of temperatures.

Activity 7

Aim: *Using melting and boiling points to identify a pure substance*

Materials: • *ice blocks* • *beaker* • *thermometer* • *burner*

Proeedure:

1. *Put some fine ice blocks in a beaker.*
2. *Use a thermometer to measure the temperature of the ice.*
3. *Stir the ice at intervals of 2 minutes with the thermometer until the ice melts. Measure the temperature of the ice at each interval and record it.*
4. *Continue taking the temperature of the molten ice* ***at 2 minute intervals*** *until all the ice has changed to liquid water.*
5. *Now heat the water using a burner. Constantly stir and take the temperature* ***every 2 minutes*** *until it starts to boil.*

Results:

1. *At what temperatures does the water melt and boil?*
2. *Draw a graph of temperature against time of heating of the ice to water vapour. Describe the shape of the graph.*
3. *Is your water pure? Explain.*

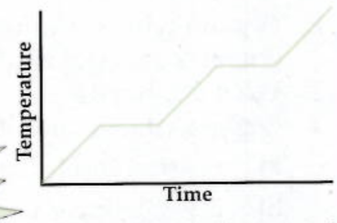


Fig. 1.4 (a): Heating curve of a pure substance

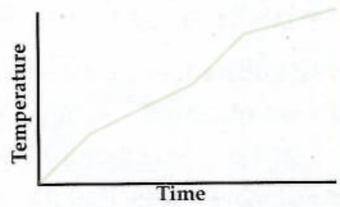


Fig. 1.4 (b): Heating curve of an impure substance

Note:

The graph line of melting and boiling points of a pure substance (*Fig. 1.4 (a)*) will show a rising gradient which will become constant at the time when the substance is melting or boiling. While the graph of an impure substance (*Fig. 1.4 (b)*) does not have constant portions (*no horizontal portions*).

The table below shows melting and boiling points of some common substances:

Substance	*Melting point* (°C)	*Boiling point* (°C)
Water	0	100
Ethanol	-114	78
Mercury	-15	357
Ammonia	-78	-27
Soldium Chloride	801	2250

2. Determining purity of a substance by using paper chromatophy

Paper chromatography is an analytical technique for separating and identifying components of mixtures that are or can be coloured. To identify each component you simply compare the chromatogram you have obtained with chromatograms of mixtures with known composition, which are called *standard chromatograms*. You may also use their *retention fraction* (R_f) for comparison since every compound has a specific retention fraction.

Retention fraction *is the ratio of distance moved by the analyte to the one moved by the solvent on the chromatography paper*. If the R_f are the same for your analyte and that component in a standard chromatogram, then the components are the same. You can calculate the R_f values by this formula:

$$R_f = \frac{\text{Distance moved by the analyte from the origin}}{\text{Distance moved by the solvent from the origin}}$$

The distances can be measured as shown in *Figure 1:4*.

For impure substances, the chromatogram gives different spots, as different components have different adsorption and retention power. A pure substance can be differentiated from impure ones since it gives only one spot which is usually equal to the standard chromatograms.

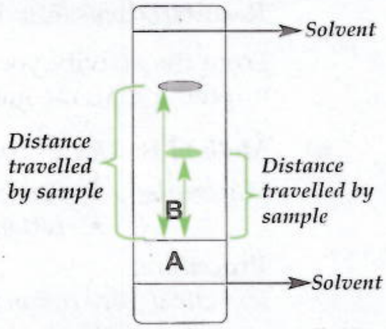


Fig. 1.4: Showing distance travelled by dyes present in ink

Exercise 3

1. Mention two properties of a substance you would use to determine its purity.
2. How would you identify a substance by using paper chromatography?
3. In chromatography, what is meant by retention fraction of the analyte?

Simple tests for identifying gases, water and ions

Different substances have varying properties and can be identified differently. In this section, you will study properties of *common gases*, *water* and *ions* that can be used to identify them.

Common gases

In this section, you will study properties of ammonia, carbon dioxide, hydrogen, oxygen and sulphur dioxide gases. You will also learn how to identify them using the same properties.

1. Ammonia gas

The ammonia gas has the following properties

a) It has no colour
b) It has a characteristic pungent smell
c) It turns moist litmus paper from red to blue
d) Turn moist universal indicator paper to blue
e) It dissolves in water to form a basic solution
f) It does not support burning.

You can, therefore, use different practical activities to identify ammonia gas. The following are some of them:

a. Test for ammonia gas

Activity 8

Aim: *Testing for Ammonia gas*

i) ***Method A: Use of Red Litmus***

Materials: • *Red litmus paper* • *beaker of ammonia solution*

Procedure:

1. *Put 20ml of ammonia solution in a beaker*
2. *Place a damp red litmus on top of the beaker containing ammonia solution and observe.*

Results/Discussion: *What colour does the litmus paper turns to?*

From the activity you might have observed that, the red litmus paper turns blue; implying that, the gas is basic or alkaline. Ammonia is a well known alkaline gas.

ii) ***Method B: Use of concentrated HCl***

Materials: • *Concentrated hydrochloric acid* • *ammonia solution* • *cotton wool* • *2 stoppers* • *glass tube.*

Procedure:

1. *Dip a piece cotton wool into the concentrated HCl and stick it to one stopper. Label this end HCl*
2. *Dip the other piece of cotton wool in ammonia solution and stick it to the other stopper. Label this end NH_3*
3. *Close the glass tube with the stoppers with cotton wool on both sides as shown in the Figure 6:*

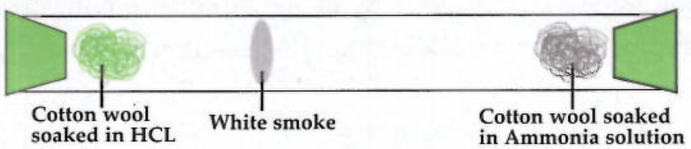


Figure 1: 5: Reaction of HCl and NH3 gases

4. *Observe for at least 10 minutes and record your observations.*

Results/Discussion:

What is formed in the tube when the two gases meet?

Ammonia gas forms a white smoke in presence of hydrochloric gas. A white smoke formed is called *ammonium chloride* [$NH_4Cl_{(s)}$] and it is formed according to the reaction:

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

b. Test for carbon dioxide gas

Carbon dioxide is one of the commonest gases found in air. It is important to plants in their food making process. Naturally all living things release carbon dioxide when they respire.

Activity 9

Aim: *Testing for Carbon dioxide gas*

Materials: • *Test tube* • *glass tube* • *lime water*

Procedure:

1. *Put 5 ml of lime water (aqueous calcium hydrogen solution) into the test tube.*
2. *Insert the glass tube into the lime water in the beaker.*
3. *Breathe out through the glass tube into the beaker* (Figure 1:6) *and observe changes in the beaker.*

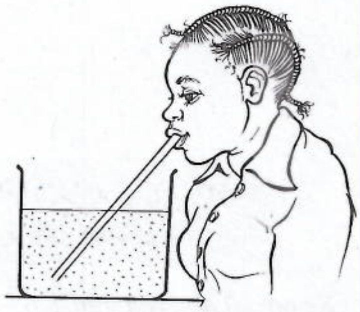


Figure 1. 6: A test for carbon dioxide

Results/Observation:

1. *What happens to the colour of the lime water?*
2. *What do you think will happen if the solution used in the beaker is water? Try it.*

Conclusion/Discussion

You have observed from the activity that lime water turns cloudy or forms a milky white precipitate. When carbon dioxide gas is bubbled through lime water, a chemical reaction takes place. It forms a white precipitate called calcium carbonate according to the equation:

$$\underbrace{Ca(OH)_{2(aq)}}_{\text{Calcium hydroxide (Lime water)}} + CO_2 \longrightarrow \underbrace{CaCO_{3(s)}}_{\text{Calcium carbonate}} + H_2O$$

However, you must be careful when conducting this experiment because when supply of carbon dioxide is in excess, the precipitate dissolves and forms Calcium hydrogen carbonate according to

$$CaCO_3 + H_2O + CO_2 \longrightarrow \underbrace{Ca(HCO_3)_{2(aq)}}_{\text{Calcium hydrogen carbonate}}$$

c. Test for Hydrogen

Hydrogen has the following properties:

i) It is a colourless gas

ii) It has no smell.

iii) It explodes when it is exposed to burning

In *Activity 10*, you will learn to test for hydrogen gas.

Activity 10

Aim: *Testing for hydrogen gas*

Materials: • *Test tube* • *hydrogen gas in a stoppered test-tube* • *lit splint*

Procedure:

1. *Hold a small test-tube of hydrogen gas up side down*
2. *Remove the stopper from the test tube*
3. *Hold a lighted splint below the open mouth of the test-tube and observe*

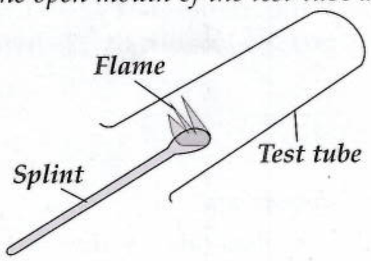


***Figure 1:** 7 Test for Hydrogen gas*

Note: If you held the test tube upright the hydrogen gas will rise and escape as it is less dense than air.

Results/Observations:

What happens when oxygen in air combines with hydrogen in the test tube?

Conclusion/Discussion:

When oxygen and hydrogen combine they react and produce hydrogen oxide. This is an explosive reaction, which makes a popping sound. This explosion occurs as hydrogen reacts with oxygen according to this equation:

Hydrogen gas + ***Oxygen gas*** $\longrightarrow$ ***Water*** + ***Energy***

This is unique for the reaction between hydrogen and oxygen to form hydrogen oxide. Remember that the common name for hydrogen oxide is water.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

d. Test for Oxygen

Oxygen is one of the important gases which is vital for all living things on earth. It has the following properties:

i) It is colourless.
ii) It has no ordour.
iii) It supports burning.

In *Activity 11*, you will test for the presence of oxygen gas:

Activity 11

Aim: *Testing for oxygen gas*

Materials: • *Matches* • *oxygen gas in the test-tube* • *splint.*

Procedure:

1. *Light the splint and then blow it out*
2. *Plunge the glowing end of the splint into the test-tube containing oxygen*

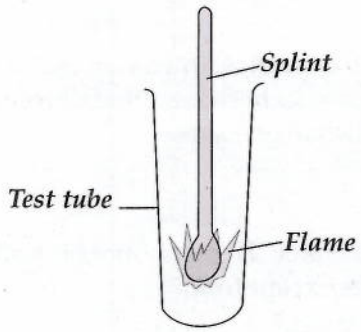


Figure 1: 8 test for Oxygen

Results/Observations:

Conclusion/Discussion:

From the activity, you have observed that oxygen gas relights a glowing splint into a flame. This means that oxygen gas supports burning.

e. Test for Sulphur dioxide

Sulphur dioxide has the following properties:

i) It is colouress.

ii) It is a poisonous gas.

iii) It dissolves in water to form an acidic solution.

In *Activity 12*, you will test for sulphur dioxide gas.

Activity 12

Aim: *testing for sulphur dioxide*

Materials: *Suphur dioxide in a beaker and potassium dichromate (vi) paper:*

Procedure:

Put a freshly made potassium dichromate (vi) paper on top of open beaker containing sulphur dioxide and observe

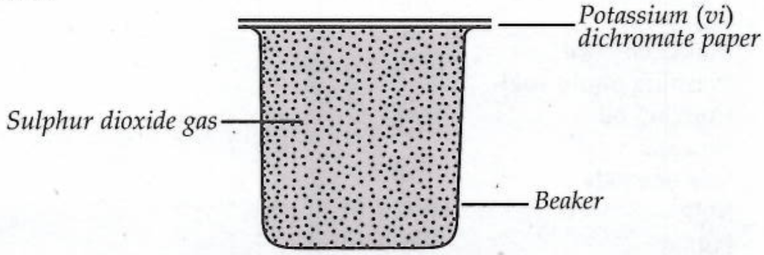


Figure 1:9: Test for sulphur dioxide gas

Results/Observations:

What happens to the colour of potassium IV paper?

Conclusion/Discussion:

Potassium dichromate (vi) paper is orange in colour and it changes to green when it is exposed to sulphur dioxide. This is so because the dichromate (vi) ion (Cr^2O^{72-}) which is orange is reduced to Cr^{3+} ion, which is green.

f. Test for water

Some substances contain water. This may not be visible until you test for it. Water has no obvious properties except that:

- It turns **anhydrous copper sulphate blue**.
  Anhydrous copper sulphate is a white powder. It changes to blue when it is mixed with water.
- It turns blue anhydrous colbalt (II) chloride paper to pink

Activities 13 and *14* will enable you to test for the presence of water in some substances using the anhydrous copper sulphate and cobalt (II) chloride paper, respectively.

Activity 13

Aim: *investigating the presence of water in different substance*

Materials:
- *Pipette*
- *Anhydrous $CuSO_4$*
- *Malambe juice*
- *Oil*
- *Tomato*
- *Milk*
- *Petri dishes*
- *Apple juice*
- *potato*
- *Vinegar*
- *paper*
- *Spatula*
- *cocacola*
- *washing liquid soap*
- *Vaseline petroleum jelly*
- *Flour*

Procedure:

1. *Put some anhydrous copper sulphate in the petridish*
2. *Draw some apple juice in the pipette*
3. *Add few drops of the apple juice onto the anhydrous copper sulphate in the Petridish, and record your observation in the table below.*
4. *Repeat steps 1 - 3 with substances given in the* Table 1.4.

No	Substance	Observation
1	Apple juice	
2	Coca cola	
3	Malambe juice	
4	Washing liquid soap	
5	Cooking oil	
6	Vinegar	
7	Vaseline jelly	
8	Milk	
9	Potato	
10	Tomato juice	
11	Paper	

Note: *Use a spatula to add anhydrous copper (II) sulphate to a solid substance.*

Conclusion/Discussion:

From *Activity 13*, you have observed that anhydrous copper sulphate turns blue when mixed with a substance that contains water such as apple juice, coca cola, vinegar, milk, potatoes, malambe juice and tomato. However, the colour of anhydrous copper sulphate does not change when mixed with a substance that does not contain water. You have learnt that both solid and liquid substance can contain water.

Activity 14

Aim: *investigating the presence of water using anhydrous cobalt (II) chloride paper*

Materials:
- *A dry anhydrous cobalt (II) chloride paper*
- *Oil*
- *Tomato juice*
- *Vinegar*
- *2 Test tubes*

Procedure:

1. *Put some oil in a test tube A and vinegar in test tube B.*
2. *Dip anhydrous cobalt (II) chloride paper in each test tube. What do you observe?*

Test-tube	Substance	Observation
A	Cooking oil	
B	Vinegar	
C	Tomato juice	

Conclusion/Discussion:

In vinegar and tomato juice, cobalt (II) chloride paper turns from blue to pink but it remains unchanged in oil. From the observations, we can conclude that vinegar and tomato juice contain water, while cooking oil does not.

Test for ions

Different ions have different properties which can be used to identify them. In this section, you will learn some properties of nitrates, sulphates and cations and how you can identify them.

1. Test for nitrates (NO_3^-)

Heating test

Some nitrate compounds release nitrogen oxide gas when heated.

a. Lithium nitrate + Heat $\rightarrow$ Lithium oxide + Nitrogen oxide

$$LiNO_3^- + \text{Heat} \rightarrow Li_2O + NO_2$$

b. Lead nitrate + Heat $\rightarrow$ Lead oxide + Nitrogen oxide

$$Pb(NO_3)_2 + \text{Heat} \rightarrow PbO + NO_2$$

However, this test does not work with sodium and potassium nitrates. Instead, we use a test called *Brown ring*. This test forms a brown ring when it is positive.

Brown ring Test

You can use this test for any nitrate. In *Activity 15*, you will investigate the presence of nitrates using the Brown ring test.

Activity 15

Aim: *To test for the presence of nitrates*

Materials:
- *2 compounds labeled X and Y* (one is a nitrate)
- *2 test tubes*
- *test tube rack*
- *dropper*
- *iron sulphate solution*
- *concentrated sulphuric acid*

Procedure:

1. *Put one 5 mls of a compound under test in a test tube*
2. *Add 3cm³ of iron (II) sulphate solution into the testtube*
3. *Carefully add a few drops of concentrated sulphuric acid*
4. *What do you observe? Which compound forms a brown ring in the test tube*

Conclusion/Discussion:

A solution of a nitrate compound gives a positive test in the activity.There is a brown ring formed when you mix a nitrate with an iron (II) sulphate solution and sulphuric acid. It is formed where the layer of concentrated sulphuric acid meets the solution of a nitrate and iron (II) sulphate solution.

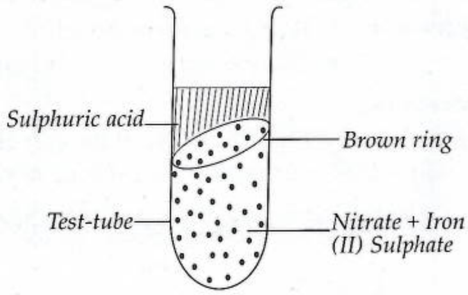


Figure 1:10 Brown ring test

2. Test for sulphates

You can test for sulphates using barium chloride solution or any barium compound solution. The reaction forms a white precipitate which does not dissolve even with the addition of hydrochloric acid. In *Activity 16*, you will test for sulphates in two unknown compounds.

Activity 16

Aim: *To test for sulphates*

Materials:
- *two unknown compounds (one of which is a sulphate)*
- *dropper*
- *barium chloride solution*
- *dilute hydrochloric acid*
- *two test-tubes*

Procedure:

1. *Half-fill the test tube with unknown solution*
2. *Add five drops of barium chloride solution and observe what happens.*
3. *Fill up the tube with dilute hydrochloric acid.*

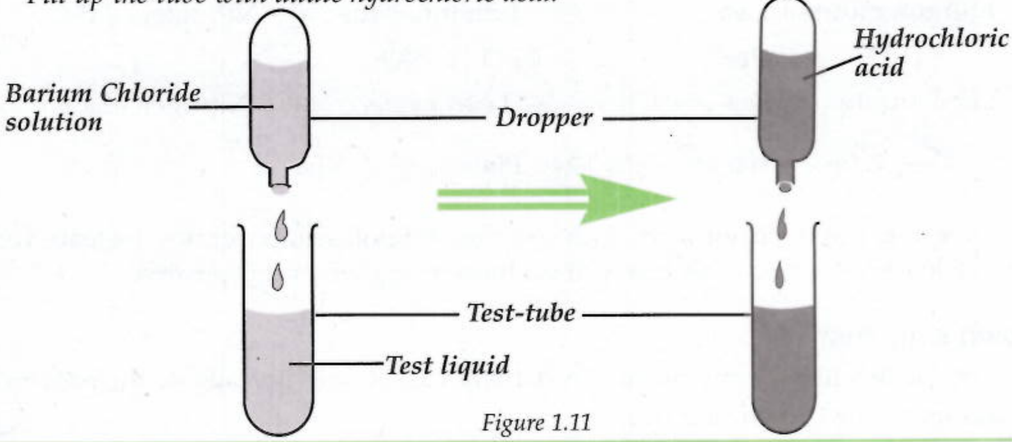


Figure 1.11

Results

1. *What did you observe when you added barium chloride solution to each of the two compounds?*
2. *Did any colour that you observed disappear with the addition of dilute hydrochloric acid?*

Conclusion/Discussion:

You have observed that one of the solutions forms a white precipitate when barium chloride solution is added. This precipitate is the salt of barium sulphate. You have also observed that the precipitate does not disappear with the addition of dilute hydrochloric acid. Look at its net equation:

$$\underset{\text{Sulphate ion}}{SO^{2-}} + \underset{\text{Barium ion}}{Ba^{2+}} \longrightarrow \underset{\text{Barium sulphate}}{BaSO_4}$$

Halides

Halides are compounds of Group 7 elements of the Peridical Table. They can either be metals, such as: copper chloride ($CuCl_2$), potassium bromide (KBr) and Silver iodide, or non-metals helides such as: hydrogen fluoride (HF) and hydrogen chloride (HCL).

Properties of halides

1. All aqueous solutions of halides, except fluorides, react with silver nitrate slutions [$AgNO_{3(aq)}$] and produce silver halide precipitates of different colours. Look at *Table 1.1*:

Solution added $AgNO_3(aq)$	$F^-(aq)$ No reaction	$Cl^-(aq)$ White ppte of AgCl	$Br^-(aq)$ Cream ppte of AgBr	$I^-(aq)$ Yellow ppte of AgI

Table 1:1

2. All silver halides precipitates, except silver iodide, dissolve in concentrated ammonia solution.
3. Except for fluorides and iodides, silver halides change colour when exposed to sunlight. See the *Table 1.2*:

Effect of sunlight	**AgF** No effect	**AgCl** White AgCl turns purple-grey	**AgBr** Cream AgBr turns green-yellow	**AgI** No effect

Table 1:2

Test for Helides

The distinctive colours of silver helides in *Table 1:1* can be used in analysis to identify an unknown halide ion.

Procedure:

1. *Add 2cm of unknown halide solution to the same volume of silve nitrate solution.*

Figure 1:12

2. *Check for the colour change.*
3. *Now identify the halide.*

Test for cations

Many cations form precipitates when sodium hydroxide or ammonia solution is added. Some of the precipitates formed disappear when excess reagent (*sodium hydroxide* or *ammonia*) is added but others do not. The precipitates formed may have different colours.

Test solution (cation)	Observations when few drops are added	Observations when excess reagent is added
Cu^{2+}	Blue precipitate	Precipitate does not disappear
Mg^{2+}	White precipitate	Precipitate does not disappear
Ca^{2+}	White precipitate	Precipitate does not disappear
NH_4^+	No precipitate	
Al^{3+}	White precipitate	Precipitate disappears
Fe^{2+}	Muddy green precipitate	Precipitate does not disappear
Li^+	No precipitate	
Pb^{2+}	White precipitate	Precipitate disappears
Fe^{3+}	Red-brown precipitate	Precipitate does not disappear
Zn^{2+}	White precipitate	Precipitate disappears

Table 1.3

In *Activity 17*, you will test for the presence of cations in the solutions.

Activity 17

Aim: *To test for the presence of cations*

Materials: 1. *Containers labelled* **A, B, C, ..., J** *containing solutions of:*

- *Copper sulphate* Cu^{2+}
- Ammonium nitrate NH_4^+
- *Lithium chloride* Li^+
- *Calcium nitrate* Ca^{2+}
- *Potassium nitrare* K^+
- *Zinc nitrate* Zn^{2+}
- *Aluminium sulphate* Al^{3+}
- *Magnesium sulphate* Mg^{2+}
- *Iron(11) sulphate* Fe^{2+}
- *Iron (111) nitrate* F^{3+}

2. *1M Sodium hydroxide solution*

Procedure:

1. *Pour each solution 1 cm deep in a separate test-tube*
2. *Add some drops of sodium hydroxide solution. Observe what happens.*
3. *If the precipitate forms, fill up the test tube with the reagent. Does the precipitate disappear?*

Results/observations:

1. *Which ions form precipitates when a reagent is added? Describe the colour of each precipitate formed?*
2. *Do the precipitates disappear when the excess reagent is added?*

Discussion

You have observed that some cations form precipitates while others do not. All group 1 metal and ammonium compounds do not form any precipitate when their solutions react with sodium hydroxide. Zn^{2+}, Al^{3+} and Pb^{2+} ions form white precipitates which disappear when excess reagent is added. Ca^{2+} and Mg^{2+} ions form white precipitates which do not disappear even when the reagent is added in excess.

Exercise 4

1. John carried out an experiment to identify K^+, Pb^{2+} and Mg^{2+} ions from the three unlabelled solutions.
   a) Name the reagent he would use to identify them.
   b) Describe the properties of each of the ions which he would use to identify them
2. a) What kind of ions are identified using Brown ring test?
   b) Name one important reagent used in Brown ring test.
3. a) Describe one property of
      i) carbon dioxide gas
      ii) hydrogen gas.
   b) What would you use to test for:
      i) oxygen gas?
      ii) water?

Unit summary

- Many types of chemical activities generate hazardous wastes which may be harmful to human beings and the environment.
- Chemical reactions you carry out in the laboratory can also generate wastes of different kinds. Safe ways of disposing waste substances from the laboratory must be put in place to prevent harmful exposure. Some of the ways are:
  i. wash the waste in the sink then through the drain,
  ii. incinerate the waste,
  iii. burying it underground or in landfills and
  iv. collect the waste in bins for the Local Authority to dispose.

  Note: *Methods of disposing waste will vary depending on how safe the waste may be.*

  v. Gaseous fumes are best removed from the laboratory through the fume hood.
- The best approach to laboratory waste management is preventing waste generation. If this cannot be achieved then consideration must be made to *reuse* and *recycle waste* before disposal.
- A pure substance is made up of one kind of particles and has fixed composition. You can identify it using its unique physical properties such as density, conductivity, melting and boiling or vapour pressure.
- To carry out a scientific investigation you have to plan and design your investigation sufficiently well following a systematic procedure.
- There are various analytical methods in chemistry, which help you to separate an analyte: identify and quantify it. These include filtration, distillation, titration and chromatography.
- You can test to see if a substance contains water by using cobalt chloride paper. In the presence of water anhydrous copper sulphate crystals turn from white to blue colour.

Unit Exercise

1. An athlete was accused of using the drug THG to enhance his running performance since his urine tested positive. How would you use chromatography to confirm this?

2. *Figure 1:13* shows a graphical presentation of a chromatogram. How many components did the mixture have? Which component is greatly adsorbed to the chromatographic paper

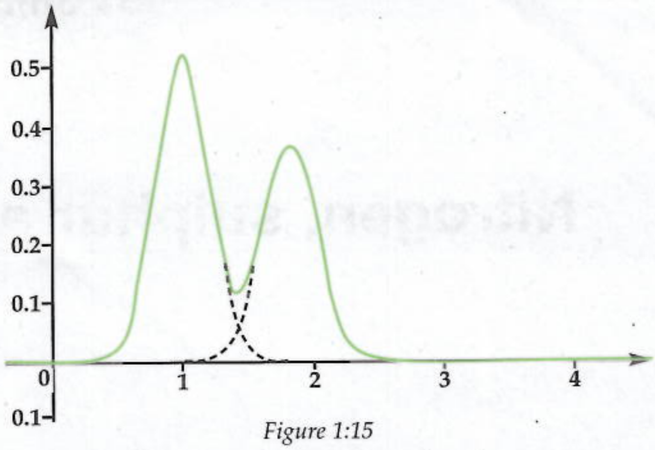


Figure 1:15

3. When a colourless liquid is strongly heated, it boils and forms a colourless vapour at 96.9°C. When the liquid is electrolysed it forms a colourless gas at the positive electrode, which is known to make fire burn brightly. At the negative electrode, a gas which burns with a popping sound is produced.
   a) Give the name of the liquid.
   b) What gases are produced at the positive and negative electrodes?
   c) Is this liquid pure? Explain.

4. *Mary and John had just been introduced to the lesson of electrolysis by their teacher, unfortunately had no sufficient time to continue with the lesson. They were told that an ionic compound when dissolved in water could conduct electricity. They tried to prove this on their on, but used different quantities of an ionic compound. The solutions conducted electricity but the bulbs lit differently. This surprised them and they decided to investigate if concentration of salt solution affects electrolysis.*

   In their investigation, state:
   a) *i.* The problem they could investigate
      ii. Two possible hypotheses
      iii. Any two factors to be controlled.
   b) explain any two ways they could use to collect their data.

5. Both carbon dioxide and hydrogen gases are colourless and ordourless. Describe an activity you would use to distinquish them.

6. a) Name the presipitate formed when hydroxide ions are added to each of the following cations
      i. copper (II) ions Cu^{2+}
      ii. calcium ions, Ca^{2+}
      iii. magnesium ions, Mg^{2+}
      iv. Aluminium ions, Al^{3+}
   b) State the colour of each precipitate formed in (a)

7. What test would you carry out to distinguish:
   a) paraffin from water
   b) ammonium sulphate from ammonium chloride solution
   c) oxygen from carbon doxide gas?

Unit 2

Nitrogen, sulphur and phosphorus

Success Criteria

By the end of this unit you should be able to:

1. describe sources of nitrogen
2. describe properties of nitrogen
3. explain uses of nitrogen and its compounds
4. describe sources and properties of sulphur
5. explain the uses of sulphur and its compounds
6. describe sources and properties of phosphorus
7. describe uses of phosphorus and its compounds

Key words:

In this unit you will find these key terms and concepts:

Haber process, Ostwald process, Contact process, α-sulphur, β-sulphur, upward delivery, tetrahedron shape

Ensure that you understand and learn how to apply them both for your academic and real life situations.

Nitrogen, sulphur and phosphorus are very important elements to human beings and the environment. They have several uses which include agricultural and industrial application. In this unit, you will study sources, properties and uses of these elements and their compounds.

Sources of Nitrogen

There are many sources of nitrogen. The sources are both natural and artificial through industrial preparations. *Figure 2.1* is a nitrogen cycle. It illustrates how nitrogen circulates in nature. Study it carefully and, in groups, answer the following questions. Report your answers to the whole class.

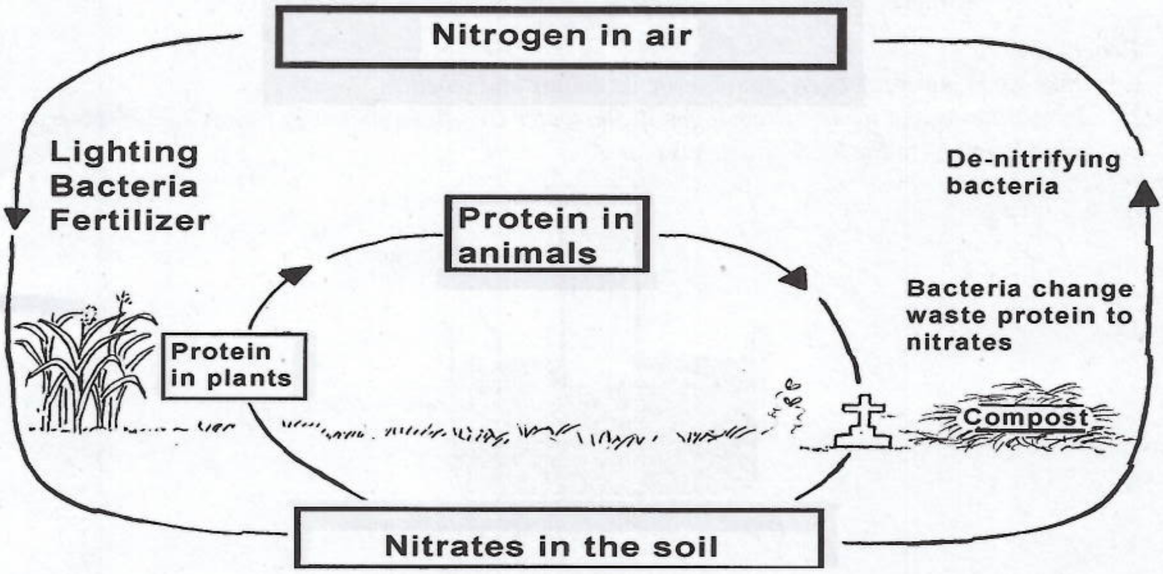


Fig. 2.1: Nitrogen cyrcle

Exercise 1

1. How many forms of nitrogen do you see in *Figure 2.1*? Mention them.
2. Where do these forms of nitrogen come from?
3. Mention any two materials that contain nitrogen.

Figure. 2.1 attempts to illustrate sources of nitrogen. In some of the sources, nitrogen exists as an element while in others as a compound. The following are the sources of nitrogen:.

1. *The Air:* A large percentage of atmospheric air is nitrogen gas. Its molecular formula is N_2. Nitrogen gas makes up 78% of the atmospheric air. It is around us. We breathe it in and out of our bodies without reacting with our blood.
2. *Soil nitrates:* The nitrogen cycle shows that some of the nitrogen in the air ends up in the soil where it forms nitrates. Nitrates are formed in three main ways:
   a) by *nitrogen-fixing bacteria* in the root modules of leguminous plants. Such bacteria fix nitrogen from the air into the soil.
   b) by l*ightning and thunder storm*: the lightinng and thunderstorm cause nitrogen to react with oxygen to become nitrogen dioxide. The nitrogen dioxide is washed by the rains into the soil where they form other compounds, including nitrates.
   c) application of *nitrogenous fertilizers* such as Urea and NPK, 23:21:0. This application supplies nitrogen into the soil, which react with oxygen in the soil to form nitrates.

Properties of Nitrogen gas

Like many other gases, nitrogen gas has both physical and chemical properties. In *Activity 1* you are going to investigate some physical and chemical properties of nitrogen gas.

Activity 1

Aim: *To investigate the physical and chemical properties of nitrogen gas.*

Materials:
- *2 wet litmus papers (blue and red)*
- *calcium hydroxide solution*
- *sulphur powder*
- *a pair of tongs*
- *wood splint*
- *a piece of magnesium ribbon*
- *6 gas jars of nitrogen gas*
- *trough of water*

Procedure:

1. *Take a gas jar of nitrogen and observe its colour and smell.*
2. *Invert the second jar of nitrogen gas in the water trough as shown in figure 2.3 (a), observe what happens to the level of water in jar A*

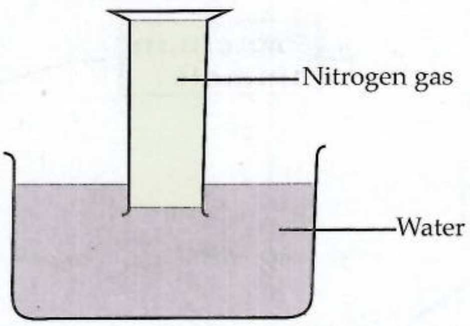


Figure 2.3 (a) Mixing nitrogen with the water

3. *Put a glowing splint in the third jar of nitrogen gas. Does the splint relight? Does the splint continue burning?*
4. *Put a burning splint in the same third jar of nitrogen. Does the splint continue burning?*

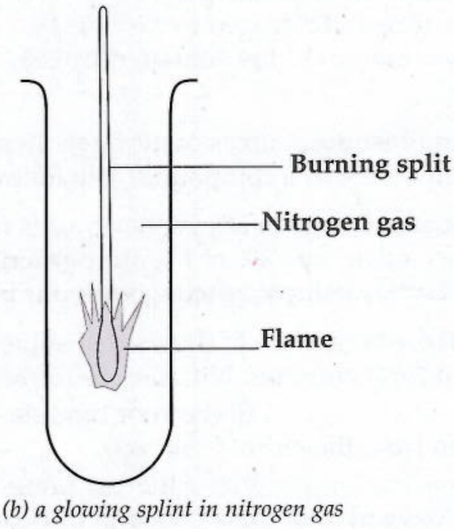


Figure 2.3 (b) a glowing splint in nitrogen gas

5. *Lower a burning magnesium ribbon on a deflagrating spoon in to the fourth gas jar of nitrogen gas as in figure 2.3 (c). Add water from a dropper onto the magnessium ash formed. Observe any gas which may be given off when water is added to the ash.*

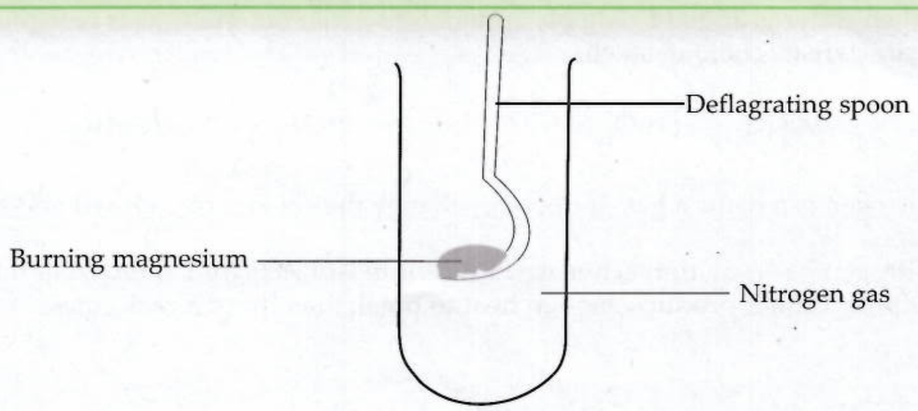


Figure 2.3 (c) magnesium burning in nitrogen gas

6. *Put wet blue and red litmus sheets of paper in the fifth jar. What colours do they become?*
7. *Lower burning sulphur on a deflagrating spoon in the sixth gas jar containing nitrogen gas as shown in Figure 2.3 (d). Do you observe what happens?*

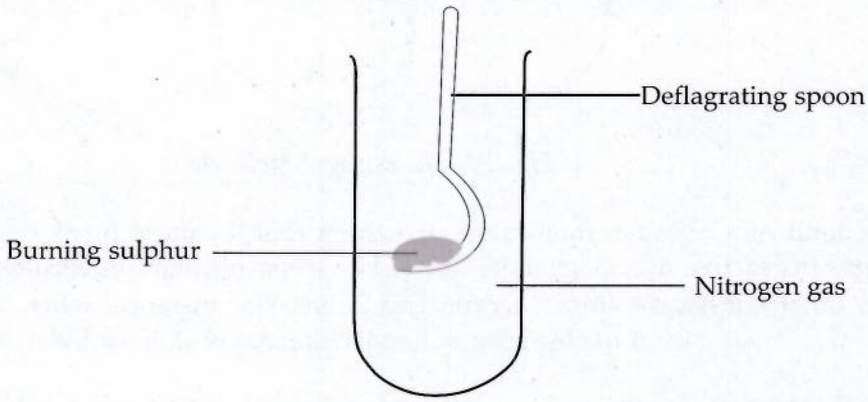


Figure 2.3 (d) burning sulphur in the nitrogen gas

Results and conclusion:

The observations you have made in this activity show both physical and chemical properties of nitrogen gas. Summarise the physical and chemical properties of nitrogen gas separately in a table.

Physical properties of nitrogen gas

1. Nitrogen is a colourless gas.
2. It has no smell (*odour*).
3. It is almost insoluble in water.

Chemical properties of nitrogen gas

1. Nitrogen gas neither burns in air nor does it support burning. Both, a glowing splint and a burning splint are extinguished when they are inserted into a gas jar of nitrogen.
2. When magnesium burns in nitrogen, it produces magnesium nitride:

$$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

When water is added to Mg_3N_2, ammonia is given off. This gas is recognized by its characteristic choking smell.

$$Mg_3N_{2(s)} + 6H_2O_{(l)} \longrightarrow 2NH_{3(g)} + 3Mg(OH)_{2(s)}$$

3. Nitrogen is a neutral gas. It does not change the colours of both red and blue litmus paper.
4. Nitrogen gas is an unreactive gas. Procedure 7 of *Activity 1* shows that the burning sulphur cannot produce enough heat to break the nitrogen molecules.

Unreactivity of nitrogen gas

The nitrogen gas molecule is made up of a pair of nitrogen atoms which are combined by three covalent bonds together forming a strong triple bond.

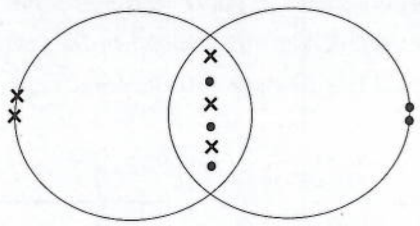


Fig. 2.4: The nitrogen molecule

The triple bond of a nitrogen molecule is so strong that it cannot break easily. This makes nitrogen gas unreactive at room temperature. However, nitrogen molecules can break and react with other substances under certain conditions. For instance, when there is a large amount of heat energy from the burning of metals like magnesium or from an electric spark.

Unlike oxygen that reacts with many metals to form oxides, nitrogen can only react with very reactive metals such as those of Groups 1 and 2 of the Periodic Table.

Exercise 2

1. State two physical properties of nitrogen gas.
2. (a) Why is nitrogen an unreactive gas?
   (b) State one condition under which nitrogen can react.

Uses of nitrogen gas

Nitrogen has several important uses. These are:

1. *Making ammonia (NH_3)*: The main use of nitrogen gas is for manufacturing ammonia, NH_3 through the Haber process. Some of this ammonia is used for making fertilizers.
2. *Freezing things:* Nitrogen boils at -196°C. So below this temperature it is a liquid. Liquid nitrogen is very cold. This makes it useful for:
   a. freezing and keeping foods frozen during transportation.
   b. storing bulls' semen for artificial insemination.
   c. in engineering with liquids. For example, it is used to freeze liquids in damaged pipes before repairing them.

3. *Providing inert atmosphere in oil storage tanks:* Nitrogen gas is pumped into spaces of oil storage tanks to prevent fires. This enables ships to transport crude oil over long distances without fire accidents.

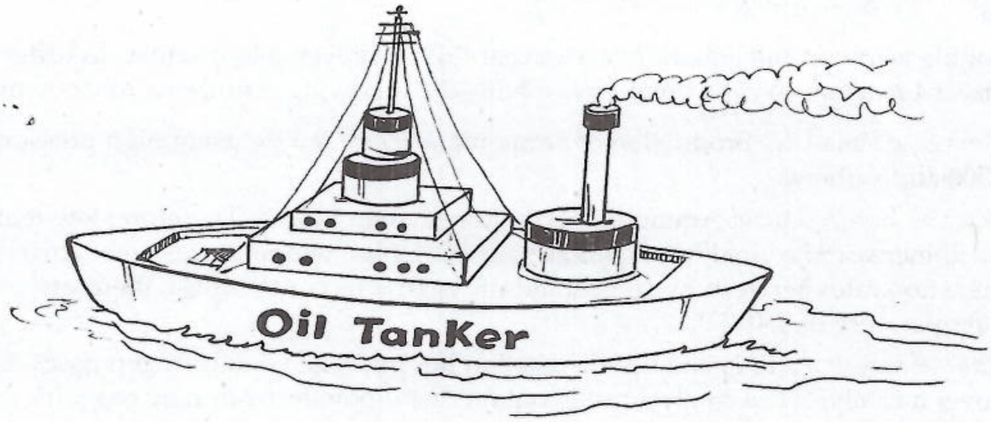


Fig. 2.5: Oil tanker (ship)

4. *Food packaging:* Nitrogen gas is also pumped into spaces of food packages to remove oxygen. This helps to keep food fresh as most bacteria cannot survive without oxygen.

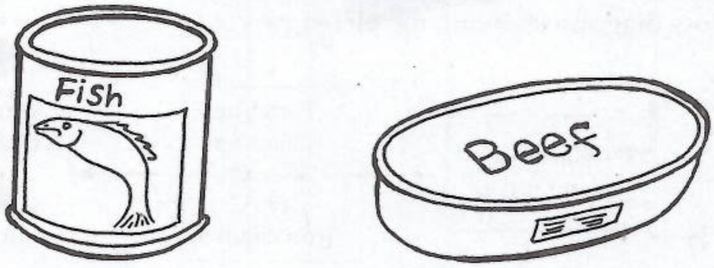


Fig. 2.6: Food packages (canned food)

Exercise 3

1. State three uses of nitrogen.
2. Explain how you can use nitrogen to:
   a) prevent fires
   b) preserve food in plastic containers.
3. Give one example of fertilizer made from ammonia.

Nitrogen compounds

Ammonia and nitric acid are some of the compounds of nitrogen. They have very important uses.

1. Ammonia

Ammonia is a compound of nitrogen and hydrogen. It has the following properties:

a. It is a gas at room temperature.
b. It has a pungent smell and it can be felt in urinaries and animal houses.
c. It is lighter than air
d. It readily dissolves in water to form a basic solution.

Industrial preparation of ammonia – the Haber process

Ammonia has the formula, NH_3. It can be made by reacting nitrogen with hydrogen gases:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The double arrow in the equation shows that this is a reversible reaction. In order to favour the forward reaction so as to obtain more NH_3, the following conditions must be provided:

1. Increase Pressure: Production of ammonia is increased by using high pressure up to 300 atmospheres.
2. Lower Temperature: Ammonia decomposes when heated. Therefore, low temperature will increase the yield of ammonia. However, high temperatures will increase the reaction rates between hydrogen and nitrogen. The compromise, therefore, is to use a temperature of 450°C.
3. Use of Catalyst: To speed up the reaction hot hydrogen and nitrogen gases are passed over a catalyst. The catalyst used contain finely powdered iron mixed with potassium and aluminium oxides.

Ammonia is manufactured in industries by a process called the Haber process: since it was developed by a German chemist, Fritz Haber.

The Haber process

Figure 2.7 is the flow diagram showing the Haber process.

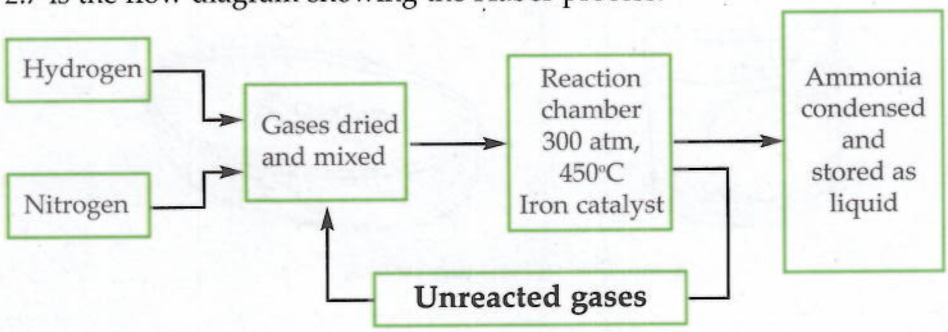


Fig. 2.7: The Haber process

How many stages does the Haber process have? The flow diagram of the Haber process shows that the process has four stages:

1st Stage: Obtaining nitrogen and hydrogen gases: Nitrogen is obtained by the fractional distillation of air. Hydrogen is obtained by the reaction of natural gas - methane, CH_4, with steam. Study the equation below:

$$\underset{\text{Methane}}{CH_4(g)} + \underset{\text{steam}}{H_2O(g)} \longrightarrow \underset{\text{Hydrogen}}{3H_2(g)} + \underset{\text{carbon monoxide}}{CO(g)}$$

Another source of hydrogen gas is the electrolysis of water. This is done when electricity is cheap.

2nd Stage: Hydrogen and nitrogen gases are dried and mixed in the ratio 3:1.

3rd Stage: Reaction of nitrogen and hydrogen gases: The mixture of nitrogen and hydrogen gases is compressed to 300 *atm* pressure and 450°C heat. The mixture is then passed over the iron catalyst. In this stage some gases react to form ammonia.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

nitrogen hydrogen ammonia

In the reaction, only 10% of the gases react to form ammonia.

4th Stage: Cooling the mixture: The mixture is cooled to liquefy the ammonia and obtain it. The unreacted gases are returned into the reaction chamber to re-start the process.

Laboratory preparation of Ammonia

Ammonia gas can be produced by **heating any ammonium salt with an alkali** such as calcium hydroxide. In *Activity* 2, you will prepare ammonia gas.

Activity 2

Aim: *Preparing ammonia gas in the laboratory.*

Materials:
- *gas jar*
- *large test tube*
- *ammonium chloride*
- *wet red litmus paper*
- *delivery tube*
- *calcium oxide*
- *calcium hydroxide*
- *holed cork stopper*
- *glass wool*
- *mortar and pestle*

Procedure:

1. *Grind the mixture of calcium hydroxide and ammonium chloride and put it in the large test tube.*

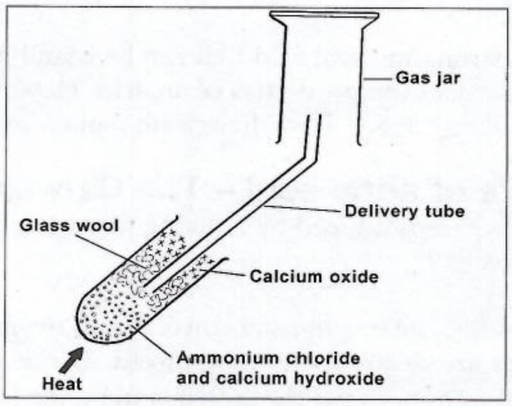


Figure 2.8 Preparing ammonia gas

3. *Heat the test tube gently. What do you observe in the delivery tube?*
4. *Put the moist red litmus paper close to the mouth of the gas jar. What colour does the red litmus paper become? Why is the gas jar positioned upside down above the delivery tube?*

Discussion:

As the mixture is heated, ammonium chloride and calcium hydroxide will react to produce calcium chloride, water and ammonia gas:

$$Ca(OH)_2(s) + 2NH_4Cl(s) \longrightarrow CaCl_2(s) + 2NH_3(g) + H_2O(l)$$

Calcium oxide is used as a **drying agent** because the usual drying agents, such as sulphuric acid and calcium chloride, react with ammonia gas.

The collecting gas jar is positioned upside down over the delivery tube because **ammonia is less dense than air** and can only be collected this way. This method of gas collection is called ***upward delivery***.

Uses of ammonia

Ammonia has many uses. The following are some of them:

1. *Making nitrogenous fertilizer:* Mostly ammonia is used for making nitrogenous fertilizers, e.g. urea.
2. *Making industrial compounds:* Ammonia is also used for making important industrial compounds such as ammonium chloride, used in dry cells; nitric acid used for making fertilizers and sodium carbonate.
3. *Refrigerant:* Liquid ammonia is used as a refrigerant for cooling large plants and factories.
4. *Household cleaners:* Ammonia is also used for making household cleaners since ammonia, like many alkalis, attacks grease.
5. Used for the *manufacture of drugs, dyes, plastics and fibers* such as nylon and terylene.

Exercise 4

1. Mention three measures that are employed in Haber process to manufacture as much ammonia as possible.
2. State three uses of ammonia.

2. Nitric acid

Nitric acid is a strong mineral acid that can be found in school laboratories. As a dilute acid, it shows most of the properties of an acid. However, in its concentrated form it is a powerful oxidizing agent. It oxidizes both metals and non-metals.

Manufacture of nitric acid – The Ostwald process

Nitric acid (HNO_3) is produced by reacting ammonia gas with oxygen over a catalyst. This process is called Ostwald process.

Ammonia from the Haber process is mixed with oxygen from the factional distillation of air. The gases are cleansed and then mixed. The mixture is then compressed to a pressure of 8 to 10 atm. In the reaction chamber, the mixture passes over a catalyst of platinum and rhodium at a temperature of 900°C. Under these conditions, the gases start to react. The reaction produces nitrogen dioxide and water as follows:

$$\text{ammonia} + \text{oxygen} \longrightarrow \text{nitrogen dioxide} + \text{water}$$

$$4NH_3 + 7O_2 \longrightarrow 4NO_2 + 6H_2O$$

The products of this reactions are then mixed with air and passed up through a tower. The nitrogen dioxide gas reacts with water flowing downwards from the tower to produce dilute nitric acid.

More concentrated acid can be produced by distillation of the dilute nitric acid.

$$4NO_2 + 2H_2O \longrightarrow 4HNO_3$$

Figure 2.9 shows the flow diagram of the process for manufacturing nitric acid.

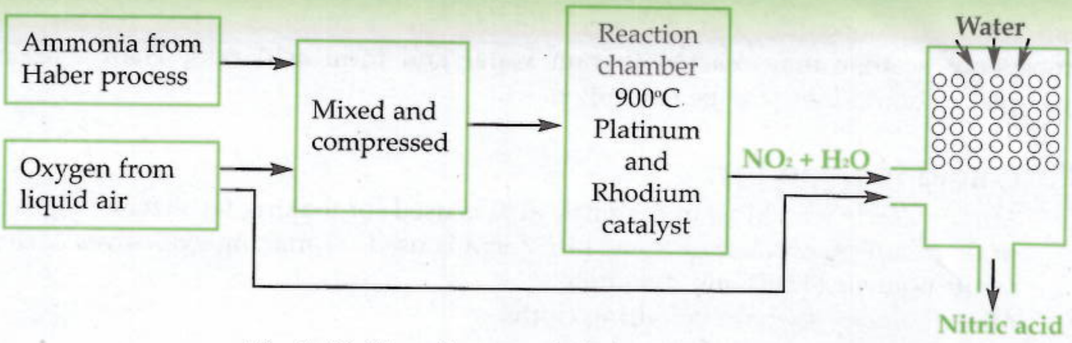


Fig. 2.10: Flow diagram of nitric acid production

You can also prepare the nitric acid in the laboratory. In *Activity 3* you will learn how to prepare nitric acid.

Activity 3

Aim: *To prepare nitric acid in the laboratory.*

Materials:
- *glass retort*
- *gas burner*
- *wire gauze*
- *round-bottomed flask*
- *potassium nitrate*
- *sand in watch-glass*
- *tripod stand*
- *concentrated sulphuric acid*
- *cold running water*

Procedure:

1. *Put some potassium nitrate and concentrated sulphuric acid in the glass retort*
2. *Set up the apparatus as in the diagram below:*

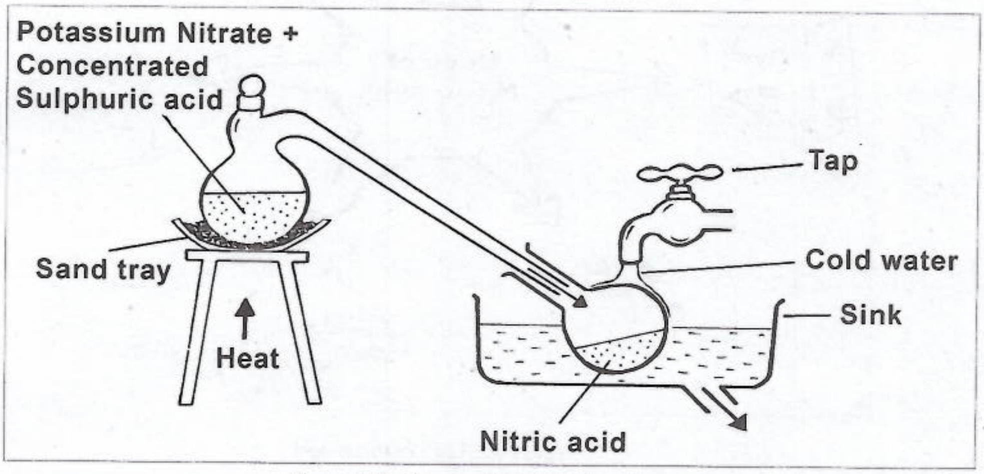


Figure 2.11: Preparation of nitric acid

3. *Heat the mixture gently and collect the acid produced in the water-cooled receiver. Why is the cold water used to cool the round-bottomed flask? What is the use of the sand.*

The equation of the reaction for the formation of nitric acid in the above activity is as follows

$$\text{Nitrogen dioxide} + \text{water} \longrightarrow \text{nitric acid} + \text{nitrogen monoxide}$$

$$3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(g) + NO(g)$$

Both nitrogen monoxide and nitrogen dioxide are a problem when released into the atmosphere as they may react with rain water and form acid rain. Their release to the atmosphere should therefore be controlled.

Nitric acid is used for:

1. *Making fertilizers:* Most of the nitric acid is used for making fertilizers.
2. *Manufacturing explosives:* Some nitric acid is used for making explosives such as trinitrotoluene (TNT) and dynamite.
3. *Manufacturing dyes* for colouring cloths.
4. *Manufacturing nylon* and *terylene*.
5. *Manufacturing drugs*.
6. *Identification of gold:* Nitric acid, being a powerful oxidizing agent, will react with many metals such as copper, iron and bronze. The acid, however, does not react with gold. This is one of the ways gold can be identified.

Uses of nitric acid

Figure 2.12 provides a summary of the uses of nitric acid

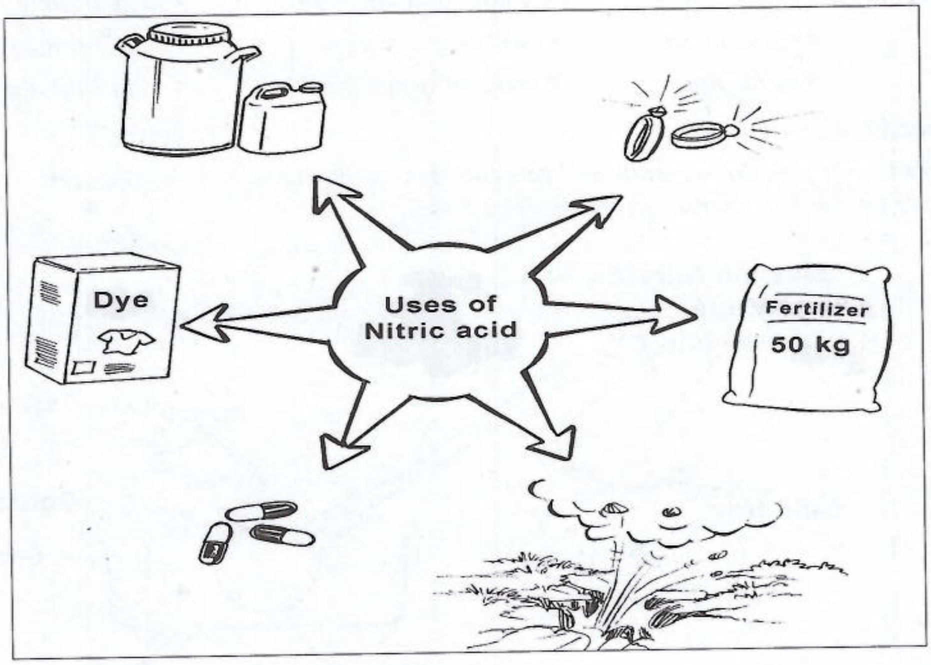


Fig. 2.12: Uses of nitric acid

Exercise 5

1. Name two raw materials used for manufacturing nitric acid.
2. State any three uses of nitric acid.

Sulphur

Sulphur is a very important element in the chemical industry. It is found in various sources and different forms. Its compound, sulphuric acid, has also several industrial uses.

Sources of sulphur

Sulphur can be found in different forms and these have different sources. The following are some of the sources of sulphur.

1. *Underground beds:* Most sulphur is found as an element in underground beds in some countries such as Poland, Mexico and the USA.
2. *Crude oil and natural gas:* Some sulphur occurs as hydrogen sulphide gas (H_2S) in crude oil and natural gas. When oil or gas is burned, the sulphur dioxide gas (SO_2) is produced. Sulphur dioxide causes air pollution.
3. *Metal ores:* Some sulphur is found combined with metals in ores. These metal ores are called *pyrites*. Some examples of pyrites are: copper ($CUFS_2$) and iron pyrites (FeS).

Extraction of Sulphur

1. *From underground deposits:* The Frasch process is used to extract sulphur from underground deposits. The process uses superheated water pumped down through one pipe to melt the sulphur. The hot compressed air is then pumped underground to push the molten sulphur upwards through a different pipe. Three pipes in all, arranged concentrically, are used. The molten sulphur is then cooled and solidified. *Figure 2.13* shows the diagram of the frasch process.

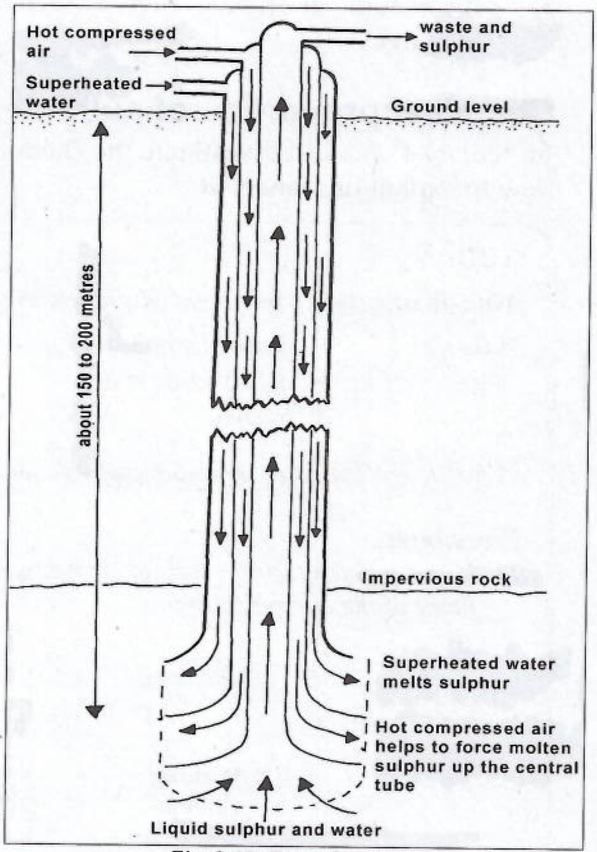


Fig. 2.13: Fransch process

2. *From crude oil:* Crude oil contains a mixture of oil gases such as *hydrogen sulphide* (H_2S), *carbon dioxide* (CO_2) and *methane* (CH_4). To separate these components crude oil is passed through an alkaline solution. Hence *hydrogen sulphide* and *carbon dioxide*, being acidic, are absorbed into the solution. They are regenerated by heating the solution. The hydrogen sulphide gas produced is later oxidized to sulphur by air.

$$2H_2S_{(g)} + O_{2(g)} \longrightarrow 2S_{(s)} + 2H_2O$$

Exercise 6

1. State any two sources of sulphur.
2. How is sulphur obtained from underground deposits? Explain.

Physical properties of sulphur

1. It is a brittle yellow non-metal solid.
2. It is made up of crown-shaped molecules with eight atoms each:

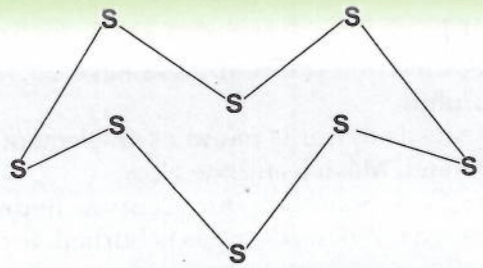


Figure 2.14: Crown-shaped molecule of sulphur

3. It is molecular and has a low melting point, 115°C.
4. It does not conduct electricity.
5. It is soluble in organic solvents such as benzene and toluene, but it does not dissolve in water.

Chemical properties of sulphur

In *Activity 4* you will investigate the chemical properties of sulphur in order for you to learn how to exploit or manage it.

Activity 4

Aim: *To investigate the chemical properties of sulphur.*

Materials:

- *crushed sulphur*
- *deflagrating spoon*
- *burner*
- *sulphuric acid*
- *gas jar*
- *test tube*
- *nitric acid*
- *wet blue + red litmus paper*
- *crucible*

Warning: Nitric acid and sulphuric acids can corrode the skin. Use protective clothing.

Procedure:

1. *Lower a deflagrating spoon of burning sulphur into a gas jar of oxygen. What colour is the flame of the burning sulphur?*

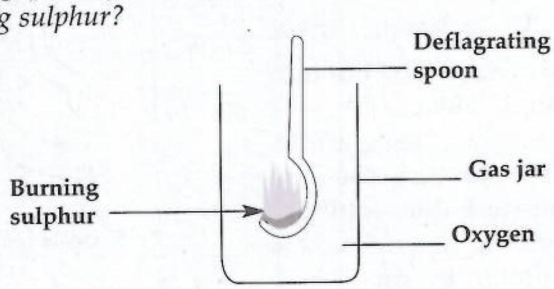


Fig. 2.15: Burning sulphur in air

2. *Test the gas produced with wet blue and red litmus paper. What colours do they become?*
3. *Mix the sulphur powder with some iron filings. Heat the mixture in a crucible. Repeat the experiment with a mixture of fine copper foil with sulphur. What do you observe?*
4. *Place some powdered sulphur and concentrated sulphuric acid in a test tube and warm it gently. Record your observation.*
5. *Repeat the experiment using the powdered sulphur and nitric acid. What do you observe?*

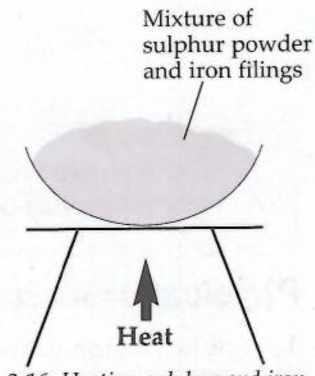


Fig. 2.16: Heating sulphur and iron

Results (chemical properties of sulphur)

1. Sulphur *burns easily in air with a blue flame* to produce acidic sulphur dioxide gas.

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

2. Sulphur reacts directly with metals to produce metallic sulphides. For example, an iron and sulphur mixture heated together produces iron (II) sulphide. A mixture of copper and sulphur when heated gives copper I sulphide.

$$Fe_{(s)} + S_{(s)} \longrightarrow FeS_{(s)}$$
$$2Cu_{(s)} + S_{(s)} \longrightarrow Cu_2S_{(s)}$$

3. When heated with sulphuric acid, sulphur is oxidized to sulphur dioxide.

$$S_{(s)} + 2H_2SO_{4(aq)} + \text{Heat} \longrightarrow 2H_2O_{(l)} + 3SO_{2(g)}$$

4. When sulphur is heated with nitric acid, sulphuric acid, water and nitrogen dioxide are produced.

$$S_{(s)} + HNO_{3(aq)} \longrightarrow H_2SO_4 + H_2O + NO_2$$

Uses of Sulphur

The following are some of the uses of sulphur:

1. Mostly for manufacturing sulphuric acid
2. For vulcanizing rubber. Ordinary rubber is soft but becomes tough, hard and elastic when heated with sulphur. This makes it suitable for making tyres.
3. For making matches, gunpowder, drugs, pesticides and paper.
4. For manufacturing of plastic flowers.
5. For manufacturing special concrete called sulphur concrete. This kind of concrete cannot be attacked by acid, so it is used for floors and walls where acid can get spilled.

Exercise 7

1. Mention three uses of sulphur.
2. What is vulcanised rubber?
3. State two: (*a*) physical properties and (*b*) chemical properties of sulphur.
4. Cu_2S is one of the metal sulphides. Describe, by using a chemical equation, how it is formed from copper, Cu and Sulphur, S.

Sulphuric acid

Sulphuric acid is a very useful compound of sulphur. It is manufactured in larger quantities than any other product. It is an oily colourless liquid. It forms a very strong acid when it is added to water.

Warning:

i. *Both, teachers and students, must take great care when using sulphuric acid. It can cause severe burns if it touches the skin.*

ii. *Always wear protective clothing and gloves when working with it.*

Manufacture of sulphuric acid

Sulphuric acid (H_2SO_4). It is made through a process called contact process. This process has four stages, namely:

Stage 1: Production of sulphur dioxide (SO_2)

Sulphur dioxide is made by burning elemental sulphur in air, though it is convenient to simply burn iron pyrites (FeS_2) in air.

$$\underset{\text{sulphur}}{S_{(s)}} + \underset{\text{oxygen}}{O_{2(g)}} \longrightarrow \underset{\text{sulphur dioxide}}{SO_{2(g)}}$$

$$\underset{\text{Iron pyrites}}{4FeS_{2(s)}} + \underset{\text{oxygen}}{11O_{2(g)}} \longrightarrow \underset{\text{iron oxide}}{2Fe_2O_{3(s)}} + \underset{\text{sulphur dioxide}}{8SO_{2(g)}}$$

Stage 2: Production of sulphur trioxide (SO_3)

Oxygen from the air and sulphur dioxide are mixed, cleaned and dried. The *cleaning process* removes dust and impurities from the mixture. They are then heated to a temperature of 450°C. The gases are then passed over a catalyst of vanadium oxide to speed up the reaction.

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$$

The reaction is a reversible reaction. So to increase the yield of sulphur trioxide, the gases are passed over several layers of the catalyst at very high temperature. Under these conditions, about 96% of sulphur dioxide and oxygen are converted into sulphur trioxide.

Stage 3: Absorption

Sulphur trioxide is then cooled and dissolved in concentrated sulphuric acid to produce a fuming sulphuric acid called ***oleum***. It is this oleum that is transported in tankers to wherever it is needed.

$$\underset{\text{sulphur trioxide}}{SO_3} + \underset{\text{sulphuric acid}}{H_2SO_4} \longrightarrow \underset{\text{Oleum}}{H_2S_2O_7}$$

Note: *Sulphur trioxide can not be directly added to water to form acid because the heat of reaction would cause a mist of sulphuric acid which will be difficult to condense.*

Stage 4: Dilution

When sulphuric acid is needed, oleum is diluted by adding it to water.

oleum + water ⟶ sulphuric acid

Diluting sulphuric acid must be done with care because of its affinity for water. The concentrated sulphuric acid should always be added to water.

Note: *Always **add acid to water** not the other way round.*

The contact process is illustrated in the flow diagram as in *Figure 2.18*.

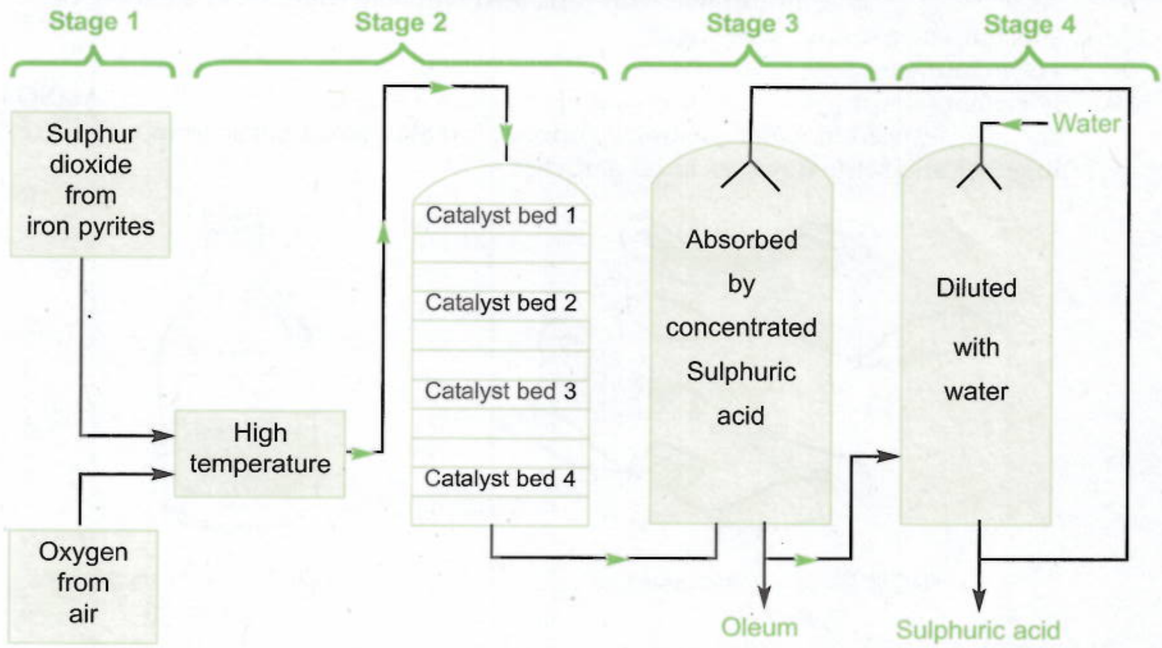


Fig. 2.18: Contact process

Uses of sulphuric acid

The following are some of the uses of sulphuric acid:

1. For manufacturing *fertilizers*: most sulphuric acid is used for manufacturing fertilizers
2. For manufacturing *paints, plastics* and *paper*
3. For manufacturing *dyes, soap* and *detergents*

Fig. 2.19: Detergents

4. As an *electrolyte* in car batteries
5. For *extraction* and *treatment of metals*
6. For manufacturing *insecticides* and *drugs*

7. For *refining crude petroleum*
8. As a *drying agent*: concentrated sulphuric acid will take water from a variety of substances including cane sugar.
9. For manufacturing *explosives*.
10. For manufacturing *synthetic fibre* and *flocculents* such as *aluminium sulphate* $Al_2(SO_4)_3$. A *flocculent* is used in water treatment process. It makes small particles in water to come together and settle down as large particles.

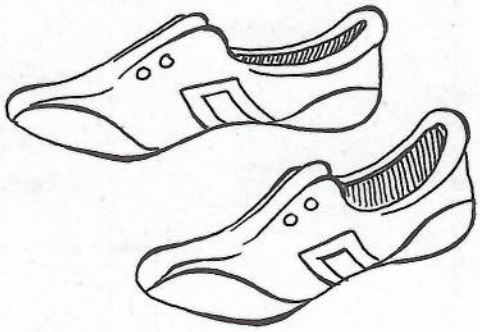


Fig 2.20 (**a**): *Sports shoes made from synthetic fibre*

Fig. 2.20 (**b**): *Aluminium sulphate* (a flocculent)

Exercise 8

1. State four uses of sulphuric acid.
2. How many stages does the contact process have?
3. Describe what is done in the first and second stages of the contact process?
4. Sulphur trioxide is the main product of the 2nd stage of the contact process. Give a balanced chemical equation that shows how sulphur trioxide is formed.
5. Sulphuric acid can be formed by directly adding sulphur trioxide to water. Why is this method of preparing sulphuric acid not a convenient one?

Sulphates

Other important compounds of sulphur are sulphates such as *iron sulphate*, ($FeSO_4$), *copper sulphate*, ($CuSO_4$) and *magnesium sulphate* ($MgSO_4$). You may have these sulphates in your school laboratory either in ***anhydrate*** or ***hydrate form***.

Sulphates are formed in three ways:

1. By *neutralizing dilute sulphuric acid* with an alkali or a carbonate:

sodium hydroxide + sulphuric acid → sodium sulphate + water

$$NaOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow Na_2SO_4 + H_2O_{(l)}$$

2. By *the action of concentrated* or *dilute sulphuric acid on appropriate metals*:

copper metal + sulphuric aid → copper sulphate + water + sulphur dioxide

$$Cu_{(s)} + 2H_2SO_{4(l)} \longrightarrow CuSO_{4(s)} + 2H_2O + SO_{2(g)}$$

3. By *double decomposition of soluble salts*:
The insoluble sulphate can be prepared by *reacting together two soluble salts*, *one of the two being a sulphate*, in a double decomposition process:

lead nitrate + sodium sulphate ⟶ lead sulphate + sodium nitrate

$Pb(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + NaNO_3(aq)$

Uses of Sulphates

The following are some uses of sulphates:

1. For making *fertilizers*, e.g. ammonium sulphate, $(NH_4 \bullet 2SO_4$.
2. For making *Plaster of Paris* (POP), e.g. *gypsum* $(CaSO_4 \bullet 22H_2O)$: used in hospitals for making plaster casts for setting broken bones.
3. For making *stomach anti-acid tablets*, e.g. *epsom salt* $(MgSO_4 \bullet H_2O)$.
4. In medicine for *eye ointments*, *mouth washers*, e.g. *hydrate zinc sulphate* $(ZnSO_4 \bullet 7H_2O)$: also used for healing various wounds.
5. As *fungicides* in orchards and *dyes*, e.g. *hydrate copper sulphate* $(CuSO_4 \bullet 5H_2O)$.
6. For making *white paints* and *diagonistic* ***X-ray*** *studies*, e.g. *Barium Sulphate* $(BaSO_4)$.

Exercise 9

1. Complete and balance the following chemical equations:
   (a) $Zn(s) + H_2SO_4(aq) \rightarrow$
   (b) $Ba(NO_3)_3(aq) + Na_2SO_4(aq) \rightarrow$
2. State one use of sulphates in
   (a) the agricultural industry
   (b) medicine

Phosphorus

Phosphorus is a non-metal element. It is from a Greek word meaning *light bearer*. The reason being that, before it was fully discovered, it was described as a white pasty substance that glowed at night and could easily catch fire.

Sources of Phosphorus

Phosphorus has various sources, such as:

1. in *banana peels*
2. in *crab shells*
3. in *rocks*
4. in *coal*

Uses of Phosphorus

Phosphorus is used for making:

1. match heads
2. detergents, eg. surf
3. fertilizers
4. rat poison
5. bombs

Physical properties of phosphorus

1. It is a *non-metal*
2. It is found in two forms (*allotropes*): *yellow* and *white*

3. It has a low melting and boiling points
4. It is composed of molecules. Each molecule of phosphorus is composed of four atoms joined together in a tetrahedron shape.

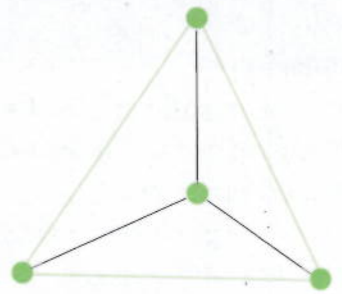


Fig. 2.22: P_4 molecule of phosphorus

Chemical properties of phosphorus

1. It burns very readily in oxygen to give a white bright flame. This combustion of phosphorus in oxygen produces phosphorus (V) oxide which dissolves in water to form phosphoric acid.

$$\underset{\text{Phosphorus}}{P_4(s)} + \underset{\text{oxygen}}{5O_2(g)} \longrightarrow \underset{\text{phosphorus pentoxide}}{2P_2O_5(s)}$$

$$\underset{\text{Phosporus oxide}}{P_2O_3(s)} + \underset{\text{water}}{3H_2O(l)} \longrightarrow \underset{\text{phosphoric acid}}{2H_3PO_4}$$

2. Phosphorus burns in chlorine gas to form phosphorus trichloride (PCl_3) [or phosphorus pentachloride (PCl_5) when excess chlorine is supplied].

$$P_4(s) + 10Cl_2(g) \longrightarrow 4PCl_5(l)$$

3. It does not react with water. Water is, therefore, a good medium for storing it; to prevent it from ignition with air.

Exercise 10

1. State two uses of phosphorus.
2. State two physical properties of phosphorus.
3. Describe combustion of phosphorus. Give the chemical equation of this reaction.

Unit summary

- Nitrogen is naturally found in air, soil nitrates, and plant and animal protein.
- Nitrogen is a colourless and odourless gas. It is also an unreactive gas. It does not burn in air nor does it support burning.
- Nitrogen gas is used for making ammonia (NH_3). Nitrogen is also used for freezing things.

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- Ammonia is a gas at room temperature; a compound of nitrogen and hydrogen.
- Ammonia is manufactured through a process called the Haber process.
- Ammonia is mainly used for making fertilizers. Some ammonia is also used for making drugs, household cleaners and for refrigeration.
- Nitric acid is made from ammonia from the Haber process and oxygen from the distillation of air. The process of making nitric acid is called Ostwald process.
- Nitric acid is used for making drugs, plastics, dyes and fertilizers.

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- Most sulphur is found as an element in underground beds in Poland, Mexico and USA. It is also found in crude oil and natural gas.
- Sulphur is a yellow non-metal solid, made up of crown-shaped molecules having eight atoms each.
- Sulphur burns easily in air with a blue flame and forms an acidic sulphur dioxide gas, SO_2.
- Sulphur reacts directly with metals to form metallic sulphides, e.g. iron II sulphide (FeS) and copper I sulphide (Cu_2S).
- Sulphur has several uses. It is used for making sulphuric acid, fungicides, matches, fireworks and vulcanized rubber.
- Sulphuric acid is the most produced industrial product in the world. It is produced from sulphur dioxide (SO_2) and oxygen (O_2) gases in a process called contact process.
- Sulphuric acid is used for making fertilizers, insecticides, drugs and paints.

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- Phosphous is a reactive non-metal. It is found in two forms: white and yellow.
- Phosphorus consists of molecules which are made up of four (4) atoms joined in tetrahedron shape.
- Phosphorus easily ignites in air to form phosphoric pentoxide. It is best stored in water.
- Phosphorus is used for making match-heads, rat poison, fertilizers and detergents.

Unit exercise

1. To obtain dry nitrogen from air, sulpuric acid and sodium hydroxide are both used and added to the test tube:
   a) What is the purpose of sulphuric acid?
   b) What is the purpose of sodium hydroxide?
   c) Write an equation for the reaction in the test tube.

2. a) State and explain what you would observe when burning magnesium is lowered into a gas jar full of nitrogen gas and water is poured onto the ashes formed.
   b) Write a balanced equation for the reaction in (a).
   c) State one use of nitrogen gas in the agriculture industry.

3. a) i) Draw the molecule of ammonia, NH_3, and ammonium ion, NH_4^+.
   ii) How do the structures you have drawn in (a) (i) differ?
   iii) What is the name of covalent bond found in the ammonium ion?
   b) When preparing ammonia gas in the laboratory, the common drying agent (sulphuric acid) is not used.
   i) Why is it not used?
   ii) Name the drying agent that is used.

4. What conditions give economic yield in:
   (a) Haber process?
   (b) the contact process?

5. The graph in *Fig. 2.27* shows the results for some experimental studies on the yield of ammonia in the Haber process. Use the graph to answer the questions that follow.

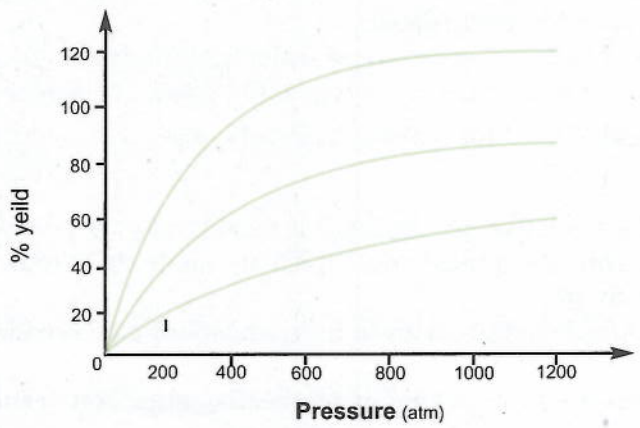


Fig. 2.23 Graph

   a) Under what conditions would a yield of 90% be achieveable?
   b) What percentage yield is obtained at the normal operating condition of 500°C and 200atm?
   c) In some modern plants a pressure of 1000 atm is used. What effect does this have on the yield at:
   i) 500°C?
   ii) 200°C?

6. a) Describe the Frasch process.
   b) What physical attributes of sulphur allows it to be extracted using superheated water.

7. In the contact process, sulphur dioxide and oxygen react in the presence of a catalyst to form compound X.
   a) i) What is a catalyst?
      ii) Name the catalyst used in the contact process.
   b) What is compound X?
   c) What is formed when compound X is dissolved in water?

8. Mention two sources of phosphorus.

9. *Fig 2.28* shows the combustion of phosphorus in oxygen. The wet blue litmus paper is put on the mouth of the gas jar to test the fumes produced from this combustion.

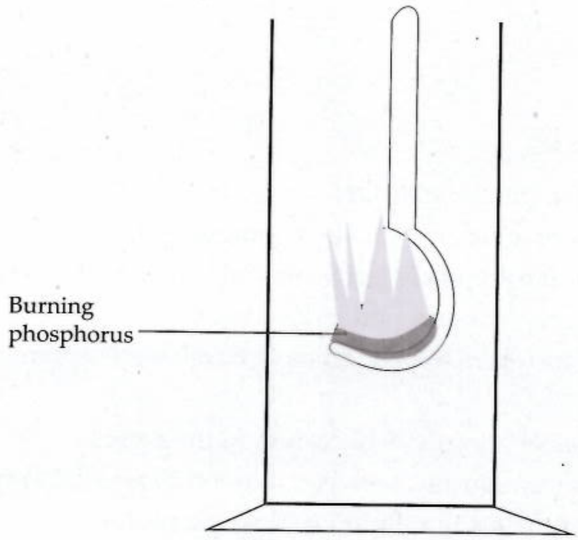


Fig. 2.24: Burning phosphorus

   a) Describe the colour of the flame of the burning phosphorus.
   b) Give the balanced chemical equation of the combustion of phosphorus in air.
   c) Describe what would become of the litmus paper in the experiment. Give a reason.
   d) Why is phosphorus stored in water?

Unit 3

Chemical Bonding and Properties of Matter

Success Criteria

By the end of this unit you should be able to:

1. Describe properties of ionic and covalent compounds
2. Explain the structural differences between ionic and covalent compounds
3. Differentiate between polar and non polar bonds
4. Relate intermolecular forces to properties of covalent compounds
5. Define allotropy
6. Relate the properties of allotropes of carbon to their uses
7. Explain the physical similarities between diamond and silicon oxide
8. Describe the uses of metals in relation to their properties
9. Explain the physical properties of alloys
10. State uses of alloys

Key words:

In this unit you will find these key terms and concepts:

bonding, stable configuration, ionic bond, covalent bonding, valence dot and cross structure, polar bond, hydrogen bonding

Ensure that you understand and learn how to apply them both for your academic and real life situations.

At your Junior Secondary level you learnt that matter is made up of small particles called atoms. These atoms cannot exist singly because they are unstable. As a result, they combine or bond with other atoms in a chemical reaction to attain a stable configuration of 2 or 8 electrons in the resulting outmost shell. This may be attained when an atom shares its valence electrons with another atom in what is called covalent bonding or when an atom completely transfers its electrons to another as in ionic bonding. You may also recall that there is a special kind of bonding between atoms of a metal lying together in a lattice called metallic bonding.

Note: *In metallic bonding, metal atoms release their loosely held valence electrons and these get shared by attraction of the negatively charged electrons to the positively charged metal atom.*

Figure 3.1, 3.2 and *3.3* show three types of bonding.

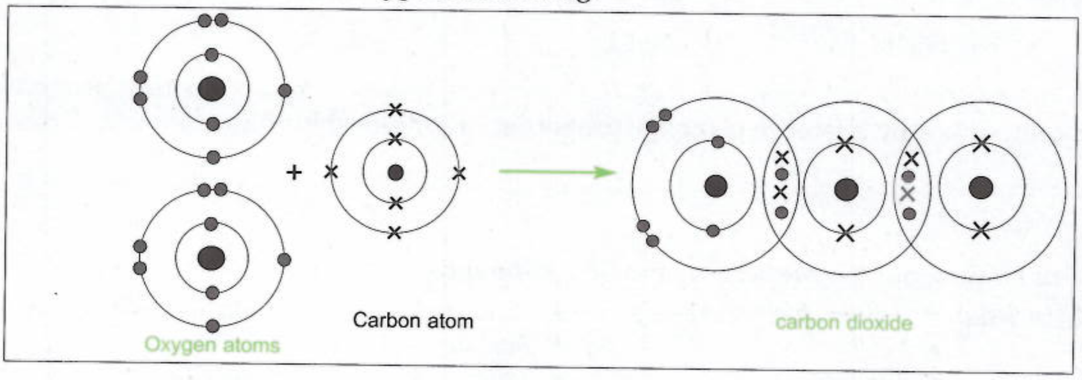


Figure 3.1: covalent bonding (Carbon and oxygen share electrons)

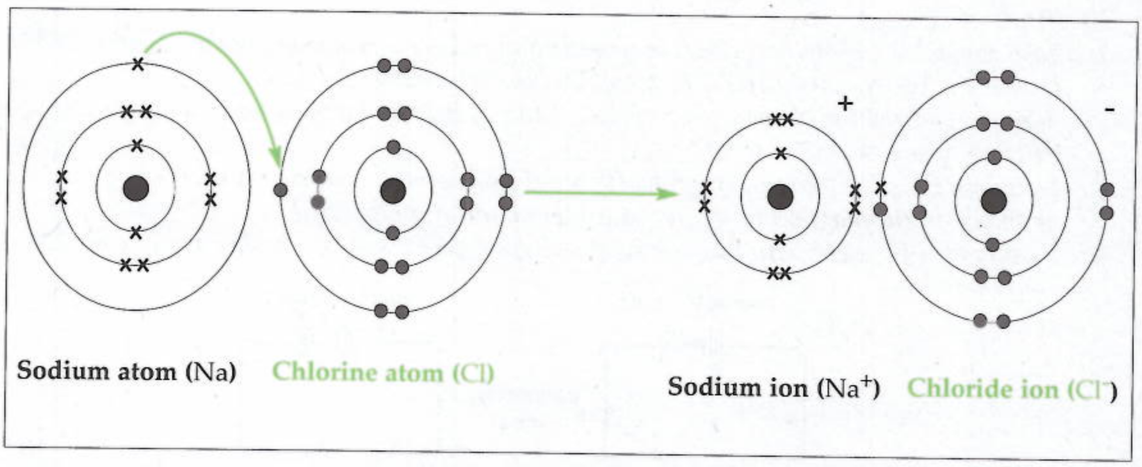


Figure 3.2: ionic bonding (A positively carged atom is attracted to a negatively charged one)

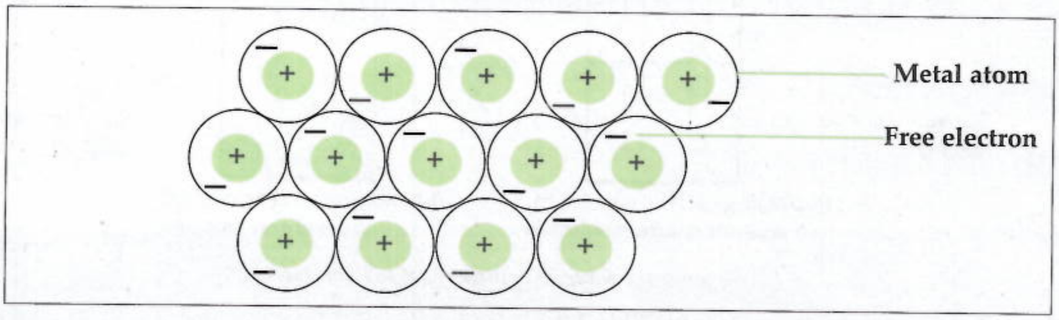


Figure 3.3: metallic bonding (Free electrons are attracted to positively charged metal atoms)

In this unit you will focus on properties of ionic, covalent and metallic bonds. You will also see how such bonding characteristics influence the properties and uses of ionic compounds, covalent compounds and metals in our daily lives.

Properties of ionic compounds

You will recall that ionic compounds are formed by ionic bonding. During such bonding, a metal atom loses its electron(s) while a non-metal atom gains the electron(s) and the two ions get attached to one another by an electrostatic force caused by attraction of opposite charges. The attraction is so great that the resulting bond is a strong bond which cannot easily be broken. For example sodium metal ion (Na^+) bonds with a chloride ion (Cl^-) to form an ionic compound called sodium chloride (NaCl).

$$Na^+ + Cl^- \longrightarrow NaCl$$

What are the properties of compounds resulting from ionic bonding? You will study properties of ionic compounds by inference using the properties of sodium chloride outlined in *Activity* 3.1.

Activity 3.1:

Aim: Finding out the properties of ionic compounds

Materials:
- *sodium chloride crystals*
- *2 cells*
- *Burner*
- *Connecting wire*
- *2 electrode*
- *spatula*
- *bulb*
- *small beaker*
- *water*
- *stirring rod*

Method:
1. *Take about 1.0 g of sodium chloride crystals and put them in a small beaker. Heat it over a burner to approximately 100°C or more. Did the salt change its state?*
2. *Take 1.0 g of sodium chloride on a spatula and put it in 25 ml of water in a beaker and stir. What do you observe?*
3. *Now take the solution resulting from (ii) above and connect it using connecting wires to a cell as shown in the diagram in figure 3.4.(a). Does the bulb light up?*
4. *Finally pass the electricity through solid sodium chloride salt in a beaker as in figure 3.4. (b).*

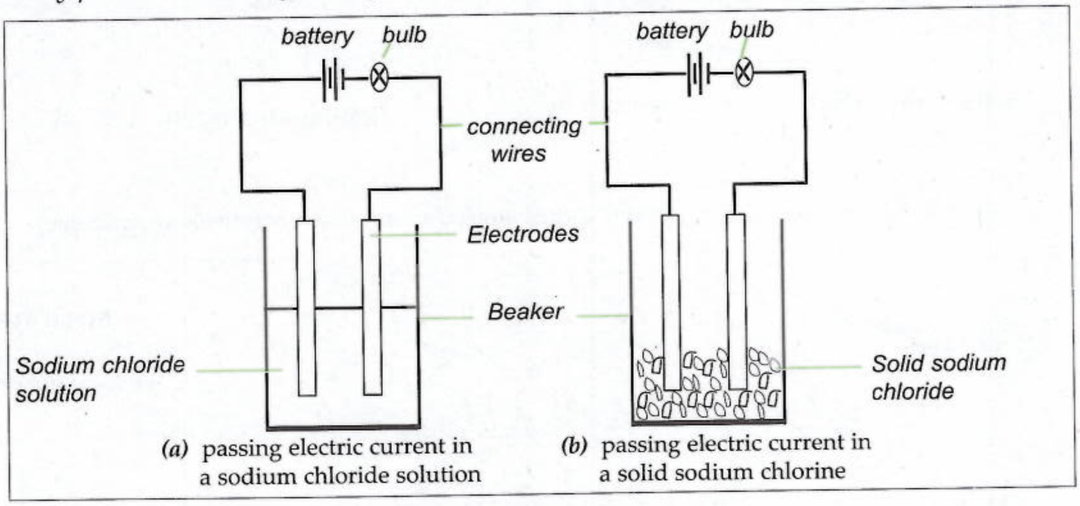


Figure 3.4: conductivity property of sodium chloride

5. *Record your observations in Table 3.1.*

Serial #	Activity	Observation
1	State of salt at room temperature.	
2	Heating the salt to 100 degrees celsius.	
3	Adding salt to water	
4	Passing a current through the solution of sodium chloride	
5	Passing a current through the solid sodium chloride	

Conclusion: *What conclusions concerning ionic compounds do you draw from this activity?*

Discussion

You may have observed that ionic compounds are solids at room temperature and do not easily melt. Often particles in a compound are strongly held together by electrostatic forces. As a result, ionic compounds have high melting and boiling points. You may have further observed that, ionic compounds easily dissolve in water. Water is one of the polar solvents: that is, solvents that are charged and can easily dissolve ionic compounds because the charges from the solvent attract opposite charges from the dissolving compound, hence making the compound to spread out.

You may also have observed that solutions of ionic compounds conduct electricity. This is true as ionic compounds are made up of ions, which are set free to move on dissolution of the compound. On the contrary, ionic compounds will not conduct electricity in their solid state.

Why do solid ionic compounds fail to conduct electricity?

Properties of covalent compounds

Note: In covalent bonding atoms share their valence electrons in order to form a molecule that has a stable configuration. During this process there is an overlap of shells as shown in *Figure 3.5*.

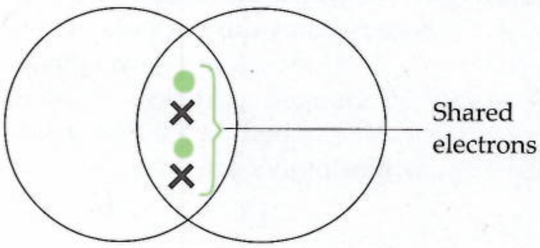


Figure 3.5: Atoms overlap forming a covalent bonding

What keeps atoms together? Why are covalent bonds strong?

The atoms stay together due to a strong binding force resulting from the covalent bonding. This force is known as *intramolecular force*. In a substance covalently bonded, molecules hold onto each other by a weak adhesive force called *intermolecular forces (IMF)*. IMF is also known as Van de Waal's forces.

The following activity will help you learn the characteristics of covalent compounds.

What do you think are the characteristics of compounds resulting from covalent bonding?

Activity 3.2:

Aim: *Finding out the properties of covalent compounds.*

Materials:

- *sugar*
- *test tube holder*
- *cells*
- *water*
- *gas burner*
- *thermometer*
- *connecting wires*
- *2 beakers*
- *test tube*
- *electrodes*
- *a bulb*

Method:

1. *Carry out your activity as in in activity 3.1 but now using sugar*
2. *Record your observations in the table below*

Serial #	Activity	Observation
1	State of sugar at room temperature	
2	Heating sugar to 100°C.	
3	Adding the sugar to water	
4	Passing a current through the solution of sugar	
5	Passing a current through the solid sugar	

Conclusion:

1. Do covalent compounds require large amounts of heat to melt and boil?
2. How do they compare with ionic compounds?

Discussion

From *Activity* 3.2 you have observed that *covalent compounds do melt easily when heated*. You have also observed that some covalent compounds, like sugar, can dissolve in water. However, *most covalent compounds do not dissolve in water* because they are not polar. Covalent compounds will *easily dissolve in non-polar solvents* such as methane and carbon tetrachloride. Due to lack of charged particles, solutions of covalent compounds *do not conduct electricity* both in solution and solid state. Some compounds are *gases* whiles others are *solids* or *liquids* at room temperature

Why do covalent compounds melt easily when heated?

Exercise 1

1. Explain the difference in bonding between a covalent and an ionic compound.
2. Can an ionic compound conduct electricity when in molten state? Explain your answer.
3. Why do covalent compounds not conduct electricity in solution and solid states?
4. Covalent compounds have low melting and boiling points. Explain the cause of this characteristic.
5. What is the difference between intra-molecular and intermolecular forces?

Structural differences of ionic and covalent compounds

1. Ionic compounds

Ionic compounds form giant structures. They are solid at room temperature due to their high melting and boiling points. The ions are packed in a regular arrangement or order forming a crystalline structure as shown in *Figure 3.12*. There are specific points which ions fill, the centre of which is called a ***lattice point***. In a single crystal, you will find many thousands of ions packed together with strong forces of attraction between positive and negative charges.

Study the structures of sodium chloride and calcium chloride in *Figure 3.12*. Both compounds are ionic.

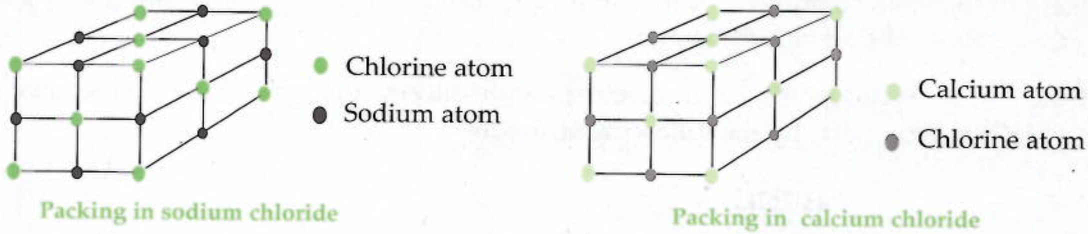


Figure 3.12: Ionic structures

2. Covalent compounds

Covalent compounds may form **simple** as well as **giant molecular structures** such as in *diamond, silicon oxide* and *graphite*. In simple molecular structures, few atoms are joined up. They are bound together by weak intermolecular forces. *As the number of molecules forming a compound become very large, giant structures are formed,* ***consequently*** *the forces holding atoms together also increase*. Atoms in giant molecular compounds are firmly held by strong ***covalent bonds***. The giant compounds are therefore strong and have high melting and boiling points.

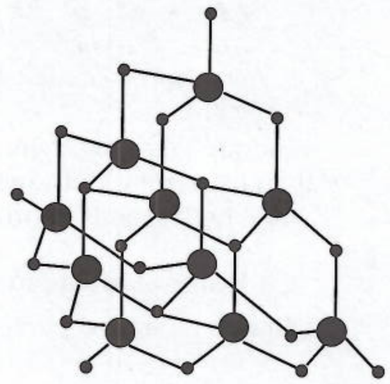


Figure 3.13: A giant covalent structure

	Ionic	Covalent
1	Form giant structures called lattice	May form simple as well as giant molecular structures
2	Are bound together by strong ionic forces	Are bound together by weak intermolecular forces
3	Have high melting and boiling points	Have relatively low melting and boiling points
4	Ions are packed in a regular arrangement or order, thus forming a crystalline structure	Molecules may not be packed in a regular arrangement

Table 3.3: Structural differences between ionic and covalent compounds

Exercise 3

1. Draw the structure of the bonding between hydrogen and oxygen to form a water molecule. Show the polarity of the bonds in the molecule using an arrow.
2. What would you suggest about the solubility in water and electrical conductivity of a solution of HCl? Why?
3. Mention two differences between covalent and ionic compounds?

Polar covalent compounds

Having studied properties of covalent compounds in the previous section, especially for *pure covalent compounds*; in which supply and sharing of electrons is equally done for the mutual benefit of the atoms involved. ***What happens if the sharing or supply of the bond electrons is unequal?*** Study the situations below:

Situation 1: Hydrogen shares its electrons with chlorine to form the covalent compound HCl as illustrated by the following structure:

Figure 3.6: the structure of hydrogen chloride showing shared electrons

a. What do you notice about the position of the shared electron pair?
b. What would be the reason for such behavior on chlorine?
c. Would the negative and positive charges balance on the side of chlorine? How about on the side of hydrogen? Suggest the result of the balance or in-balance.

Chlorine takes the shared electrons closely to its side because it is more *electronegative* than hydrogen. Chlorine becomes surrounded by more negative than positive charges, while hydrogen is deprived of electrons and has more positive charges than negatives.

The result of situation 1 is the development of *partial negative charge* (δ^-) on the chlorine atom and *partial positive charge* (δ^+) on the hydrogen atom, respectively.

$$\overset{\text{Drift of electrons}}{\longrightarrow}$$
$$\delta^{+}\mathrm{H}-\mathrm{Cl}\,\delta^{-}$$

Figure 3.7: A polar bond in hydrogen chloride

There is a *permanent dipole* formed by the drift of the electron cloud in the direction of chlorine as shown by the arrow. The resulting bond is said to be *polar* (*or charged*) and the compound is a *polar covalent compound*. In this situation *the bond that was supposed to be purely covalent has ionic characteristics. The greater the difference in electronegativities of the atoms sharing their electrons, in a covalent bond, the greater the ionic character of the bond*. Polar compounds easily dissolve in water to form an electrolyte. Some polar compounds are acidic.

Situation 2: Sometimes both electrons invovled in the formation of a bond are donated by one atom. Such bonds are called *co-ordinate* or *dative bond*. A coordinate bond is one of the covalent bonds and is formed between a positive ion, such as a hydrogen ion, H^+, and a compound having a lone pair of electrons such as ammonia (NH_3) or water

(H_2O). **Recall** that, a lone pair of electrons consist of non-bonding electrons, which do not usually take part in a chemical reaction. Study the following examples:

a. Co-ordinate bond between ammonia and hydrogen ion

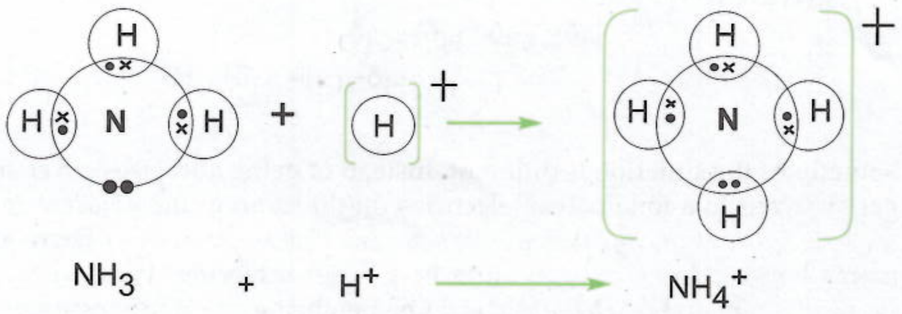


Figure 3.8: Formatiom of an ammonium ion from lone pair of electrons in ammonia and positive charge in a H-ion

The positive ion adds itself to the lone pair of electrons by overlapping with it. Since ammonia is neutral, the positive charge of hydrogen ion is transfered to the resulting ammonium ion, NH_4^+.

Note:

a) *The shared pair of electrons is provided by one of the atoms in dative covalent bonding.*

b) *The atom that provides both electrons of the covalent bond is called the* ***donor.*** *In this case nitrogen is the donor*

c) *The atom that accepts both electrons is called* ***acceptor.*** *In the illustration above, hydrogen ion, H^+ is the acceptor*

b. Co-ordinate bond between water and hydrogen ion (H_3O^+)

The co-ordinate bond in a ***hydronium ion*** is formed when a hydrogen ion adds to the water molecule by overlapping with the lone pair of electrons on water. *Figure 3.9* illustrates how the hydronium ion is formed.

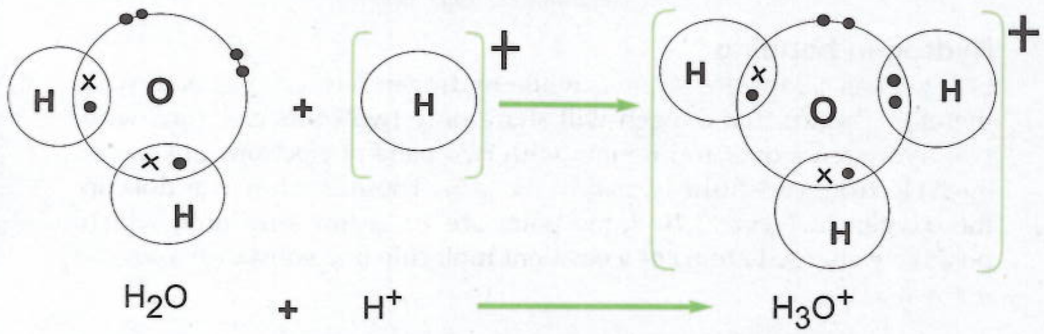


Figure 3.9: Formatiom of a hydronium ion from a lone pair of electrons on a water molecule and a hydrogen ion

What happens in a polar covalent substance?

We have alewady seen the hydrogen chloride molecule (H-Cl), that the bond **H-Cl** behaves like a dipole ($H^{\delta+}$-$Cl^{\delta-}$). In the HCl substance there exists a number of such dipole bonds creating *permanent dipole-dipole* interactions between adjacent molecules. The *positive charge* in one molecule *attracts a negative charge* from another molecule as shown below:

$H^{\delta+}-Cl^{\delta-}\ \ H^{\delta+}-Cl^{\delta-}$

$H^{\delta+}-Cl^{\delta-}\ \ H^{\delta+}-Cl^{\delta-}$

$H^{\delta+}-Cl^{\delta-}\ \ H^{\delta+}-Cl^{\delta-}$

Sometimes the situation is different. Instead of being attracted to a charge, an atom may get attracted to a lone pair of electrons on the atom in the adjacent molecule. (*You will see how this happens by studying the behavior of water molecules.*) There are two hydrogen atoms bonded to an oxygen atom in a water molecule. ***Would you expect the water molecule to be polar?*** *Definitely, yes*! Oxygen being so electronegative ***attracts the bond electrons*** to itself more than they get attracted to the side of hydrogen. The result is as shown in *Figure 3.10*:

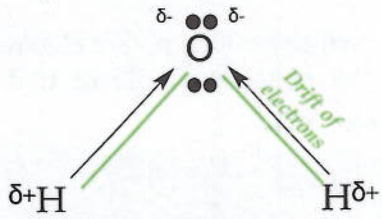


Figure 3.10: water molecule

Water molecules would then bond to each other through dipole-dipole attraction. Study *Figure 3.11*.

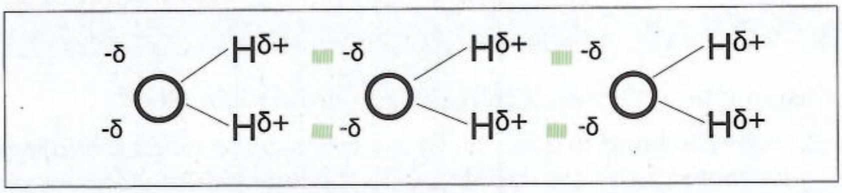


Figure 3.11: A positively charged hydrogen attracked to a negatively charged oxygen in a different water molecule

Hydrogen bonding

Oxygen has *six valence electrons* while hydrogen has *only one*. You will, therefore, notice that oxygen will share only two of its electrons with two hydrogen atoms and remain with two pairs of electrons unshared. Such electrons constitute ***lone electron pairs***. These are shown as dots on the oxygen in *Figure 3.10*. Lone pairs are ***potential sites*** onto which positively charged atoms of a covalent molecule in a substance ***may get attracted***.

How do two atoms share their electrons?

This type of bonding, involves partially positive hydrogen atoms and other partially negative particles, is called ***Hydrogen bonding***.

In hygrogen bonding there is a special interaction between the lone pairs of oxygen and the partially positively charged hydrogen atoms of the other molecule as shown in *Figure 3.11*. Similar attractions may take place when hydrogen is bonded to more electronegative atoms such as *fluorine* (F) and *nitrogen* (N) which also have lone pairs of electrons.

The difference between polar and non-polar compounds can now be summarised as in the *Table 3.3*.

Polar compounds	Non-polar compounds
Soluble in water	Insoluble in water
High surface tension	Low surface tension
High melting and boiling points	Low melting and boiling points

Table 3.3

Exercise 2

1. Which atom would attract the other with its lone pair of electrons in HCl and H_2O?
2. Why do polar compounds dissolve in water?

Allotropy

What is an allotropy?

You have seen that atoms of different elements combine with other atoms so that they attain stability. The atoms may be of the same or different kinds. When atoms of the same kind join up they form *elements*. The manner in which these atoms re-arrange themselves result in different structures with different chemical and physical properties such as *color*, *state* and *density*. Such forms are called ***allotropes***. Therefore, ***allotropy** is the existence of an element in two or more forms at the same physical state*. The physical properties may differ widely, but identical chemical compounds can be formed from the various allotropes of the same element.

Acitvity 3.5:

Aim: Discovering substances that exist in allotropy

1. *Find out which elements exist as allotropes and list them in a table.*
2. *State the names and characteristics of the allotropes of each element.*

You must have found out that many elements naturally exist in allotropy. Notable examples include *carbon*, *oxygen*, *sulphur* and *phosphorus*. We will look at some of these allotropes and study them in more detail.

1. Allotropes of Oxygen

Only two oxygen allotropes are known and these are *oxygen* (O_2) and *ozone* (O_3) as shown in *Figure 3.8*. At room temperature and pressure both oxygen and ozone are gases. Ozone is usually made from oxygen by reaction with an oxygen atom. It:

a. has a characteristic *sharp smell* or *ordour*.

b. is very unstable and *changes spontaneously* to oxygen.

c. is quite useful in the atmosphere as it *helps to absorb ultraviolet* (UV) *light from the sun*, thereby protecting the earth from the sun's UV radiation.

d. It is also a *very powerful oxidizing agent*.

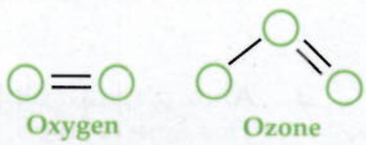


Figure 3.15: Structures of Oxygen and Ozone

2. Allotropes of sulphur

Sulphur exists in four different allotropes. The following are the allotropes of sulphur:

a. Rhombic sulphur
b. Monoclinic sulphur
c. Plastic sulphur
d. Amorphous sulphur

a. Rhombic sulphur

This is also called *alpha sulphur* (α sulphur). It consists of large yellow crystals with an *octahedral shape*. Rhombic sulphur *crystallizes from a solution of carbon disulphide* at room temperature. The crystals of this allotrope *are stable below* **96°C**. Above this temperature, they s*lowly change to the monoclinic form*.

The S_8 molecules of rhombic sulphur interlock with each other.

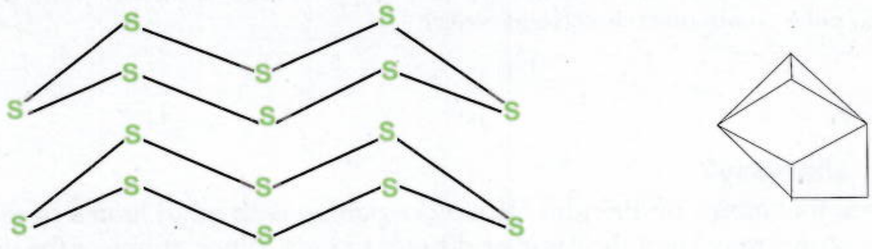


Fig. 3.16 Interlocking pattern of arrangement of S_8 molecules in rhombic sulphur

b. Monolinic sulphur

Monoclinic sulphur is also called **beta sulphur** (***β-sulphur***). It consists of *needle-shaped transparent crystals* which are stable above 96°C. Below this temperature they gradually change to rhombic form. The S_8 molecules of monoclinic sulphur are stacked.

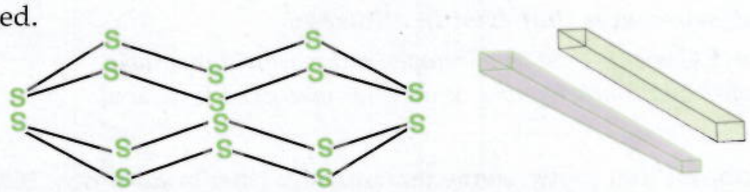


Fig. 3.17: Slacked pattern of arrangement of S_8 molecule in monoclinic sulphur

c. Plastic sulphur

Plastic sulphur *behaves like rubber*. It is formed *when boiling sulphur is poured into cold water*. Then, the sulphur liquid gradually takes the form of rhombic sulphur.

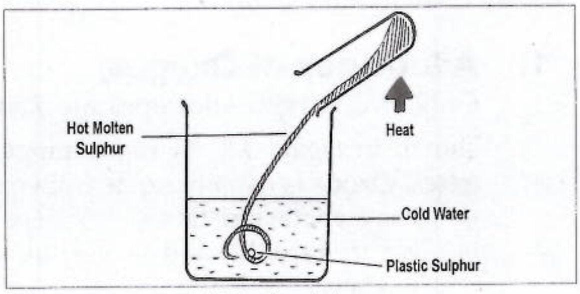


Fig. 3.18 Plastic sulphur

d. Amorphous sulphur

This form of sulphur is formed from a solution of *hydrogen sulphide* (H_2S) in water. When a saturated solution of this gas in water is exposed to air, a white powder of amorphous sulphur is produced.

3. Allotropes of carbon

Carbon is a non-metal, which exists in two main allotropes: graphite and diamond.

a. Diamond

This is a giant structure of carbon atoms in which each atom is covalently bonded to other four atoms. This bonding style results into a giant structure of clear, colorless crystals of tetrahedrally arranged atoms as shown in *Figure 3.19*.

What characteristics would such a tetrahedrally arranged structure have?

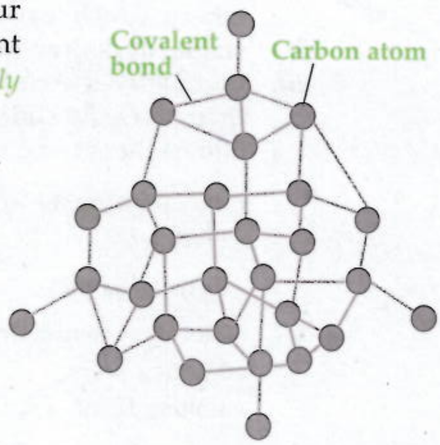


Figure 3.19: Structure of diamond

The bonding arrangement of tetrahedral carbon atoms enables the diamond to be extremely hard with atoms very closely packed. Since all electrons are involved in the formation of covalent bonds, no delocalized or mobile electrons are found. Therefore, the diamond does not conduct electricity.

The diamond:

i. is the hardest known, and the least compressible substance as the atoms are closely packed.

ii. has very high melting and boiling points.

iii. is a good absoorber of heat at room temperature, such that it is used as a heat sink to rapidly cool electronic components.

iv. is useful for making cutting tool in industry and surgery due to its hardness.

b. Graphite

Graphite is made up plane layers of carbon atoms. Each atom is bonded to three others using covalent bonds. This results into a structure with fused hexagonal rings arranged in an infinite three-dimensional network. The layers are attracted to each other by weak cohesive forces (the **van der Waals forces**). *Figure 3.20* illustrates this structure.

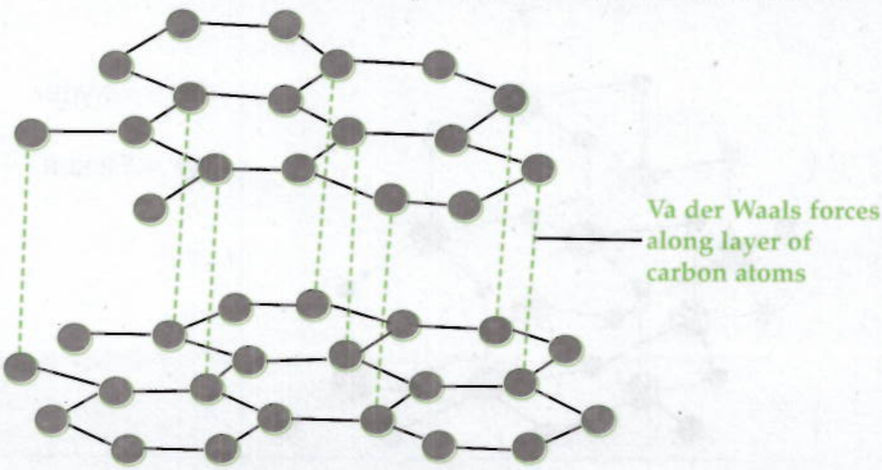


Figure 3.20: The structure of graphite

Basing on what you have seen so far, of what use would graphite be?

Some of the characteristics of graphites are:

i. ***can conduct electricity:*** since the carbon atoms are bonded to only three others, which leaves one electron free to move about.

ii. stable, dark grey in color and waxy.

iii. ***lubricants:*** the weak, non-bonding, interaction between the layers, allows them to easily slide over each other; hence soft and slippery. This property allows them to be used as lubricants.

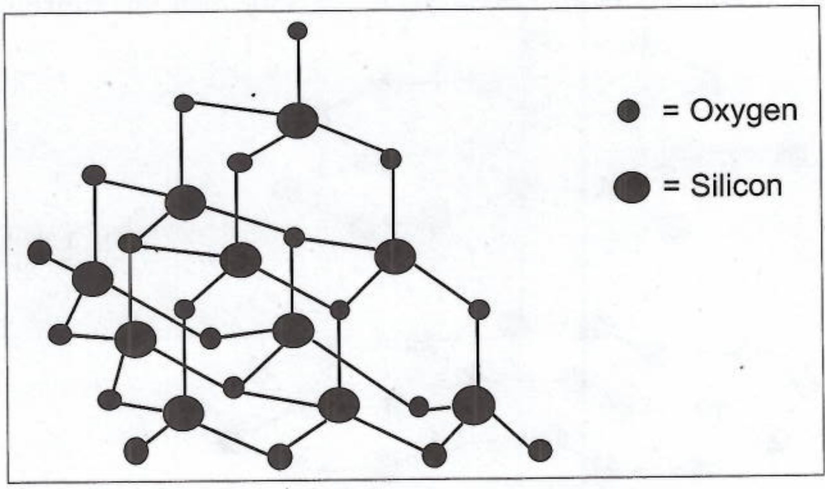
Property	Diamond	Graphite
Hardness	Very hard	Soft with slippery feel
Density (g/cm^3)	Higher density	Lower density
Electrical conductivity	Does not conduct	Can conduct
Melting point	Higher	Lower
Boiling Point	Higher	Lower
Use	Making jewellery, cutters	Making lubricants, pencils and electrodes

Table 3.2: Differences in physical properties between diamond and graphite

Comparison between Diamond and Silicon oxide

Silicon is an element in *Group* 4 of the Periodic Table. It therefore has a valence of four and the atoms may bond tetrahedrally to other *Si atoms*, just as in carbon diamond. In its elemental state silicon is a *non-conductor of electricity*. But when its temperature is increased, it becomes an *excellent semi-conductor* (as you will see later).

Silicon reacts with oxygen to form silicon oxide. In normal situations, we expect silicon and oxygen to form a compound similar to carbon dioxide (CO_2), which is silicon dioxide (SiO_2). Then, ***what is the formula and structure of silicon dioxide?*** Silicon dioxide is very unstable as it easily forms ***silicon oxide*** (SiO). In silicon oxide, *silicon gets bonded tetrahedrally to four oxygen atoms*, resulting into a giant structure. An example of silicon oxide is *common sand* or *silica*.



What properties would this giant structure have?

Figure 3.21: Structure of silicon oxide

Let's investigate some of these properties in *Activity 3.6*:

Activity 3.6

Aim: *Finding out the physical properties of Silicon oxide*

Materials:
- *River sand*
- *Beaker*
- *spatula*
- *pestle*
- *gas burner*
- *Crucible*
- *100ml of water*

Procedure:

1. *Take some river sand. Put approximately 10g on a spatula and heat over a burner flame.*
2. *Dissolve a spatula full of sand in 100 ml of water in a beaker. What happens?*
3. *Crush some sand in a crucible using a pestle. How easily does this happen?*

Results/observations:

What happens when you heat river sand? Does it melt? Does the sand really dissolve in water? How easly would you crush the sand?

You will have noticed from *Activity 3.6* that sand (*silicon oxide*) is:

1. ***quite hard***; *it cannot easily be crushed*
2. *cannot dissolve in water* and
3. has *high melting point* (1160°C).

How do you compare the characteristics of silicon with those of carbon diamond?

Silicon oxide forms a giant molecular structure, which is rigid and hard. The structure brings about strong covalent bonding between atoms, which cannot easily be broken.

Table 3.3 summarises some similarities between silicon oxide and diamond:

Property	Diamond	Silicon oxide
Bonding arrangement	Tetrahedral	Tetrahedral
Packing of atoms	Very close	Very close
Hardness	Very hard	Very hard
Density (g/cm^3)	3.51	2.25
Electrical conductivity	Does not conduct as there are no free electrons	Does not conduct as there are no free electrons
Melting point	High	High
Boiling Point	High	High

Table 3.3 similarities in physical properties of silicon oxide and diamond

Exercise 4

1. Mention any three properties of metals and the activities to which these are put to use.
2. Explain why metals are good conductors of heat and electricity.
3. Define the term allotropy
4. Which three elements exist as allotropes? Name their allotropes.
5. What are the similarities between silicon oxide and diamond? Mention any three.

Properties of metals

A metal comprises of atoms packed together in a crystal lattice. At your Junior Secondary level, you learnt that, the atoms have their electrons delocalized all over the entire lattice. There are great cohesive forces between the central metal atoms and the delocalized electrons, which keeps all the atoms together. *Figure 3.14* depicts this metallic lattice.

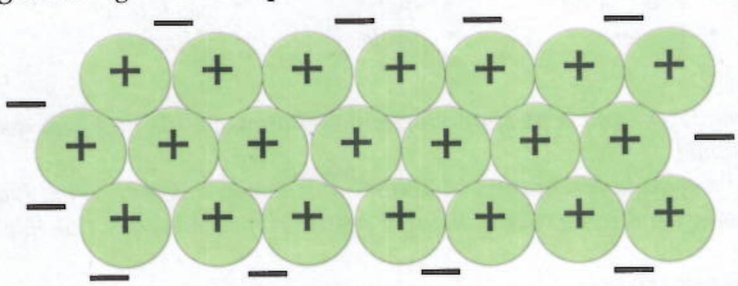


Figure 3.14: A close packing metal lattice

Think of the characteristics of metals in terms of their solubility in water, density, melting and boiling points and heat or electrical conductivity. You can verify your suggestions by carrying out the *Activity 3.3*.

What physical properties do metals have?

Activity 3.3

Aim: *Finding out the properties of metals*

Materials:
- *Pieces of copper metal*
- *Connecting wires*
- *water*
- *2 cells*
- *ammeter*

Method:
1. *Carry out your activity as in* Activity 3.1 *but now use a sheet of copper metal.*
2. *Record your observations in Table 3.7.*

Serial #	Activity	Observation
1	*Heating metal as much as possible.*	
2	*Putting the metal in water*	
3	*Passing a current through copper metal*	

Table 3. 8

Conclusion:

You must have noticed that metals do not readily melt, because their melting and boiling points are very high. When you connect a piece of copper metal to an electric circuit current flows. This shows that most metals are good conductors of electricity. Metals are also good conductors of heat. Their delocalised electrons enable metals to conduct heat and electricity. Metals will not dissolve in water but they react with water to form a metal hydroxide. Metallic particles are closely packed hence metals have very high densities. Metals are solids at room temperature, except mercury. Some are ductile, sonorous and malleable.

What properties of metal do you see from this activity?

Table 3.4 summarises characteristics of ionic compounds, covalent compounds and metals.

Characteristic	Ionic compounds	Covalent Compounds	Metals
Formation	Formed by *electron transfer* from a metal atom (X) to non-metal atom (Y), producing *oppositely charged ions*, X^+ and Y^- which *attract to form a compound* of nature X^+Y^-	Formed *when electrons are shared* between two or more atoms. 	The *positive ions occupy fixed positions in a lattice* and the *delocalized electrons move freely* throughout the lattice (*refer to Figure 3.14*)
Melting and Boiling points	*High melting* and *boiling points* due to availability of *strong electrostatic forces* between ions in the solid lattice	Low *melting* and *boiling points* due to presence of *weak forces between molecules* in simple molecular structures, **except** in giant molecular structures	*High melting and boiling points* due to strong *metallic bonds* between positive ions and negative electrons in the lattice
Conductivity	*Non-conductors of electricity in solid state*, but when melted or dissolved in water because *then the ionic lattice breaks down and ions are free to move* as mobile charge carriers	*Non-conductors of electricity* as there are no free or mobile charged particles	*Good thermal and electrical conductors* due to mobile, delocalized electrons that conduct heat and electricity, *even in the solid state*
Solubility	Usually *dissolve in polar solvents* (e.g. *water*) since *solvent molecules attract ions* in the lattice and surround each one of them (*hydration*)	Molecular structures often *soluble in non-polar solvents* (e.g. *hexane*)	*Insoluble* in polar and non-polar solvents

Table 3.4 Comparisons of ionic and covalent compounds and metals

Uses of metals

In the previous section, you saw that metals have closely packed particles with high cohesive forces between them. These properties make metals:

1. *strong with high density* and *melting points.*
2. hard but *ductile* and *malleable*: i.e. you can bend or hammer it into various shapes.
3. delocalized electrons enable them to be *good conductors of heat and electricity.*

What uses of metals do you know from your daily activities at home and school?

These properties make metals very useful.

Acitivity 3.4

Aim: *Finding out uses of metals*

Procedure:

1. Table 3.6 *outlines some characteristics of metals*
2. *Complete column 2 and 3 by filling the possible use and example of metal used for the purpose.*

	Metal Characteristic	Possible use	Example of metal
1	High melting points	for making cooking pots	Aluminium
2	Conduct electricity		Copper
3	Ductile	For electrical wiring	
4	Malleable		Aluminium
5	Sonorous	For making bell	
6	High rusting resistance		Zinc
7	Unreactive and non-toxic	Protecting steel from rusting	
8			

Table 3.6: Properties and uses of metals

3. *Report to the whole class how you filled up your table.*

Alloys

In the last section, you saw that metals are used for various activities due to their great strength, ductility, malleability and their ability to conduct heat and electricity. They are also used to *reinforce structures* such as bridges, storey buildings and to make cooking utensils, cables for electricity and water pipes.

However, metals can be amplified so that they are even better. One method is to make metals into alloys. An *alloy is a mixture of a metal with another metal*. For instance, a metal can be mixed with carbon. The mixing is done in deffferent proportions while the metal is in molten state so as to obtain desirable qualities. Some best examples of alloys are **steel**, **bronze** and **brass**. *Table* 3.7 outlines some alloys, their properties and uses.

How can you improve the quality of the performance of metals and more useful?

Mixture	Alloy	Special property	Use
70% Iron, 20% Chronium, 10% Nickel	Stainless steel	Does not rust	Making car bodies, cutting tools, chisels, razor blades
95% Copper, 5% Tin	Bronze	harder than brass, does not corrode, sonorous	Making statues, ornaments, musical instruments, church bells
30% Lead, 70% Tin + some antimony	Pewter	Malleable	for making plates, mugs and ornaments
30% Lead, 70% Tin	Solder	low melting point	Joining wires and pipes (*welding*)
70% Aluminium, 30% Magnesium	Magnalium	Light and tough	Construction of aircraft
70% Copper, 25% Nickel	Cupro-nickel	Hard wearing, attractive colour	Making coins
70% Copper, 30% Zinc	Brass	harder than copper, does not corrode	Making machine bearing, mineral instrument, door furniture, jewellery and electrical connections.

Table 3.7: Alloy mixtures and their uses

Unit Summary

- Ionic compounds are *made up of charged ions*, which are *bonded* together by *strong electrostatic forces*. The ions are in close packing and arranged in a regular pattern in a crystal lattice.
- Covalent compounds consist of atoms joined together by *weak intermolecular forces*. The atoms do not form a regular pattern.
- In some covalent bonds the electrons are *unequally shared because one atom is more electronegative than the other*. In such a case a *polar bond develops as the more electronegative atom attracts electrons to itself* and become negatively charged while the other atom becomes positively charged.
- *Allotropy is a state in which a substance made of atoms of the same element exist in different forms* due to differences in which the atoms join up. Many substances exist as allotropes, for example, *carbon diamond and graphite*, *rhombic and monoclinic sulphur*, and *oxygen and ozone*. Allotropes allow different substances to be put to different use.
- Metal atoms *bond up in a pool of delocalized electrons* forming a metallic crystal. As a result, metals are strong, hard and can conduct heat and electricity.
- The properties of metals *may be improved when you mix one metal with another in a molten state to form an alloy*. Alloys have better properties and uses than the original metals.

Unit Exercise

1. You have ionic and molecular compounds X and Y. The following are some of their properties:

	X	Y
Melting point	782	-77
Boiling point	1600	-34
Solubility in water	Soluble	____
Electrical conductivity	____	____

   a) Which substance would be calcium chloride, why?
   b) Complete the table by filling in information for solubility and electrical conductivity.

2. Discuss the differences in the properties of sodium metal, chlorine gas and sodium chloride in terms of:
   a) state at room temperature
   b) solubility
   c) electrical conductivity

3. Explain the following:
   a) Ionic compounds only conducts electricity only when in solution or molten
   b) Ionic compounds are solid, whereas molecular compounds can exist in all three states of matter at room temperature.

4. Explain why:
   a) graphite can conduct electricity, while diamond does not
   b) graphite is good lubricant.

5. Describe the structural difference between graphite and diamond.

6. describe three differences between graphite and diamond in terms of their:
   a) hardness
   b) density
   c) uses

7. What is the difference between these words: *malleable* and *ductility* as regards metals?

8. What is an alloy?

9. Explain why tin is used to coat food cans.

10. Name an alloy that:
    a) has a low meelting point
    b) never rust
    c) is used for making musical instruments

11. Expalin the use of each of the alloys you have mentioned in question 7 (a) and (b).

12. Explain one difference in composition betweeen brass and bronze.

Unit 4

Stoichiometry

Success Criteria

By the end of this unit you should be able to:

1. Write a balanced equation
2. Work out the relative formula mass of a compound.
3. Define the mole of a substance
4. Convert moles into other units of measurements
5. Determine percentage of water in molecular and hydrated ionic compounds
6. Deduce empirical and molecular formulae from relevant data
7. Calculate concentration of solutions
8. Prepare standard solution
9. Determine the concentration of a solution using titration
10. Determine the yield in a chemical reaction

Key words:

In this unit you will find these key terms and concepts:

chemical formula, Relative Formula Mass, mole, avogadro constant, standard solution, titration

Ensure that you understand and learn how to apply them both for your academic and real life situations.

What is Stoichiometry?

At your Junior Secondary level, you learnt how to write chemical formulae of both elements and compounds. You learnt that elements and compounds that react are called *reactants* and those produced are *products*, and also that a *balanced chemical equation shows the right proportions* of these reactants and products.

Stoichiometry is the study of the amounts of reactants and products of a chemical reaction. It is a branch of chemistry that is *based on the Law of conservation of mass*.

Chemical Equation

A *chemical equation* is the short and accurate *representation of a chemical reaction*. It has **reactants** on the left and **products** on the right *separated by an arrow*.

Writing balanced equation

A good chemical equation must be balanced. **Remember:**

1. a balanced chemical equation will show the *same number of atoms* of each kind that make up *reactants on one side* and *products on the other*.
2. the *kinds of atoms* involved in a reaction *do not increase* or *decrease after the reaction* but remain the same.

The following *steps* should help to remind you on how to *balance a chemical equation*:

1. First write the *word equation* with reactants on the left and products on the right
2. Write a *correct symbol* or *formula* for each reactant and product.
3. *Balance the numbers of atoms* on both sides by *putting a coefficient against the molecules*. **Do not change the subscripts** of the atoms in the formula.
4. *Write the physical states* after each reactant and product in brackets. Use (*s*) for solid, (*l*) for liquid, (*g*) for gas and (*aq*) for aqueous solution.

Examples

1. Carbon burns in air to produce carbon dioxide gas. Write and balance the chemical equation for the reaction.

Word equation: Carbon + Oxygen $\longrightarrow$ Carbon dioxide

Chemical equation: $C + O_2 \longrightarrow CO_2$

Check the number of carbon and oxygen atoms on both sides. Are they the same?

This equation is balanced since the numbers of atoms of each kind of atom on both sides are the same.

Hence the final equation will be: $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$

2. Sodium reacts with chlorine gas to form sodium chloride. Write and balance the chemical equation for the reaction.

Word equation: Sodium + chlorine $\longrightarrow$ Sodium chloride

Chemical equation: $Na + Cl_2 \longrightarrow NaCl$

Now looking at the equation, we have 2(Cl atoms) on the left and 1(Cl atom) on the right. *So chlorine does not balance*. How would you then make the atoms balanced?

Suggestion: multiply **NaCl** by **2**

This becomes $Na + Cl_2 \longrightarrow 2NaCl$

Chlorine is balanced but sodium is not.

Suggestion: multiply **Na** by **2**

The equation now becomes $2Na + Cl_2 \longrightarrow 2NaCl$

It is balanced. Now put in the physical states of both reactants and product.

The final equation will now be: $2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$

Exercise 1

1. Write and balance the chemical equation for each of the reactions below:
   a) Magnesium reacts with oxygen to produce magnesium oxide
   b) Potassium reacts with chlorine to produce potassium chloride
   c) Aluminium reacts with fluorine to produce aluminium fluoride.

2. Balance the following chemical equations:
   a) $Li_{(s)} + Cl_{2(g)} \longrightarrow LiCl_{(s)}$
   b) $Mg_{(s)} + F_{2(g)} \longrightarrow MgF_{2(s)}$
   c) $Na_{(s)}\ H_2O_{(l)} \longrightarrow NaOH_{(aq)} + H_2(g)$
   d) $Al_{(s)} + O_{2(g)} \longrightarrow Al_2O_{3(s)}$

Relative formula mass for compounds

Recap of your Junior Secondary course work, you saw that *relative formular mass* (RFM) is *the mass of a molecule relative to the mass of carbon -12*. Hence to work out the RFM of a compound, you have to write the ***correct chemical formula*** first. Then you **add up** *the relative atomic masses* (RAM) of the elements *that make up the compound to get the relative formula mass*. *Table 4.1* will be useful to you in working out RFM.

Element	Symbol	Atomic number	Relative atomic Number	Element	Symbol	Atomic number	Relative atomic Number
Aluminium	Al	13	27	Nitrogen	N	7	14
Argon	Ar	18	40	Oxygen	O	8	16
Beryllium	Be	4	9	Phosphorus	P	15	31
Boron	B	8	11	Potassium	k	19	39
Calcium	Ca	20	40	Silicon	Si	14	29
Carbon	C	6	12	Sodium	Na	11	23
Chlorine	Cl	17	35.5	Sulphur	S	16	32
Flourine	Fl	9	19	Iron	Fe	26	56
Helium	He	2	4	Copper	Cu	29	64
Hydrogen	H	1	1	Zinc	Zn	30	65
Lithium	Li	3	7	Silver	Ag	47	108
Magnesium	Mg	12	24	Manganese	Mn	25	55
Neon	Ne	10	20	Chromium	Cr	24	52

Table 4.1: Symbols, Atomic Numbers and Relative Atomic Masses of some elements

Note: We always find *relative molecular mass* (RMM) of a substance if it is *molecular* and *relative formula mass* (RFM) if the substance is *ionic*.

Examples

Find the relative formula mass (RFM) or relative molecular mass (RMM) of each of the following compounds:

1. $MgSO_4$
2. NaCl
3. CO_2

Solutions

1. RFM of $MgSO_4$

	RAM	
Mg	24	1 x 24 = 24
S	32	1 x 32 = 32
4(O)	16	4 x 16 = 64
		120

Therefore, the RFM of $MgSO_4$ is 120.

2. RFM of NaCl

	RAM	
Na	23	1 x 23 = 23
Cl	35.5	1 x 35.5 = 35.5
		58.5

The RFM of NaCl is 58.5

3. RFM of CO_2

	RAM	
C	12	1 x 12 = 12
2(O)	16	2 x 16 = 32
		44

The RMM of CO_2 is 44

Exercise 2

Use *Table 4.1* for RAMs of the elements to answer the following questions:

1. Find the RFM or RMM of the following ionic compounds:
   (a) Na_2CO_3
   (b) $LiNO_3$
2. Find the RMM of the following molecular compounds:
   (a) $C_2H_{12}O_6$
   (b) CCl_4

The Mole

In English Language, we use certain terms to represent quantities. For example, *a pair of socks, a dozen of eggs, a unit of mangoes, a decade, a century, etc*. Look at *Table* 4.2 to see their meanings.

Term	Meaning
A pair of socks	2 socks
Three pairs of mangoes	6 mangoes
A dozen of eggs	12 eggs
Two dozens of shirts	24 shirts
A unit of oranges	10 oranges
Two units of notebooks	20 notebooks
A decade	10 years
A century	100 years

Table 4.2: terms representing quantities

Likewise, in chemistry, we use the term **mole** to represent ***a quantity***. Mole is an ideal measure since *molecules, atoms and ions have very small masses which do not help to make our calculations easily*. When some scientists carried out an experiment involving masses of compounds or elements, they use grams to represent their quantities. When RAM, RMM and RFM of substances are given in grams, it is found that all give the same number of particles 602 300 000 000 000 000 000 000. This is the number which is called ***a mole***. *Is it practical to work with such a number?*

Why do we use mole to mean quantity in chemistry? What makes up one mole?

This large number of particles is simplified by presenting it to three significant figures as 6.02×10^{23}. This number was first discovered by an Italian chemist ***Amedev Avogadro***. It was later named ***Avogadro's number*** or ***constant*** to honour him.

A mole is used to replace the terms *atomic mass*, *molecular mass* and *formula mass units*. One mole therefore is the:

- ***Amount of a substance which contains an Avogadro's number of particles 6.02×10^{23}***
- ***Atomic mass of an element in grams***
- ***Molecular or formula mass in grams***

Study *Table* 4.3 to understand this concept of a mole better.

Particle	Symbol or formula	RAM RMM RFM	Number of moles	Number of particles	Mass in grams
Atoms/ Elements	C	12	1	6.02×10^{23}	12g
	O	16	1	6.02×10^{23}	16g
	N	14	1	6.02×10^{23}	14g
	N_2	28	1	6.02×10^{23}	28g
	Mg	24	1	6.02×10^{23}	24g
	O_2	32	1	6.02×10^{23}	32g
Compound	CO_2	44	1	6.02×10^{23}	44g
	Na_2O	62	1	6.02×10^{23}	62g
	$Mg\ SO_4$	120	1	6.02×10^{23}	120g

Table 4.3: Relationship between molar mass and Avogadro Number

Definition of a mole

A mole of a substance is a quantity containing 6.02×10^{23} particles with their RAM, RMM and RFM expressed in grammes.

> **Note:** One mole of any substance is the same as RAM, RMM or RFM of the substance but in grammes. This is called the **molar mass** of a substance

Using the mole in calculations

Now that you know that 1 mole = 6.02×10^{23} = RFM (RMM, RAM) ***in grammes***, there are several formulae you can derive from this relationship. You can use the following formulae:

1. **No. of moles of atoms (N)** = $\dfrac{\textit{Mass of an element in grammes}}{\textit{Molar mass in grammes}}$

$$N = \frac{M}{RAM}$$

You can use this formula to convert mass of a substance to the number of moles or vice versa.

Example 1

How many moles of sodium atoms are there in 69g of sodium metal?

Solution

RAM of Na = 23

1 mole of Na = 23g

∴ in 69g of Na there are = *m*/RAM

$= \dfrac{69g \times 1 \text{ mole}}{23g}$

= **3 moles**

2. **No. of moles of molecules or ions** = $\dfrac{\text{Mass of a substance in grammes}}{\text{Molar Mass in grammes}}$

Example 2

Work out the masses of the following gases:

1. 2 moles of oxygen gas
2. 5 moles of Nitrogen gas

Solution

1. RMM of O_2 = 2 x 16 = 32

∴ 1 mole O_2 = 32 g

2 moles O_2 = ?

∴ $\dfrac{\textit{2 moles} \times 32\ g}{\textit{1 mole}}$

= 2 x 32 g

= **64 g**

2. RMM of N_2 = 2 x 14 = 28

$\therefore$ 1 mole N_2 = 28 g

5 moles N_2 = ?

$$\frac{5\ moles \times 28\ g}{1\ mole}$$

= **140 g**

3. **No. of moles of particles** = $\dfrac{\textit{No of particles}}{\textit{Avogadro number}}$

You can use this formula to find the number of moles of particles when you are given their quantity (*number*) or vice versa.

Example 1

Taking the Avogadro number as 6×10^{23}, work out the number of moles in:

1. 6,000,000,000 **Na** atoms
2. 12,000,000,000,000 **sulphur** atoms

Solution

1. 6×10^{23} = 1 mole

   6,000,000,000 = 6×10^{9}

   6×10^{9} = ?

   $$\therefore \frac{(6 \times 10^{9}) \times 1\ mole}{6 \times 10^{23}}$$

   $$= \frac{\mathbf{1\ mole}}{\mathbf{10^{14}}} \quad \text{Or} \quad \mathbf{10^{-14}\ moles}$$

2. 12,000,000,000,000) = 1.2×10^{13}

   $$\therefore \frac{1.2 \times 10^{13} \times 1\ mole}{6 \times 10^{23}}$$

   = 0.2×10^{-10} mole

   = **2×10^{-11} mole**

Example 2

Work out the number of moles that are in 32g of sodium hydroxide (NaOH).

Solution

RFM of NaOH = 23 + 16 + 1 = 40

$\therefore$ 40g NaOH = 1 mole

32g NaOH = ?

No of moles NaOH = $\dfrac{32g\ NaOH \times 1\ mole}{40\ g}$

= **0.8 moles**

Exercise 3

1. What is a mole?
2. What number is equivalent to
   a) 1 mole
   b) 0.5 moles.
3. Work out the number of moles in the following:
   a) 168g of CaO
   b) 9g of H_2O
4. Find the number of moles in
   a) 1.8×10^{24} molecules of CO_2
   b) 3×10^{23} ions of Na

Molar volume of gases

Many substances exist as gases. Gases are difficult to weigh or work with because they easily expand and contract. If we want to find the number of moles of gases, we have to measure their *volume* rather than their *mass*. We can measure the volume of a gas using a *syringe*, *measuring cylinder* or *graduated gas jar* as shown in *Figure 4.1.*

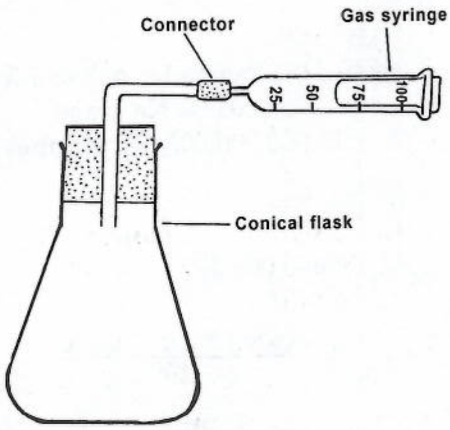


Fig 4.1 : Apparatus for measuring and collecting gas

Chemists work with masses and volumes of gases at two conditions, namely:

1. At *standard temperature* (0°C) *and pressure 1 atmosphere or 760mm Hg* (STP): 1 mole of any gas occupies a volume of 22.4dm³. This is called *molar volume of gas at STP*.
2. At *room temperature* (25°C) *and pressure 1 atmosphere or 760 mm Hg* (RTP): 1 mole of any gas occupies a volume of 24 dm³. This is called *the molar volume of a gas at RTP*.

Example 1

Calculate the number of moles of ammonia gas NH_3 occupying a volume of 44.8 dm³ at STP.

Solution

At STP 22.4 dm³ of NH_3 gas = 1 mole

44.8 dm³ = ? more

$$\therefore \frac{44.8\ dm^3 \times 1\ mole}{22.4\ dm^3}$$

= **2 moles**

Example 2

Calculate the volume of 3 moles of carbon dioxide gas (CO_2) at RTP.

Solution

At RTP 1 mole of a gas occupies 24 dm^3

∴ 3 moles of a gas occupy **?**

$\therefore \frac{3 \text{ moles} \times 24 \text{ dm}^3}{1 \text{ mole}}$

$= \underline{\mathbf{72\ dm^3}}$

Example 3

Calculate at STP the volume occupied by 12 g of oxygen gas.

Solution

At STP 1 mole of O_2 will occupy 22.4 dm^3 volume

But 1 mole O_2 = 2 x 16 = 32g

∴ 32g of O_2 occupies 22.4 dm^3

12g of O_2 will occupy?

$\therefore \frac{12 \text{ g} \times 22.4 \text{ dm}^3}{32 \text{ g}}$

$= \underline{\mathbf{8.4\ dm^3}}$

Exercise 4

1. What is the value of molar volume of any gas at
   a) STP?
   b) RTP?
2. Calculate the volume occupied by
   a) 11g of CO_2 gas at STP
   b) 7g of N_2 gas at RTP
3. Work out the masses of the following gases at STP
   a) 11.2 dm^3 of CO_2
   b) 33.6 dm^3 of O_2

Percentage of water in hydrated ionic and molecular compounds

Crystals of some compounds, such as sugar and copper II sulphate, contain water, which becomes part of the substance.

1. Water in sugar

In the following activity, you will investigate what happens when sugar is heated. You will later be able to find *the percentage of water in sugar.*

Activity 1

Aim: *To investigate percentage of water in sugar*

Materials:

- *Dry test tube*
- *Gauze wire*
- *Tripod stand*
- *Bunsen burner*
- *Accurate balance*
- *Test tube holder*
- *20 g of sugar*

Procedure:

1. *Weigh a dry test tube*
2. *Add 20g of sugar crystals; accurately weighed into the test tube*
3. *Heat the crystals strongly until there is no further visible change*
4. *Cool and weigh*
5. *Heat the contents strongly again for 1 minute*
6. *Cool and reweigh*
7. *If there is no more mass change subtract the **final mass of the test tube + residues** from the **initial mass of test tube + sugar** to find **the mass of water that has evaporated.***
8. *Finally find the percentage of water in sugar using the formula below:*

$$\textit{Percentage of water} = \frac{\textit{Mass of evaporated water}}{\textit{Mass of sugar}\text{ (before heating)}} \times 100$$

Observation / Results:

1. *On heating sugar, what did you observe?*
2. *If there are any residues, what do you think they made up of?*

Conclusion

1. What percentage of sugar is water?
2. You have observed steam rising as the sugar is heat. The colour of sugar also changes to black. The black substance that finally becomes residues is made up of *carbon*.
3. The equation below shows the results of disintergration of the sugar after heating:

$$\underset{\text{sugar}}{C_{12}H_{22}O_{11}} \xrightarrow{\text{Heat}} \underset{\text{carbon}}{12C_{(s)}} + \underset{\text{water}}{11H_2O_{(g)}}$$

2. Water in hydrous copper II sulphate

You can also investigate *the percentage of water in hydrous copper II sulphate*. Hydrous copper sulphate is *blue*. *Activity* 2 will help you carry out this investigation.

Activity 2

Aim: *To investigate the percentage of water in copper II sulphate*

Materials:

- *eye goggles*
- *A pair of tongs*
- *Tipod stand*
- *dry test tube*
- *Accurate balance*
- *Copper II sulphate crystals*
- *Bunsen burner*
- *Wire-gauze*

Procedure:

1. *Weigh a dry test tube*

2. *Weigh about 5g of copper II sulphate crystals into the test-tube*
3. *Heat the crystals strongly until there is no visible change*
4. *Cool and reweigh*
5. *Heat the test-tube with its contents strongly again for 1 minute*

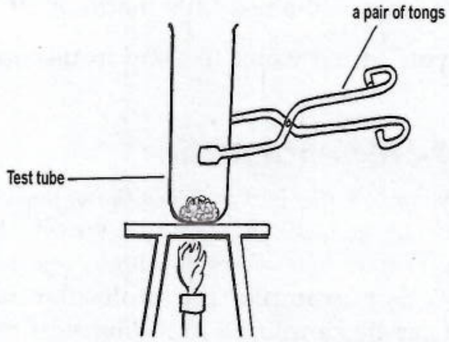


Figure 4.2: heating copper II sulphate

6. *Cool and reweigh*
7. *Repeat steps 5 to 6 until there is no further change in mass*
8. *Calculate the mass of residues and hence the mass of water lost from the crystals*
9. *Find the percentage of water in copper II sulphate*

$$\textit{Percentage of water} = \frac{\textit{Mass of evaporated water}}{\textit{Mass of copper II sulphate before heating}} \times 100$$

Results / Observations:

What colour did the residues become when you completely heat the copper II sulphate?

Conclusion

1. What is the percentage of water in copper II sulphate?
2. You have observed that, when you completely heat copper II sulphate, it turns into a white powder. This powder is called *anhydrous copper II sulphate*. The change of colour comes about because of the *loss of water from hydrous copper II sulphate*.
3. The molecules of sugar are made up of *carbon*, *hydrogen* and *oxygen atoms* as in the formula $C_{12}H_{22}O_{11}$. When sugar is heat the molecules break up. *Hydrogen and oxygen atoms combine to form water*. Carbon remains in the test tube as the black residues.
4. The water that evaporate from the copper II sulphate exist inside the molecules of $CuSO_4$. The water gets incorporated into the molecules of the compound during the formation of the crystals. This water is called *water of crystallization* and the compound is said to be *hydrated*. Hydrated or hydrous copper II sulphate has the formula $CuSO_4.5H_2O$. The $5H_2O$ is part of the molecule. When heat the water evaporates to leave behind a white powder of *anhydrous copper II sulphate*:

$$CuSO_4.5H_2O \xrightarrow{\text{Heat}} CuSO_4 + 5H_2O$$

Exercise 5

1. 50 g of glucose $C_6H_{12}O_6$ was strongly heat in a dry test-tube until a black stuff appeared
   a) work out the equation of this decomposition of glucose.
   b) What is the black residue in the test tube made up of?
2. What would happen if you added water to anhydrous copper sulphate?

Empirical and molecular formula

So far you have seen that a *molecular formula represents the numbers and types of atoms that bond together to form a molecule*. Hence, the empirical formula of a compound *is the simplest formula that shows the atoms that make up the compound in their lowest ratio*. For example, the molecular formula of *benzene* is C_6H_6. This formula can be simplified in its simplest ratio. To do this you have to divide the numbers of the atoms by their lowest number. So in C_6H_6, we will divide each subscript by **6**: $C_{6/6}$, $H_{6/6}$ becomes **CH**. Therefore, CH *is the empirical formula of benzene*.

What is the empirical formula of a compound?

Find the empirical formula of $C_4H_8O_2$.

Compare your answer to the following solution:

In $C_4H_8O_2$, the lowest value is **2**. So we divide each value with 2.

$C_{4/2}H_{8/2}O_{2/2} = C_2H_4O$

So the simplest ratio of the atoms **C:H:O** is **2:4:1**.

The empirical formula of $C_4H_8O_2$ is, therefore, C_2H_4O.

Calculating empirical formula

1. Using masses of elements in the compound

We can work out the empirical formula of a compound if we know the masses of the elements that make up a compound. The masses can be found by using an experiment.

Finding empirical formula of magnesium oxide experimentally

When magnesium ribbon is heat strongly, it burns very brightly to form a white powder called *magnesium oxide*. The data *Table 4.6* was obtained from an experiment to find the formula for magnesium oxide.

Mass of crucible	14.63 g
Mass of crucible + magnesium	14.87 g
Mass of crucible + magnesium oxide	15.03 g
Mass of magnesium used (14.87g - 14.63g)	0.24 g
Mass of oxygen which reacted with magnesium (15.03g - 14.87g)	0.16 g

Table 4.6: Data of an experiment

The experiment was carried out using the apparatus shown in *Figure 4.3*:

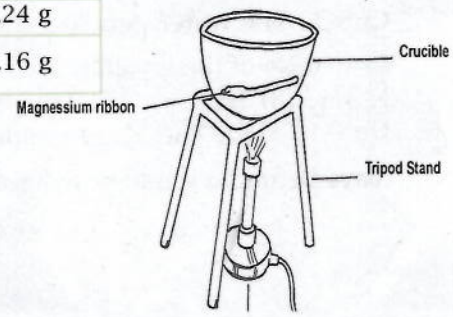


Figure 4.3 aparatus used in burning magnesium ribbon

Looking at *Table 4.6*, it is clear that 0.24 g of magnesium reacted with 0.16g of oxygen to form magnesium oxide.

Mg	*O*
0.24 g	0.16g

Then find the number of moles of each element **M/RAM**:

	Mg	*O*
	$\frac{0.24g \times 1 \text{ mole}}{24g}$	$\frac{0.16g \times 1 \text{ mole}}{16g}$
	= 0.01	= 0.01
Then find ratio of the moles	$= \frac{0.01}{0.01}$	$= \frac{0.01}{0.01}$
	= **1**	= **1**

From the ratio of moles, we can now write the empirical formula of the compound (*magnesium oxide*). Since the ratio of number of moles of magnesium to that of oxygen is **1:1**, then the empirical formula is MgO.

Example

In an experiment an unknown organic compound was found to contain 0.12g of carbon and 0.02 g of hydrogen. Work out the empirical formula of the compound

Solution

	C	H
Masses of elements in the compound	0.12g	0.02g
Number of moles of the elements	$\frac{0.12g \times 1 \text{ mole}}{12g}$	$\frac{0.02g \times 1 \text{ mole}}{1g}$
	= 0.01	= 0.02
The ratio of the moles	$= \frac{0.01}{0.01}$	$= \frac{0.02}{0.01}$
	= **1**	= **2**

Therefore the empirical formula is CH_2

2. Using percentage by mass of the elements

The empirical formula can also be worked out using the percentages by mass of the elements that make up a compound. There is no major difference from the first method, since in this method we take the *percentage as mass of the elements*.

Example 1

Work out the empirical formula of the compound with the following percentage composition by mass of elements: **C** = 40%, **H** = 6.67%, **O** = 53.33%.

Solution

	C	:	H	:	O
Percentages of elements	40	:	6.67	:	53.33
Divide the percentages by the RAM to find the number of moles	$\frac{40}{12}$	:	$\frac{6.67}{1}$	:	$\frac{53.33}{16}$
=	3.33	:	6.67	:	3.33
Divide the moles by their lowest value	$\frac{3.33}{3.33}$	:	$\frac{6.67}{3.33}$	:	$\frac{3.33}{3.33}$
=	1	:	2	:	1

∴ The empirical formula of the compound is CH_2O

Working out molecular formula of a compound from its empirical formula

The molecular formula *is a true formula of a compound*. It shows the elements making up the compound in their right numbers. For example, $C_6H_{12}O_6$ is a *molecular formula of glucose*. We can work out the molecular formula of a compound using its empirical formula if we know its ***relative molecular mass*** or ***relative formula mass***.

Example 2

Calculate the empirical formula of an organic compound containing 92.3% carbon and 7.7% hydrogen by mass. If the RMM of the original compound is 78. What is its molecular formula? (RAM: **C** = 12, **H** = 1)

Solution

	C	:	H
% by mass	92.3	:	7.7
No of moles	$\frac{92.3}{12}$	:	$\frac{7.7}{1}$
=	7.7	:	7.7
Ratio of moles =	**1**		**1**

∴ Empirical formula is CH.

But the mass of the empirical formula unit = 12 + 1 = 13g

Mass of original compound = 78g

∴ No. of empirical formula units present in 78g

$$= \frac{\textit{Mass of a compound}}{\textit{Mass of empirical formula}}$$

$$= \frac{78}{13}$$

$$= 6$$

∴ The molecular formula of the organic compound is 6 x (*empirical formula*)

= 6 x (CH)

= C_6H_6 (Benzene)

Exercise 6

1. What is the empirical formula?
2. a) An oxide of hydrogen has a percentage composition by mass of H = 5.9% and O = 94.1%. Its relative formula mass is 34. Calculate:
   i) the empirical formula of the oxide
   ii) the molecular formula of the oxide

   b) What is the name of the oxide in (a)?

Concentration of a solution

Many chemical reactions take place in *aqueous solutions*. The *solutes* of these solutions are the *reactants*. It is, therefore, important to know the *concentration of the solution* you are working with. Concentration of a solution is *the amount of solute dissolved in a specific volume of solvent*.

Ways of expressing concentration of a solution

There are two ways of expressing concentration of a solution:

1. Ratio of mass of solute to volume of solvent

This is the *number of grammes of solute dissolved in 1 dm³ of a solvent*. This will be expressed as grammes per cubic decimeter (g/dm^3)

Example

If 20g of sodium chloride, NaCl, is dissolved in $5dm^3$ of water, what will be the concentration of the solution in g/dm^3

Solution

$$\text{Concentration in } g/dm^3 = \frac{\text{mass } (g)}{\text{volume } (dm^3)}$$

$$= \frac{20\text{ g}}{5dm^3}$$

$$= 4\ g/dm^3$$

2. Ratio of number of moles of solute to volume of solvent

This is the *number of moles of solute dissolved in 1 dm³ of a solvent*, expressed as moles per cubic decimeter (m/dm^3). This is called *molarity*. Molarity of a solution is the concentration of a solution expressed in moles per cubic decimetre ($moles/dm^3$).

Example 2

- 2 moles of a solute dissolved in dm^3 of water, implies that, a solution has a molarity of 2 $moles/dm^3$. This can also be written as $2mol/dm^3$ or 2molar or 2M.
- 3 moles of a solute dissolved in $1dm^3$ of solution has a molarity of 3 mol/dm^3 or 3M
- 0.5 moles of a solute dissolved in m/dm^3 of solution has a molarity of 0.5 $moles/dm^3$ or 0.5M

Note: $1dm^3 = 1000\ cm^3 = 1000ml = 1$ litre

Example 2

If 2 moles of sucrose is dissolved in 10 dm^3 of water, what will be the concentration o f the solution in moles/dm^3.

Solution

$$\text{Molarity} = \frac{\text{Number of moles}}{\text{Volume } (dm^3)}$$

$$= \frac{2 \text{ moles}}{10 \; dm^3}$$

$$= \mathbf{0.2 \; moles/dm^3} \quad \text{or} \quad \mathbf{0.2M}$$

Example 3

What would be the concentration of a solution in mol/dm^3 when 0.2 moles of a solute dissolves in 0.25 dm^3?

Solution

$$\text{Molarity of a solution} = \frac{\text{No. of moles}}{\text{Vol } (dm^3)} = \frac{0.2 \text{ moles}}{0.25 \; dm^3} = \underline{\mathbf{0.8M}}$$

From the two expressions of concentration of a solution we have looked at, you can see that there are several calculations we can make involving *mass*, *moles* and *molarity*. Sometimes you will need one quantity in order to work out the other. For example, *you will need mass of the solute in order to work out the number of moles* or vice versa. You will also need moles of a solute in order to work out the molarity of a solution or vice versa.

1. Calculating molarity from number of grammes of solute dissolved

Example

Calculate the molarity of a solution made when 4.0g sodium hydroxide (NaOH) is dissolved in 250 cm^3 of solution (Na = 23, O = 16, H = 1):

Solution

Step 1: Find the number of moles from the mass of the solute:

$$\text{No. of moles} = \frac{\text{mass (g)}}{\text{Molar mass } (g \; mol^{-1})}$$

$$\text{RFM of NaOH} = 23 + 16 + 1 = 40 \; (\textit{molar mass})$$

$$= \frac{4.0g \times mol}{40g}$$

$$= \underline{\mathbf{0.1 \; moles}}$$

Step 2: Convert the volume to dm^3 if it is in other units.

Converting 250 cm^3 to dm^3

$$= \frac{250 \; cm^3 \times dm^3}{1000 \; cm^3}$$

$$= \underline{0.25dm^3}$$

Step 3: Divide the number of moles of the solute by the volume in dm^3 to find the molarity of the solution:

$$= \frac{0.1 \text{ moles}}{0.25 dm^3}$$

$$= \underline{0.4 \text{ mol/dm}^3}$$

This is in agreement with the formula:

$$Molarity = \frac{\text{No of moles of the solute}}{\text{Volume of the solvent}}$$

2. Calculating number of moles of the dissolved solute

Sometimes you are given a solution of a known molarity and volume. *How can you work out the number of moles of the solute dissolved in it?*

Example

Work out the number of moles of sodium hydroxide dissolved in $200cm^3$ of a solution of concentration 0.33M

Solution

Molarity = $0.33 mol/dm^3$
Volume = $200cm^3$

$$\therefore \text{No moles} = \frac{0.33 \text{ mol/dm}^3 \times 200cm^3 \times 1dm^3}{1000cm^3}$$

$= 0.066$
$= \underline{\mathbf{0.07 \text{ moles}}}$

The formula we have used in this example is: *No of moles* = *Molarity* x *volume*

3. Calculating concentration of solution in g/dm^3 given its molarity

Concentration of a solution in g/dm^3 can easily be worked out if its molarity is known. This is done by *multiplying the molar mass of the solute by the molarity of the solution*.

Example

Work out the concentration in g/dm^3 of 0.08M sodium carbonate (Na_2CO_3) solution.

Solution

Molarity = $0.08 mol/dm^3$

But 1mole Na_2CO_3 has molar mass of $(23 \times 2 + 12 + 3 \times 16) = 46 + 12 + 48 = 106g$

$\therefore$ 0.08 mole Na_2CO_3 has mass of $\frac{0.08 \text{ mole} \times 106g}{1\text{mole}}$

$\therefore$ *concn* = $\underline{\mathbf{8.48 \ g/dm^3}}$

In short, you can use the formula: *Concentration in* g/dm^3 = *molarity* x RFM

Exercise 7

1. What is molarity of a solution?
2. Calculate the molarity of the following solutions:
   a) 20 g of NaOH dissolved in $500cm^3$ of solution
   b) 3.5 g of NaCl dissolved in $600cm^3$ of solution.
3. Work out the number of moles of solute dissolved in the following solutions:
   a) $500cm^3$ of O.2M LiCl solution
   b) $5dm^3$ of 0.1M $CaSO_4$ solution.

Preparing standard solution

A standard solution is *a solution of a known concentration*. For example, 0.5M NaCl solution is a standard solution because we know its concentration.

What is a standard solution? Is a sugar solution of concentration $2g/dm^3$ a standard solution?

Standard solutions are very important as they enable you to see how to handle or use them. Secondly, they are important because we can use them to determine concentrations of other solutions.

1. Preparing standard solutions by dissolution

In this method, you *measure out an exact mass of a solute* and *dissolve it in the required volume* of a solvent. We *use an accurate balance to measure the exact mass* and *accurate volumetric apparatus* to measure accurate volume.

When preparing a standard solution, there are several steps you should follow carefully, such as:

Step 1: You need to know how much solute is required to make the solution

Step 2: Weigh the required mass of solute in the beaker using an accurate balance as in *Figure 4.5* (a).

Step 3: Add distilled water to the beaker and stir with a rod.

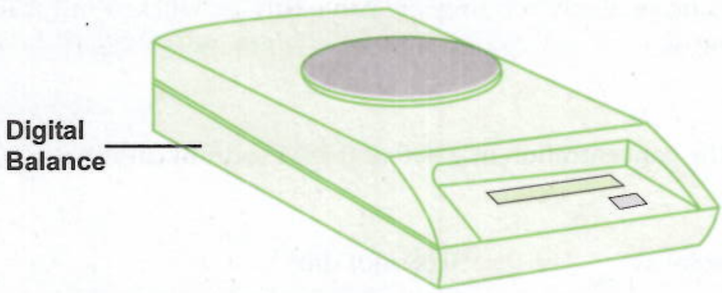


Figure 4.4: Apparatus used in standard solution preparation

Step 4: Using a filter funnel, transfer the salt solution into a 1 dm^3 volumetric flask.

Step 5: Rinse the beaker and funnel with plenty of distilled water.

Step 6: Then add more distilled water to the graduated mark.

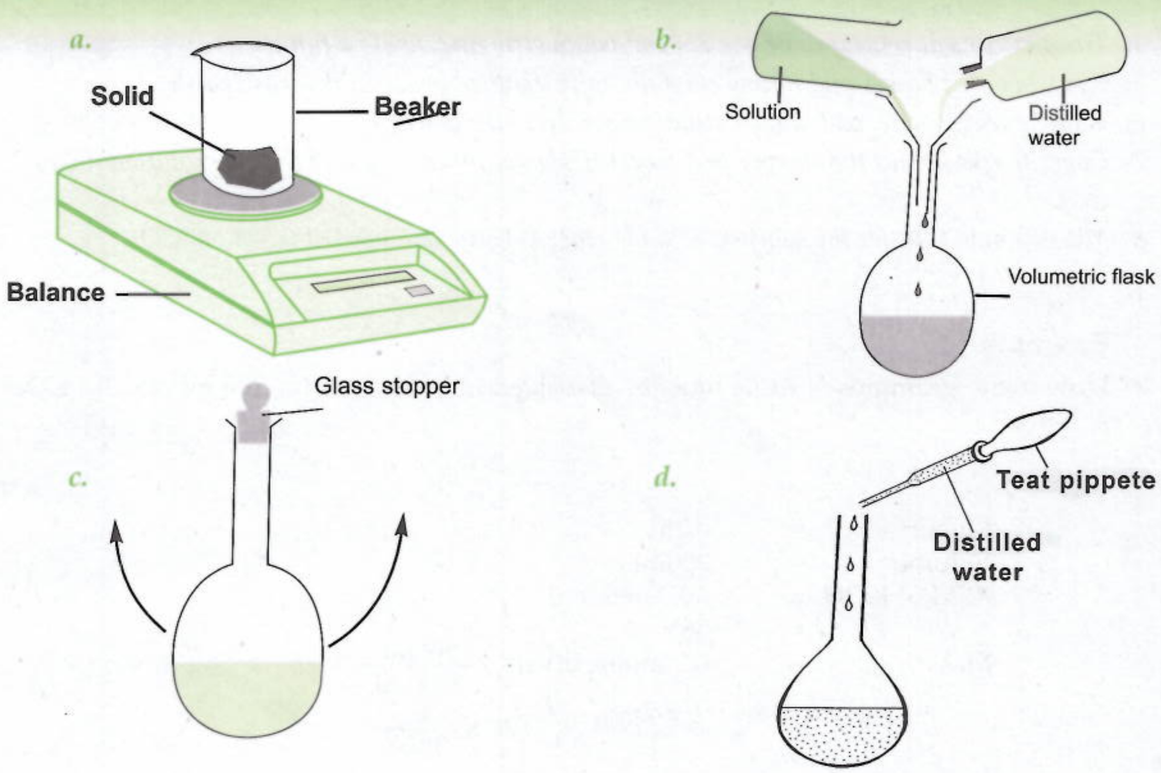


Figure 4.5 Preparation of a standard solution

In *Activity* 3, you will practice how to make standard solutions by dissolution.

Activity 3

Aim: *To prepare $250cm^3$ Na_2CO_3 Solution of 0.5M concentration by dissolution.*

Materials:

- *balance*
- *wash bottle*
- *funnel*
- *$250cm^3$ volumetric flask*
- *distilled water*
- *sodium carbonate*
- *glass-rod stirrer*
- *beaker*

Procedure:

1. *Work out the exact mass you need. You do this by using the formula:*

   mass = molarity x volume x RFM

   But ***molarity*** *= 0.5 mol/dm^3,* ***vol*** *= $250cm^3$,* ***RFM*** ***of Na_2CO_3***

   $= 2 \times 23 + 12 + 3 \times 16 = 46 = 12 = 48 = 106$

   $\therefore$ *the mass required* $= \dfrac{0.5\ mol/dm^3 \times 250cm^3 \times 106 \times 1dm^3}{1\ mol \times 1000cm^3}$

   $= 0.12 \times 106\ g$

   $\ldots. = 12.25\ g$

2. *Weigh 12.25 g of sodium carbonate crystals in a beaker*
3. *Add distilled water into the beaker and stir continuously until all the salt dissolves*

4. *Transfer the salt solution into the 250cm³ volumetric flask using a funnel*
5. *Rinse both the beaker and funnel carefully with distilled water in the wash bottle*
6. *Using a teat pipette, add more distilled water up to the mark*
7. *Cover the flask with the stopper and invert it several times to ensure that the solution mixes well.*
8. *You can now transfer the solution into the reagent bottle and mark it 0.5M* Na_2CO_3

Example

How many grammes of KOH must be dissolved in order to make 200 ml of 0.2M KOH solution?

Solution

Molalrity	=	0.2M
Volume	=	200ml
RFM of KOH	=	39 + 16 + 1
	=	56
Mass	=	$0.2 \text{ moles/litre} \times \frac{200\text{ml}}{1000\text{ml}} \times l \times 56\text{g/mole}$
	=	$\frac{0.2 \text{ mole}}{1\,l} \times 0.2l \times \frac{56\text{g}}{\text{mole}}$
	=	$0.2 \times 0.2 \times 56\text{g}$
	=	**0.24g**

2. Preparing standard solution by dilution

Standard solution can be prepared by diluting a solution of known concentration. we do this by adding more solvent in order to reduce the concentration of the original solution.

You will find the formula below very useful when preparing standard solution by dilution.

$M_1 \times V_1 = M_2 \times V_2$ where M_1 = *initial concentration*

V_1 = initial volume

M_2 = final concentration

V_2 = final volume

Example

A laboratory assistant has 1M NaOH solution. Describe how he would prepare 250cm³ of 0.5M NaOH solution from the standard solution.

Solution

$$M_1 \times V_1 = M_2 \times V_2$$

$$M_1 = 1\text{mol/dm}^3$$

V_1 = ? (*not known*)
M_2 = $0.5\ mol/dm^3$
V_2 = $250cm^3$ (*substitution*)
$1mol/dm^3 \times V_1 = 0.5mol/dm^3 \times 250cm^3$

Divide both sides by $1mol/dm^3$ to find V_1

$$\frac{1moldm^3 \times V_1}{1\ mol/dm^3} = \frac{0.5\ mol/dm^3 \times 250cm^3}{1\ mol/dm^3}$$

$1mol/dm^3 = 1mol/dm^3$
V_1 = $\underline{\mathbf{125cm^3}}$

Then measure exactly $125cm^3$ of the standard solution. Place it in a $250cm^3$ volumetric flask and make it up to $250cm^3$ mark using distilled water.

Example 1

Work out the volume of 0.8M NaOH required toprepare the 250 ml solution of 0.2M NaOH.

Solution

$M_1V_1 = M_2V_2$
M_1 = 0.8M
V_1 = ?
M_2 = 0.2M
V_2 = 250 ml

Substuting in the equation:

$0.8M \times V_1 = 0.2M \times 250\ ml$

Making V1 the subject of the formula:

$$V1 = \frac{0.2M \times 250\ ml}{0.8M}$$

$= \underline{\mathbf{62.5\ ml}}$

This means that you have to:

a. Measure 62.5 ml of the original (stock) solution using a measuring cylinder.
b. Put the volume into a connical flask
c. Add distilled water to the connical flask up to the 250 cm^3 mark.
d. Shake the solution thoroughly. This helps to distribute the concentration evenly.

Activity 4

Aim: *To prepare 0.2M HCL by dilution.*

Materials:

- *500 cm^3 volumetric flask*
- *Connical flask*
- *distilled water*
- *1 M HCl solution*
- *measuring cylinder*

Procedure:

1. *Work out the volume of 1M HCl solution required to dilute to make 500 cm³ of 0.2 M HCl.*
2. *Measure the volume you have already worked out in 1, of 1M HCl using a measuring cylinder and put it in a 500 cm³ volumetric flask.*
3. *Rinse the measuring cylinder with distilled water and add the water to the volumetric flask.*
4. *Add more water to the volumetric flaskup to the 500 cm³ mark.*
5. *Shake the solution thoroughly then label it.*

Exercise 8

1. What is a standard solution?
2. Calculate the number of moles and the number of grams of the following solutes dissolved in the solutions.
   a) 50 cm3 sodium hydroxide solution of concentration 0.1 M
   b) 15 cm3 lead nitrate solution concentration 0.6 M
3. State one difference between making standard solution by dilution and dissolution methods.
4. Explain how you would prepare 500 cm³ solution of 0.2M HCl from 1M stock solution.
5. Work out the volume of 2M LiOH required to prepare 250 cm³ of 0.5M solution.

Determining the concentration of a solution by titration

Titration is the gradual addition of one solution from a burrette to the standard solution *until the reaction between the two solutions is complete*. The best known examples of titration are the ***neutralization*** *of an acid* by a base and ***redox reactions***.

What do you think the term titration mean?

Acid-base titration

As stated earlier, we can use titration to determine the concentration of a solution of unknown concetration. For instance, you can determine concetration of a solution of hydrochloric acid using a 0.1mol/dm³ solution of sodium hydroxide. The concentration of sodium hydroxide is known but that of hydrochloric acid is not. In this titration, you add acid to the base in the flask. Before adding the acid, the base is mixed with a small volume of indicator, e.g. *phenolphthalein*. With phenolphthalein indicator, the base will change its colour to *pink*. When you add the acid, the pink colour of the base slowly disappears. *The colour will completely disappears when you added enough acid to the base*. This is called the ***end-point*** *of titration*. It is shown by the ***change of colour of an indicator*** such as phenolphthalein

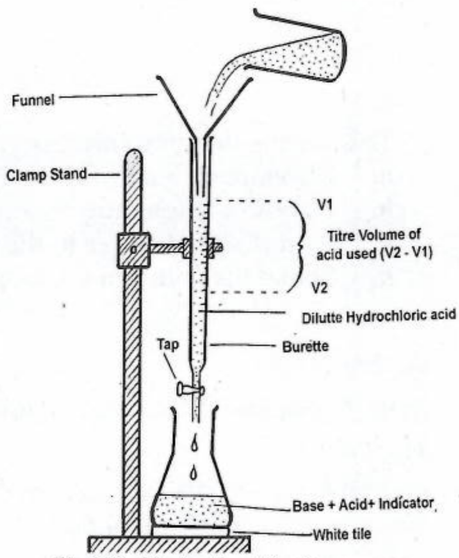


Fig 4.6: Illustrating titration process

Remember: The *neutralisation process* of an acid and base is a *chemical reaction*.

Look at this reaction equation:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl + H_2O \quad \text{..........................} \quad \textit{equation (a).}$$

From the equation, we can see that **1 mole** of HCl requires **1 mole** of NaOH to react.

Since **No. of moles** = **molarity x volume** we can then substitute the number of moles with *molarity x volume* (MV) in the chemical equation, *but on the side of reactants only*.

$$M_1 \times V_1 \text{ (NaOH)} = M_2 V_2 \text{ (HCl)}$$

Where: M_1 and V_1 are for the standard solution while M_2 and V_2 are for the other solution. V_2 is the volume of a solution which you are adding to the standard solution. *You find* V_2 *after reaching the end point*. Thereafter you work out M_2 from the equation.

Example

$30cm^3$ of 0.2M NaOH solution is used to titrate hydrochloric acid. If the volume of hydrochloric acid used is $20cm^3$, find the molarity of the acid.

$$M_1 \times V_1 = M_2 \times V_2$$

$$\therefore 0.2M \times 30cm^3 = M_2 \times 20\ cm^3$$

Divide both sides by $20cm^3$ to find M_2

$$\frac{M_2 \times 20cm^3}{20cm^3} = \frac{0.2M \times 30\ cm^3}{20cm^3}$$

$$\therefore \quad M_2 = \underline{0.3\ M}$$

Chemical equations with coefficients more than one

In *equation (a)*, the coefficients for HCl and NaOH is 1, respectively. You will sometimes be asked to titrate bases and acids as shown in *equation (b)*:

$$2HCl + Na_2CO_3 \longrightarrow 2NaCL + CO_2 + H_2O \quad \text{........................} \quad \textit{equation (b).}$$

In this equation HCl has coefficient 2. In cases where one or more reactants have coefficients more than 1, the following formula is used:

$$\frac{M_1 \times V_1}{N^1} = \frac{M_2 \times V_2}{N^2}$$

Where N_1 and N_2 are coefficients of the formulae of the respective reactant.

So $N_1 = 1$ and $N_2 = 2$

Example

$20.5cm^3$ of hydrochloric acid was required to neutralize $25cm^3$ of 0.25 M $Ca(OH)_2$ solution. Work out the molarity of the acid.

As you have already seen, it is necessary to write the balanced equation of the reaction before you start working out. This will help you to find out the number of moles of reactants used.

The *balanced equation* of a reaction between HCl and $Ca(OH)_2$ is:

$$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$$

Formula to use: $\frac{M_1 \times V_1}{N_1} = \frac{M_2 \times V_2}{N_2}$

Where M_1 = 0.25M $Ca(OH)_2$

V_1 = $25cm^3$ $Ca(OH)_2$

V_1 = 1

V_2 = $20.5\ cm^3$ HCl

M_2 = ? HCl

Substituting the above values into the equation:

$$\frac{0.25M \times 25cm^3}{2} = \frac{M_2 \times 20.5cm^3}{1}$$

By cross-multiplying you will have:

$$0.25M \times 25cm^3 \times 2 = M_2 \times 20.5cm^3 \times 1$$

$$12.5Mcm^3 = 20.5cm^3 \times M_2$$

Making M_2 the subject:

$$M_2 = \frac{12.5M\cancel{cm^3}}{20.5\cancel{cm^3}}$$

$$M_2 = \underline{0.6M}$$

Activity 4

Aim: *To find molarity of hydrochloric acid by titration*

Materials:

- *$25cm^3$ of 0.5M Na_2CO_3 in a conical flask*
- *hydrochloric acid (unknown molarity)*
- *Burette*
- *Phenolphthalein*
- *A dropper*
- *Refort stand*
- *Funnel*

Procedure:

1. *Fill the burette with dilute hydrochloric provided using a funnel*
2. *Add 2 - 3 drops of phenolphthalein indicator to Na_2CO_3 in the flask. Observe the colour change*
3. *Titrate until the end-point. Do you remember how to identify the end-point?*
4. *Record the volumes of the acid used in three trials.*

5. *Find the average volume of the acid used*
6. *Now using the formula:*

$\frac{M_1 V_1}{N_1} = \frac{M_2 V_2}{N_2}$, *find the molarity of the acid.*

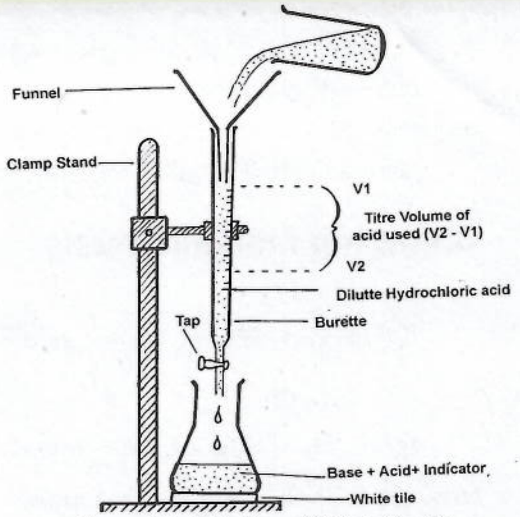


Figure 4.7: Tritration of HCl to Na_2CO_3

Exercise 9

1. 11.3cm³ sulphuric acid was used to neutralize 12.5cm³ of 0.1m sodium hydroxide solution. Calculate the concentration of sulphuric acid.
2. 20cm³ of hydrochloric acid neutralizes 25cm³ of 0.5M sodium hydroxide. Work out the molarity of the hydrochloric acid.

Determining the yield of chemical reactions

1. Calculating the theoretical yield

You will be required to work out the quantities of the yields or products in the chemical reaction. This is a very important aspect in stoichiometry. Now look at the following chemical reactions:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \quad \text{...} \quad (i)$$

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)} \quad \text{...} \quad (ii)$$

In *equation* (*i*) *one atom of carbon* requires *one molecule of oxygen* to produce *1 molecule of carbon dioxide*. In *equation* (*ii*) *two molecules of hydrogen* gas requires *one molecule of oxygen* gas *to produce two molecules of water*.

2. Changing the statements to moles

In *equation* (*i*) one mole of carbon requires one mole of oxygen gas to produce one mole of carbon dioxide gas.

Similarly in *equation* (*ii*) *two moles of hydrogen* gas requires *one mole of oxygen* gas to produce two moles of water. If *equation* (*ii*) *was multiplied by 5*, we would have:

$$10H_{2(g)} + 5O_{2(g)} \longrightarrow 10H_2O_{(l)}$$

We will say:

a. Ten moles of hydrogen gas requires five moles of oxygen gas to produce ten moles of water.

b. It is clear from the statements concerning *equation* (*ii*) that it is very important to have a *correctly balanced equation* to obtain correct quantities.

3. Changing moles to mass

a. Equation (1)

$12g$ of C + $32g$ of $O_2 \longrightarrow$ 44g of CO_2

b. Equation (2)

4g of H_2 + 32g of $O_2 \longrightarrow$ 36g of H_2O

Now let us look at another reaction.

Magnesium reacts with *oxygen* to produce *magnesium oxide* according to this equation:

$2Mg_{(s)} + O_{2(g)} \longrightarrow 2MgO_{(s)}$ .... (3)

Attempt the following questions that refer to *equation* (3):

1. How many moles of oxygen will be needed to react with magnesium in order to produce 2 moles of MgO?
2. How many moles of oxygen will be needed to produce four molecules of magnesium oxide?
3. How many moles of magnesium will react with 4 moles of oxygen gas?

If you answer the above questions correctly you should be able to follow and understand the following example:

Example

Propane (C_3H_8) burns in oxygen to produce carbon dioxide CO_2 and water H_2O according to the equation:

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O.$$

Work out:

a) the mass of CO_2 produced if 88g of propane is used.

b) the mass of propane required if 20g of water is produced.

Working

According to the equation of the reaction $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$, one mole of propane reacts with five moles of oxygen gas to produce three moles of carbon dioxide and four moles of water.

$$1 \text{ mole } C_3H_8 + 5 \text{ mole } O_2 \longrightarrow 3\text{mole } CO_2 + 4 \text{ mole } H_2O$$

or

$$44g \text{ of } C_3H_8 + 160g \text{ of } O_2 \longrightarrow 132g \text{ of } CO_2 + 72g \text{ of } H_2O$$

a) If 44g C_3H_8 produces 132g CO_2

88g C_3H_8 produces ?

$$= \frac{88g \times 132g\ CO_2}{44g}$$

= **264g of CO_2 will be produced**

b) 72g H_2O is produced from 44g C_3H_8

27g H_2O will be produced from ?

$$= \frac{27g \times 44g\ C_3H_8}{72g}$$

= **16.5g of propane will be required**

Limiting and excess reagents

Sometimes in a chemical reaction one reactant may be in short supply when the other is more than enough. The reactant which is not enough for the reaction process has a *limiting quantity* while the other is in excess. Look at *Figure 4:8:*

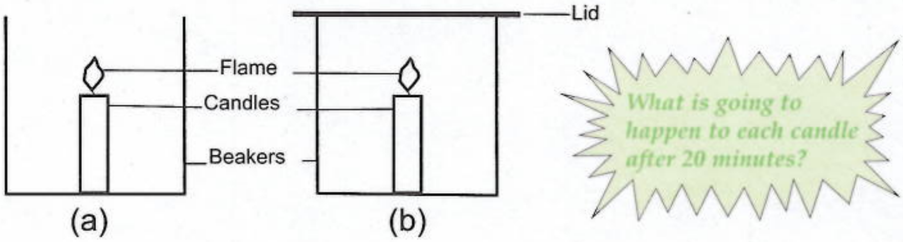


Fig 4.8 Limiting and excess reagents

It is true that candle (*a*) will still continue burning whereas candle (*b*) will get extinguished. The reason being that the air is more than enough or it is in excess in (*a*) while in (*b*) it is in short supply since the beaker is covered. The air in (*b*) has a limiting effect on the extent of the reaction process. *Limiting reactants determine the quantity of the products of the chemical reaction*.

Example

Mg burns in oxygen to produce magnesium oxide:

$$2\ Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

If 6g of magnesium is used in the reaction, what mass of magnesium oxide will be produced?

2 moles Mg produces 2 moles MgO

$2 \times 24g \longrightarrow 2(24 + 16)$

48g Mg $\longrightarrow$ 80g MgO

6g Mg $\longrightarrow$? MgO

$$\frac{6g\ Mg \times 80g\ MgO}{48g\ Mg}$$

= **10g MgO will be produced**

In the example, just given, the amount of magnesium metal used is a limiting factor of the reaction whereas oxygen is in excess.

Exercise 10

1. Calcium burns in oxygen to produce calcium oxide according to this equation

$$2Ca + O_2 \longrightarrow 2CaO \qquad (\text{RAMs: } \mathbf{Ca} = 4O, \mathbf{O} = 16)$$

a) If 10g of calcium were used, work out the mass of calcium oxide produced in the reaction.

b) If 80g of calcium were used, work out the mass of calcium oxide produced.

Unit Summary

- A chemical equation is *the short and accurate representation* of a chemical reaction. A *chemical equation must always be balanced*.
- When balancing a chemical equation, *the subscripts of the elements **do not** change*.
- Relative formula mass of a compound is worked out by adding the RAM of the element that make up the compound.
- A mole is a unit in stoichiometry. It is equivalent to a large number 6.02×10^{23}. This large number is called *Avogadro number*.
- Molar mass is the RFM or RMM or RAM of a substance in grammes. It is *the mass of one mole of a substance*.
- Molar volume is *the volume of a gas occupied by one mole* of any gas either at standard temperature and pressure (STP) or a room temperature and pressure (RTP)
- At STP (O°C and 1 atmosphere pressure), one mole of any gas has a volume of 24 dm^3.
- *Water of crystallization* of an ionic compound is the amount of water which is incorporated into the crystals as the compound crystallizes.
- Empirical formula of a compound is *the simplest formula of a compound* in which atoms are in their simplest ratio. You can work out the empirical formula from the masses or percentages by mass of elements that make up the compound.
- The concentration of a solution can be expressed as g/dm^3 **or** mol/dm^3
- Molarity of a solution is *the number of moles of the solute per dm^3* of a solution.
- Number of moles of a solute dissolved in a solution is the product of molarity and volume (No of moles = molarity x volume)
- Concentration in g/dm^3 is equal to the product of molarity and RFM of a compound (concentration in g/dm^3 = molarity x RFM)
- Standard solution is the one whose concentration is known
- Titration is *the gradual addition of one solution* from a burrette into the standard solution
- Titration is used to determine the concentration of one solution
- A reactant with a limit quantity determinies the mass of the product

Unit Exercise

1. Balance the following equations:
   a) $Mg_{(s)}$ + $HCL_{(ag)}$ ⟶ $MgCL_{2(s)} + H_{2(g)}$
   b) C_4H_{10} + O_2 ⟶ $CO_2 + H_2O$
   c) $Fe(OH)_2$ + H_2SO_4 ⟶ $FeSO_4 + H_2O$

2. Work out the Relative formula mass of the compound: (RAMs: Na = 23, C = 12, O = 16, Fe = 56, AL = 27, N = 7)
   a) Na_2CO_3
   b) feO
   c) $AL(NO_3)_3$

3. Work out the number of moles of each of the following substances:
   a) 23g of Na
   b) 135g of H_2SO_4
   c) 10g of CaO

4. How many molecule are in
   a) 11g of CO_2?
   b) 2g of O_2?

5. Calculate at STP
   a) The density of oxygen gas
   b) The mass of $100dm^3$ of hydrogen gas
   c) The volume occupied by 146g of HCl.

6. Determine the empirical formula of
   a) An oxide of calcium formed when 0.4g of calcium reacts with 0.16g of oxygen.
   b) An organic compound which contains 80% by mass of carbon and 20% by mass of hydrogen. If the RFM of the compound is 30, what is the molecular formula of the compound?

7. Calculate mass in grammes of $Ca(OH)_2$ contained in each of the following:
   a) $2\ dm^3$ of 1M $Ca(OH)_2$
   b) $500cm^3$ of O.4M $Ca(OH)_2$

8. Calculate the molarity of each of the solutions:
   a) 14.8g of $Ca(OH)_2$ per dm^3
   b) 60g of Na OH in $3dm^3$
   c) 11.2g of KOH in $200cm^3$

9. Explain how the following solutions can be prepared
   a) 3 litres of 1.5M $CuSO_4$ solution
   b) 2 litres of 1.0M NaOH solution

10. Work out the mass of
    a) KOH you would dissolve to make a 250 cm^3 solution of 0.4M KOH

b) $CuSO_4$ you would dissolve to make a 100 cm^3 solution of 0.2M $CuSO_4$.

11. Describe how you would prepare a 500 cm^3 solution of 0.2M KCl by dissolution.

12. Explain how you would prepare:
    a) 200 cm^3 solution of 0.5M HCl from 2M stock solution.
    b) 150 cm^3 solution of 0.1M NaOH from 1M stock solution.
    c) 250 cm^3 solution of 0.3M HSO_4 from 1M stock solution.

13. 10cm^3 of dilute sulphiric acid neutralises 12cm^3 of sodium hydroxide whose concentration is 0.2 moles. Describe an experiment that would assist you determine the concentration of the acid. What is the concentration of the acid by calculation?

14. 69g of sodium burns in excess oxygen to produce sodium oxide:

    $$Na + O_2 \longrightarrow Na_2O.$$

    Find the mass of sodium oxide Na_2O produced in the reaction.

Unit 5

Heats of Reactions

Success Criteria

By the end of this unit you should be able to:

1. Define the terms *exothermic* and *endothermic* in relation to heat changes
2. Describe temperature changes in exothermic and endothermic reactions and processes
3. Draw energy level diagrams for exothermic and endothermic reactions
4. Describe energy changes involved in bond breaking and bond formation processes
5. Determine whether the reaction is exothermic or endothermic using bond energies

Key words:

In this unit you will find these key terms and concepts:

exothermic, endothermic, thermo-chemical reactions, energy level diagram, enthalpy, bond energy

Ensure that you understand and learn how to apply them both for your academic and real life situations.

Enthalpy of a substance

This is the amount of heat energy contained in a substance. The symbol for enthalpy is H. All matter contains energy in the form of *kinetic energy* and *potential energy*. Enthalpy of a substance can change due to various factors.

Formation and breaking of bonds, *change of state of matter* and *dissolution of substances* in water cause changes in enthalpy of that substance. Enthalpy change is denoted by the symbols, ΔH. The change itself is denoted by the symbol Δ. The heat change is calculated as *the difference between the heat content of the products of the reaction and that of the reactants*.

$$\text{Enthalpy change} = \text{Enthalpy of products} - \text{Enthalpy of reactants}$$

$$\Delta H = H_2(\textit{products}) - H_1(\textit{reactants})$$

$$\Delta H = H_2 - H_1$$ where H_2 and H_1 are enthalpies of products and reactants, respectively

For example, in the combustion of carbon in air:

$$\underset{\text{Reactants}}{C_{(s)} + O_{2(g)}} \longrightarrow \underset{\text{product}}{CO_{2(g)}}$$

The *heat change* will be found as follows:

$$\Delta H = H_2(CO_2) - H_1(C_{(s)} \quad + \quad O_{2(g)})$$
$$= H_2(CO_2) - H_1(C_{(s)}) - H_1(O_{2(g)})$$

Combustion** is a chemical change that **releases heat energy** to the surrounding*. There are also other processes, both physical and chemical, that affect energy change in the surroundings. Chemical reactions that cause change in temperature are called ***thermo-chemical reactions. In this unit we will look at how chemical reactions and processes cause energy change.

Exothermic and Endothermic Processes

1. What is an exothermic reaction or process?

Any *reaction* or *process that gives off heat* to the surroundings is ***exothermic***. The value of ΔH for an exothermic reaction is always *negative* because H_2 has a lower value than H_1: *implying that the **reactants** have **more energy** than the **products***. The *temperature of the mixture* in which the exothermic reaction occurs *will be higher* than the initial temperature.

$$H_2 - H_1 = -\Delta H$$

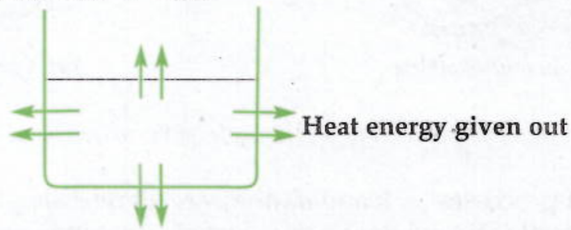


Figure 5.1: An exothermic reaction

In an exothermic change heat energy is released to the surroundings. The final temperature is higher than the initial temperature. Therefore, heat change is negative ($-\Delta H$).

2. What is endothermic reaction or process?

An *endothermic reaction* or *process is the one that absorbs heat from the surroundings*. In an endothermic process the values of ΔH are always positive because heat content in the products is more than the heat content in the reactants: i.e. H_2 is greater than H_1.

Since this process absorbs heat from the surroundings, the temperature of the mixture becomes lower than the initial temperature:

$$H_2 - H_1 = +\Delta H$$

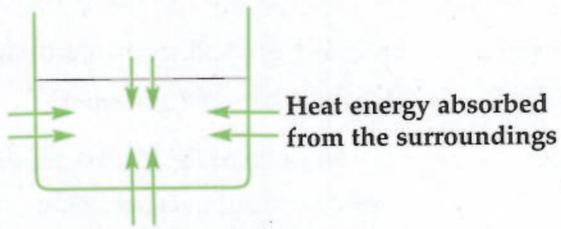
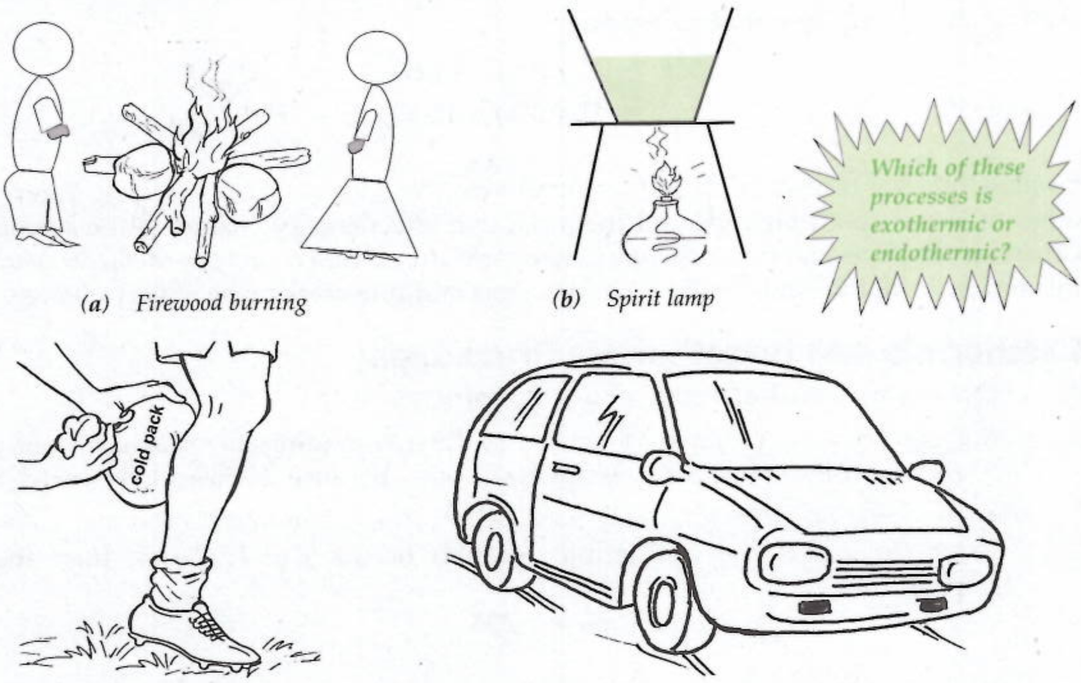


Figure 5.2: Endorthermic reaction

Examples of thermal changes in everyday life

There are so many activities that we do or see that are exothermic or endothermic. Study *Fig. 5.3* and classify each activity as endothermic or exothermic.



(a) Firewood burning

(b) Spirit lamp

(c) An injured athlete

(d) A car using petrol to run

Fig. 5.3: Examples of thermo reactions

When you look at processes (*a, b* and *d*), they are all releasing heat to their surroundings. This heat is keeping people around the fireplace warm, causing water in the beaker to boil, and the car to move. So these *processes are exothermic*.

However, in (*c*), where a cold compress is applied to the injured athlete's calf muscle, causes a freezing sensation that releases pain. A cold compress contains water and ammonium nitrate separated by a barrier. When the barrier is broken and ammonium nitrate mixes with water, then heat energy is absorbed from the water hence lowering the temperature of the compress. This *process is endothermic*.

Exercise 1

1. What is
   a) endothermic reaction?
   b) exothermic reaction?
2. Give any two examples of
   a) an endothermic process
   b) exothermic process

Heat changes during dissolution of substances in water

Heat changes occur when substances dissolve in water. Some substances *absorb* while others *give off heat*. The heat change that occurs when one mole of a substance completely dissolves in water to form a solution of a stated concentration is called ***heat of solution***.

In *Activity 1*, you will investigate heat changes during dissolution of substances in water.

Activity 1

Aim: *Investigating exothermic and endothermic processes in dissolution of substances in water.*

Materials:

- *ammonium nitrate or silver nitrate*
- *sodium hydroxide pellets*
- *2 beakers*
- *2 glass rods*
- *2 spoons/spatulas*
- *water*
- *thermometer*

Procedure:

1. *Put about 25 cm^3 of water in each beaker.*
2. *Measure the temperature of the water and record it.*
3. *Put 3 spatulaful of ammonium nitrate into one beaker.*
4. *Put 3 spatulaful of sodium hydroxide pellets into the second beaker.*
5. *Stir each mixture with a glass-rod.*
6. *Measure the temperature of each solution with a thermometer.*

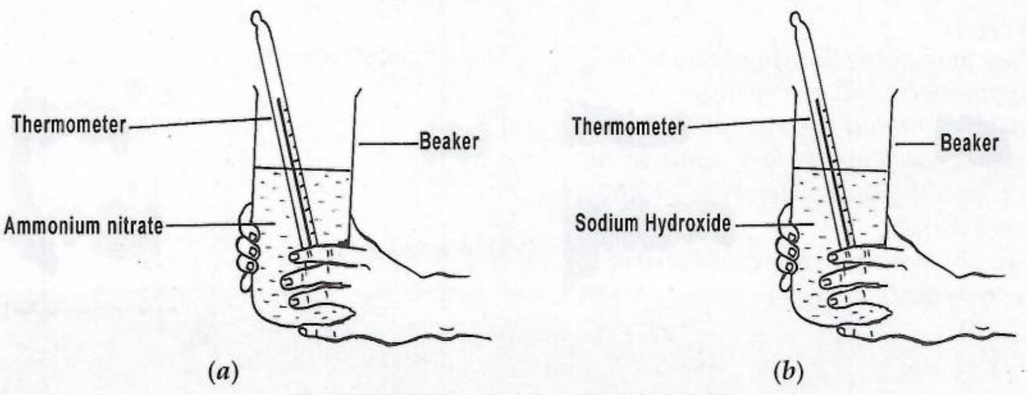


Fig. 5.4: Endothermic and exothermic reactions

7. *Put each beaker in your hand and observe how it feels.*
8. *Record any temperature change in each beaker.*

Results and observations:

	Amonium nitrate (NH_4NO_3)	Sodium hydroxide (NaOH)
Initial temperature		
Final temperature		
Temperature change		

Table 5.1: Temperature results

How did each beaker feel when you held it in your hands? In which one did the temperature rise or fall?

You may have observed that in the beaker containing a mixture of water and ammonium nitrate, the temperature went down. You also felt cold when you held this beaker in your hands. The chemical reaction that occurs between water and ammonium nitrate is *endothermic*: i.e. it absorbs energy from its surroundings including your hands.

You may have also observed that the temperature went up in the mixture of water and sodium hydroxide. The beaker containing the mixture felt warm in your hands. The reaction taking place in the beaker is *exothermic*: i.e. it releases heat energy to the surroundings including your hands.

Heat changes during neutralisation

Heat energy changes also occur *when an acid is neutralised by an alkali*. When 1M solution of an acid is neutralized by 1M solution of an alkali to form one mole of water, the heat change that takes place is called ***heat of neutralisation***.

In *Activity* 2, you will investigate heat energy changes in the neutralisation reaction.

Activity 2

Aim: *To investigate energy changes in a neutralisation reaction*

Materials:

- *2 beakers (150cm³ each)*
- *thermometer*
- *25cm³ 0.1M HCl*
- *cotton wool*
- *25cm³ 0.1M NaOH*

Procedure:

1. *Take and record the temperatures of hydrochloric acid and sodium hydroxide solutions separately.*
2. *Mix the acid and the base solutions in a 150cm³ beaker insulated with cotton wool outside.*
3. *Stir the mixture and record the final temperature.*

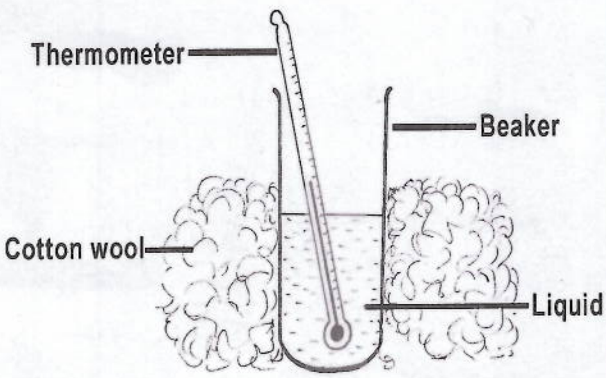


Fig. 5.5: Energy changes during neutralisation

Results and observations:

Solution	Temperature
Hydrochloric acid	
Sodium hydroxide	
After mixing (*final temperature*)	
Temperature change	

Table 5.2

Is the final temperature higher or lower?

You must have found out that the neutralisation reaction between an acid and a base also releases heat energy into the surroundings. So this reaction is *exothermic*.

The surroundings that exchange heat energy with the chemical reaction include *the beaker, hand holding the beaker, thermometer* and *air around it*.

Heat energy changes during changes of states of substances

When you heat ice in a beaker, *it melts into liquid*. If you continue heating, all the ice will change to liquid, then it will start boiling; consequently *change to vapour in the process*.

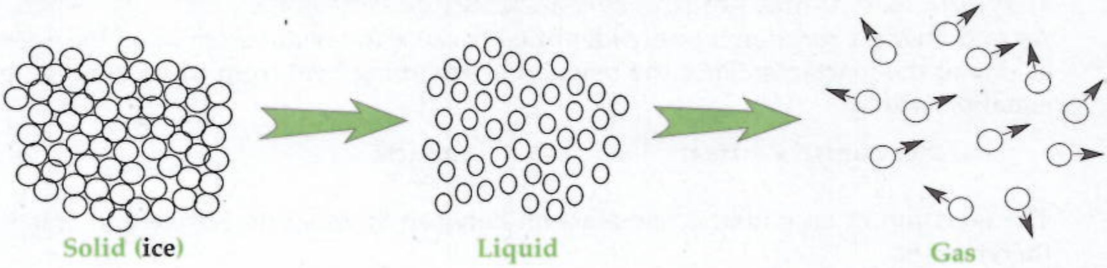


Fig.5.6: Ice melting to liquid and liquid evaporating

When any substance melts or boils, energy is absorbed by the particles. This energy breaks the forces that hold the particles together. So *changes in states from solid to liquid and from liquid to gas are endothermic*.

But *when a gas changes to liquid and liquid to a solid, heat is given off to the surroundings.* These changes are, therefore, exothermic.

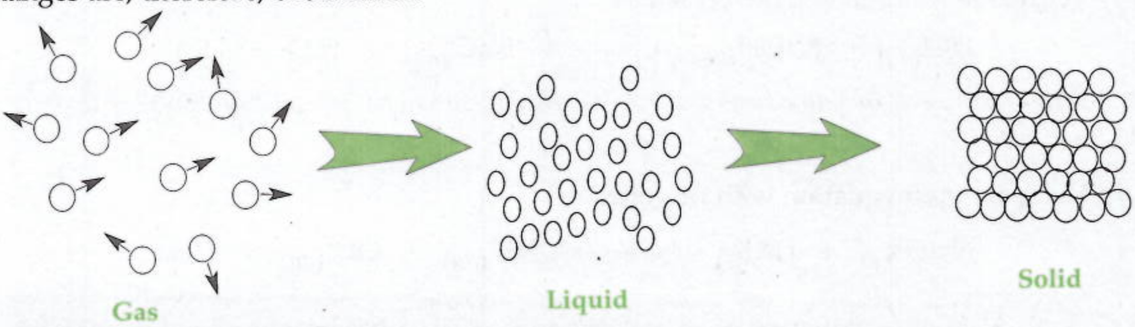


Fig 5.7: Gas condenses to liquid and liquid solidifies

Exercise 5.2

Complete the sentences by filling in the gaps:

1. A beaker of endothermic reaction feels ____________ because it __________ heat energy ______________ the ______________
2. A beaker of exothermic reaction feels ___________ because it _____________ heat energy ____________ the ______________

Energy level diagrams

Thermo-chemical reactions are *processes that release or absorb heat to or from their surroundings thereby affecting temperature*. These processes can be identified using their equations and graphs.

Identifying endothermic and exothermic reactions from their thermo-chemical equations

Chemical equations of thermo-reactions can tell us whether the reaction is endothermic or exothermic. They are presented with heat energy either as one of the reactants or one of the products.

1. Thermo-chemical equations of endothermic reactions

An endothermic reaction is easily identified because its chemical equation includes heat as one of the reactants. Since the reaction is absorbing heat from the surroundings, its equation will be:

Reactants + Heat ⟶ Products

The equation of an endothermic reaction between ammonium nitrate and water will, therefore, be:

$$NH_4NO_{3(s)} + H_2O_{(l)} + \text{Heat} \longrightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$$

2. Thermo-chemical equations of exothermic reactions

An exothermic process releases heat to the surroundings. Its equation will, therefore, include heat as one of the products:

Reactants ⟶ Products + Heat

The equation of the neutralisation reaction between hydrochloric acid and sodium hydroxide solution will be written as:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O + \text{Heat}$$

What would be the equation of a chemical reaction between sodium hydroxide pellets and water?

Compare your equation with this one:

$$NaOH_{(s)} + H_2O_{(l)} \longrightarrow Na^+{}_{(aq)} + OH^-{}_{(aq)} + \text{Heat}$$

Note: A chemical equation of an endothermic reaction has heat as one of its reactants while that of exothermic has heat as one of its products.

Graphs of thermo-reactions

We can use graphs to explain energy changes that accompany chemical reactions or physical processes. These graphs are called *energy level diagrams*.

1. Energy level diagrams for endothermic reaction

If a person lifts a brick from the ground to the top of a wall there is change of its energy level:

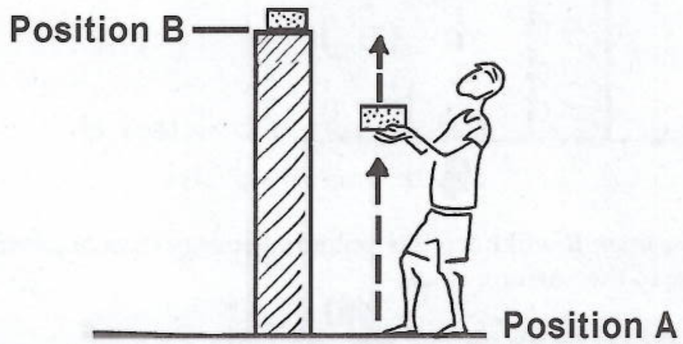


Fig. 5.8– A man lifting up a brick

As the brick is being lifted from *position* **A** to *position* **B**; it gains potential energy of height. If we take the brick at *position* **A** as a *reactant* and at *position* **B** as the *product*, we can say that ***the reactant has absorbed energy** from the surroundings* ***to become a product*** at *position* **B**. So **at** *position* **B the product has more energy than the reactant at** *position* **A.** The energy level diagram of an endothermic reaction will then be like *Figure 5.9*:

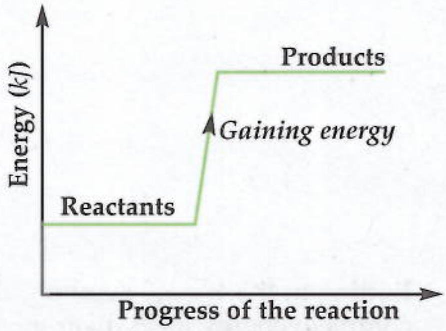


Fig. 5.9: Energy level diagram of an endothermic reaction

Now using the symbols H_1, H_2, ΔH, the complete energy level diagram of an endothermic reaction will now look like this:

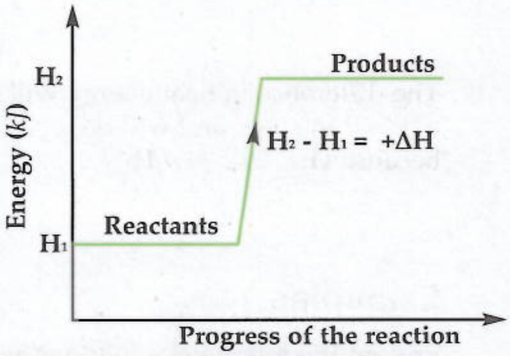


Fig. 5.10: Energy level diagram of an endothermic reaction

The *heat energy* (ΔH) *difference* of an endothermic reaction is positive because H_2 is greater than H_1.

2. Energy level diagrams for exothermic reaction

The picture below shows a brick being lowered from the top of a wall to the ground.

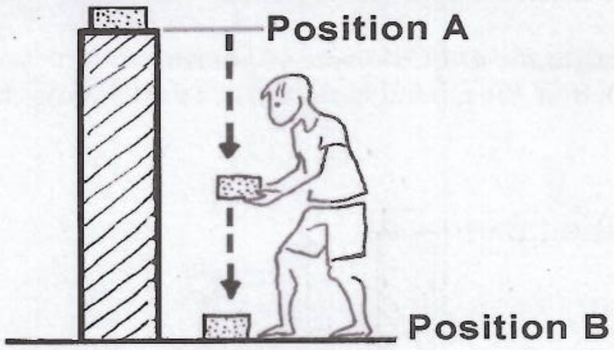


Fig. 5.11: A man lowering a brick

The brick at *position* B will have less potential energy than at *position* A. This brick has lost its energy to the surroundings.

If we take *position* A as the *reactant* and *position* B as the *product*, then the energy level diagram of an exothermic reaction will be like *Fig 5.12*:

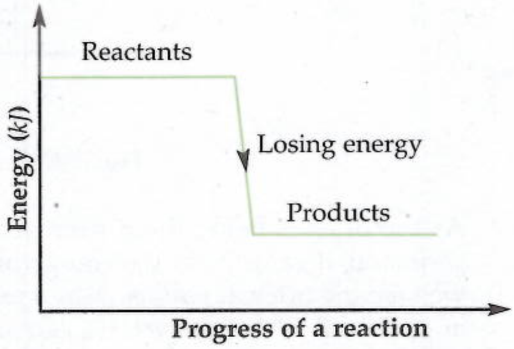


Fig. 5.12: Reaction process

Putting in the energy symbols, the complete energy level diagram of a exothermic reaction will be like in *Fig. 5.13*:

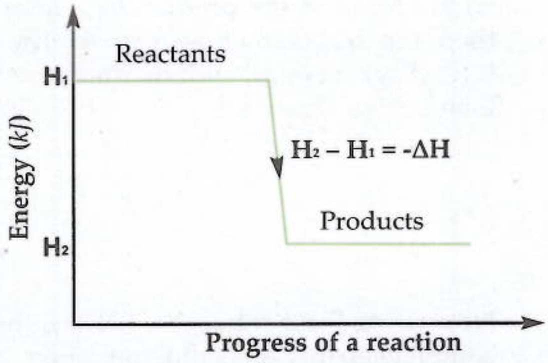


Fig. 5.13.: Energy level diagram

The difference in heat energy will be $\Delta H = H_2 - H_1$ and will be *negative* because H_2 *is less than* H_1.

Examples

Present the following equations as energy level diagrams for the thermo-reactions.

1. X + 1200 kj ⟶ Y

Solution

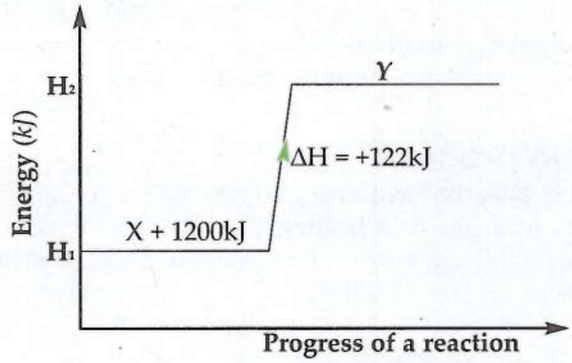


Fig. 5.14

This is an endothermic reaction. The energy (1200 Kj) from the surroundings is one of the reactants. So **Y** has more energy than **X**.

2. **A + B ⟶ AB + 350 Kj**

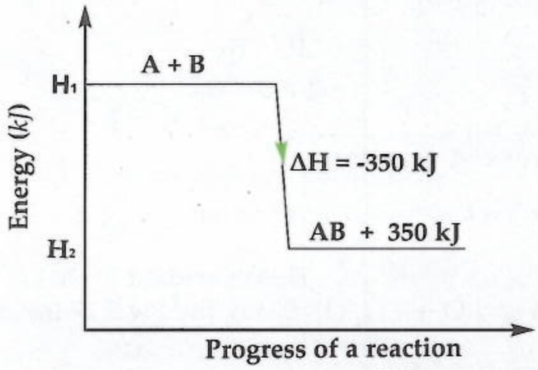


Figure 5.15

This is an exothermic reaction. The heat energy (350 Kj) is given off as one of the products. So the product (**AB**) has less energy than the reactants (**A** and **B**).

Exercise 3

1. Tell whether each of the reactions below is exothermic or endothermic
   a) M + 300 kj ⟶ N
   b) P + Q ⟶ PQ + 500 kj
   c) AB + 600kj ⟶ A + B
   d) U + V ⟶ UV + 3000kj
2. Draw energy level diagrams for each of the reactions in question 1.

Energy changes in bond breaking and bond formation

In every chemical reaction *bonds break* or *form*. The breaking or formation of chemical bonds in a reaction involves energy.

Bond energy

Bond energy could be defined as *the amount of energy in **kilojoules** (kj) associated with breaking or forming of 1 mole of chemical bonds in an element or compound*. Some bonds are stronger than others. Strong bonds require more heat energy to break. The *amount of energy required to break a bond is the same as the amount of energy given off* when the same is formed.

Bond breaking - endothermic process

When a solid melts to form a liquid, it absorbs heat energy from its surroundings. When a liquid changes to a gas at a boiling point, it absorbs more heat energy. Likewise, *when a bond breaks to form individual atoms, it absorbs heat energy from the surroundings*.

What form of process is bond breaking?

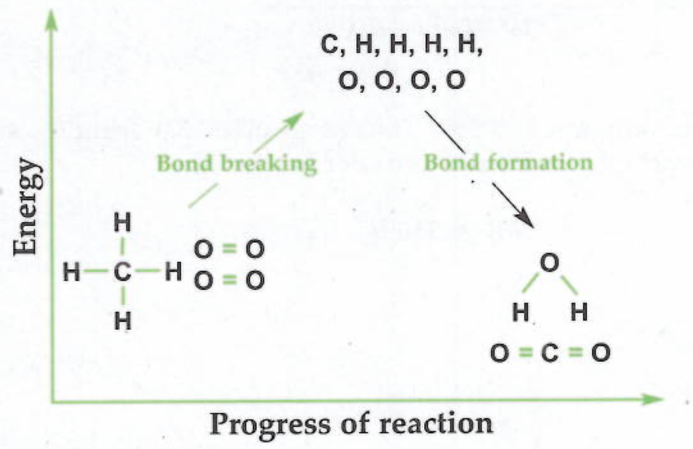


Fig 5.16: Illustrating bond breaking and formation

In Figure 5.16, it is clear that *bond breaking absorbs energy*. This is evident in that (CH_4 and $2O_2$) are at a lower level than (**C, H, H, H, H** and **O, O, O, O**). Study the *Fig.5.17* (*a*) and find out how comparative it is to *Fig 5.17* (*b*).

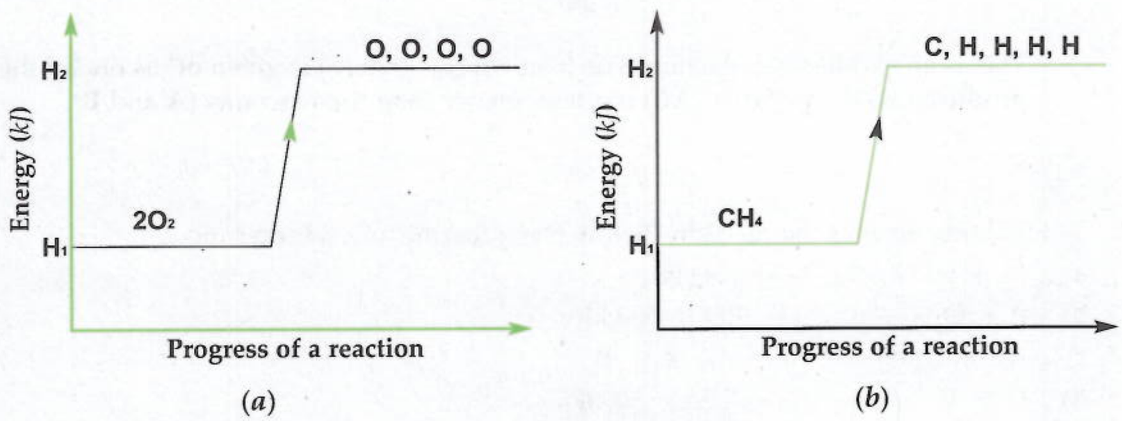


Fig.5.17: Breaking of chemical bonds

In both graphs the breaking of chemical bonds is absorbing energy from the surrounding. In (*a*) O = O absorbs energy to produce O and O and in (*b*) CH_4 absorbs energy to produce C, H, H, H, H.

Bond formation (exothermic process)

When water vapour is cooled, particles join together to form a liquid. When liquid water is cooled further it freezes to form ice. Both *condensation and freezing processes are exothermic.*

*When atoms join together **to form a bond**, energy is also given off*. So *bond formation is an exothermic process* like condensation and freezing, as it gives off heat to the surroundings.

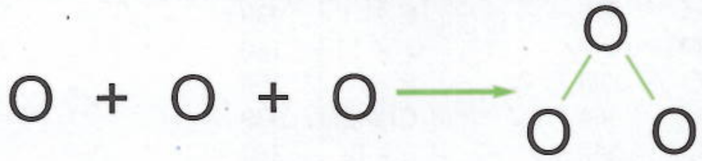


Fig 5.18: Atoms joining to form bonds – an exothermic process

Now study the *Figure 5.19*:

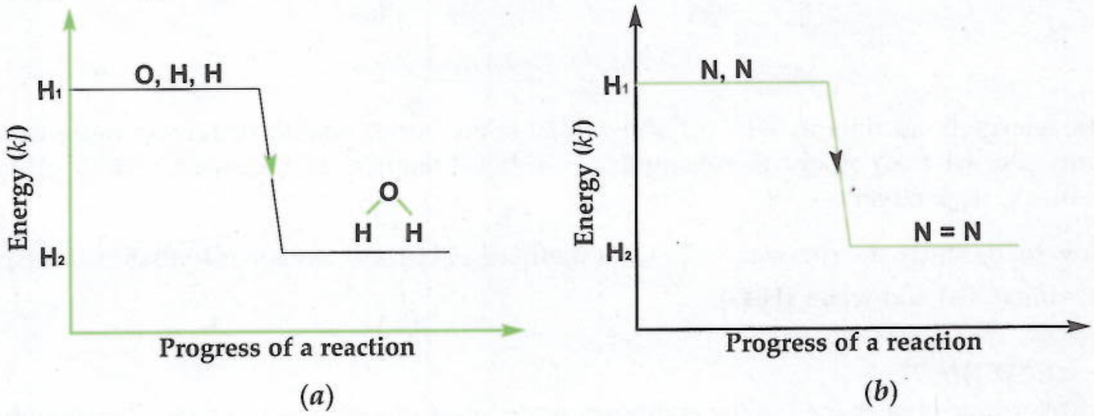


Fig 5.19: Formation of chemical bonds

In *Fig. 5.19* (*a*) oxygen and hydrogen atoms combine to form water. Energy is lost to the surroundings. In *Fig. 5.19* (*b*) nitrogen atoms are combine to form nitrogen molecules. Even here energy is lost into the surroundings. So in both reactions, bonds are formed and energy is given off.

Exercise 4

1. What is bond energy?
2. Why is change of state from solid to liquid endothermic?
3. In the following reactions:

a) $NH_3 \longrightarrow N + 3H$
b) $O_2 + 2H_2 \longrightarrow 2H_2O$
c) $SO_2 \longrightarrow S + O + O$

identify which is exothermic or endothermic.

Determining the kind of reaction using bond energy

You can determine whether a reaction is exothermic or endothermic from the value of the overall energy change. The *overall energy change* is worked out by *subtracting energy given out when bonds form from the energy absorbed when bonds break*.

Calculating the overall energy change using bond energy

Since exothermic change gives out energy while endothermic change takes in energy, the overall energy change can easily be worked out using the data of some of the bond energies below:

Bond	Bond energy ($Kjmol^{-1}$)	Bond	Bond energy ($Kjmol^{-1}$)
C – H	435	C - Cl	339
O = O	497	H - H	160
C = O	803	F - F	156
H – O	464	Cl – Cl	339
C – C	347	C – Br	280
C – O	358	Br – Br	193
C = C	610	O – O	166
C ≡ C	835	N – N	390
C – F	495	H - Br	365

Table 5.1: The bond energies

The energy bond data in *Table 5.1* shows that some bonds absorb or release more energy to form or break than others. For example, C = O bond requires and releases $803 kj/mol^{-1}$ to form or break, respectively.

Now let us study the chemical *reaction* of methane (CH_4) and oxygen (O_2) that *produces* carbon dioxide (CO_2) and water (H_2O).

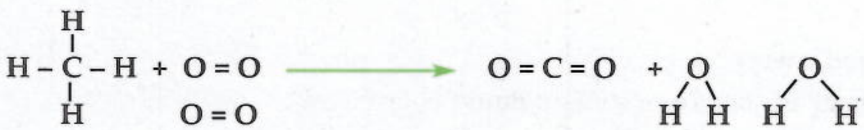
Example 1

Determine heat change in the combustion of 1 mole of methane to produce carbon dioxide and water. Is the reaction endothermic or exothermic?

Solution

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

If we show the bonds in each molecule, we will have the following equation:



In the equation above, there are 4 C – H and 2O = O bonds to break according to the given reactants. There are also 2C = O and 4 O – H bonds to form according to the products formed. Energy required (*absorbed*) for breaking bonds will be as follows:

4C – H = 4 x 435 = 1,740 Kj

2O = O = 2 x 497 = <u>994 Kj</u>

Total energy required = **2734 kj**

According to the products, the following are the bonds to form: 2 C = O and 4 O - H. These bonds will release the following energy as they form:

2C = O 2 x 803 kJ = 1606 kJ

4O – H = 4 x 464 kJ = <u>1856 kJ</u>

Total energy released = **3462 kJ**

The ***energy difference*** (ΔH) will be as follows:

Energy required to break bonds* – *Energy given off to form bonds

= 2734 – 3462

= **-728 kJ**

The enthalpy change is negative

This chemical reaction is exothermic. Therefore, *the overall energy released to the surroundings is* **728 KJ**.

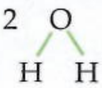
Example 2

Determine whether the following reaction is exothermic or endothermic.

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

Solution

Number of the bonds of the reactants to break:

2  = 4 O – H bonds

The amount of energy to be absorbed by the reactants:

O – H requires 464 kJ

Therefore,

4 O – H requires 4 x 464 kJ

= **1856 kJ**

Number of bonds to be formed by the products:

O = O + 2 H – H

Amount of energy to be given out:

O = O gives off 497 kJ

2 H – H gives off 2 x 160 = 320 kJ

Total energy given off = **817 kJ**

Energy difference (ΔH)

Energy absorbed* – *Energy given off

1856 Kj – 817 Kj

= **+1039 KJ**

The enthalpy change is positive

The chemical reaction is endothermic. The *energy absorbed is more than the energy given out*.

Exercise 5.5

Use the bond energy data in *Table 5.1* to answer the following questions:

1. Work out the heat energy for the following reactions:
   (a) $H_2O \longrightarrow H_2 + O$
   (b) $CO_2 \longrightarrow C + O_2$
   (c) $6CO_2 + 6H_2O \longrightarrow C_6H_2O_6 + 6O_2$
2. In each of the reactions in *question 1* state whether it is endothermic or exothermic.

Unit summary

- An *exothermic reaction releases heat to the surroundings*. The temperature reading of the thermometer rises and the surroundings become warmer when it is taking place. *Combustion of fuels* are good examples of exothermic reactions.
- An *endothermic reaction absorbs heat from the surroundings*. The surroundings, therefore, become cooler. Examples of endothermic reactions are:
  - ***a.*** *reactions of ammonium* or *potassium nitrate with water*.
  - ***b.*** *reactions of the sodium hydroxide pellets with water*
- *Heat energy change* of endothermic reaction *is positive*, while for *exothermic reaction is negative*.
- Graphs of thermo-chemical reactions are called energy *level diagrams*.
- *Bond breaking absorbs energy* from the surroundings while *bond formation releases energy*.
- Bond energy is defined as *the amount of energy in kJ associated with breaking* or *forming of 1 mole of chemical bonds* in a molecular element or molecular compound.
- The overall heat energy change is the *heat of reaction*.

Unit exercise

1. Which of the following reactions are exothermic?
   a) Steam condensing on your skin
   b). Burning of carbon
   c) Dissolving ammonium nitrate in water
   d) Dissolving salt in water

2. *Figure 5.20* shows the graph obtained when calcium oxide is added to dilute sulphuric acid.

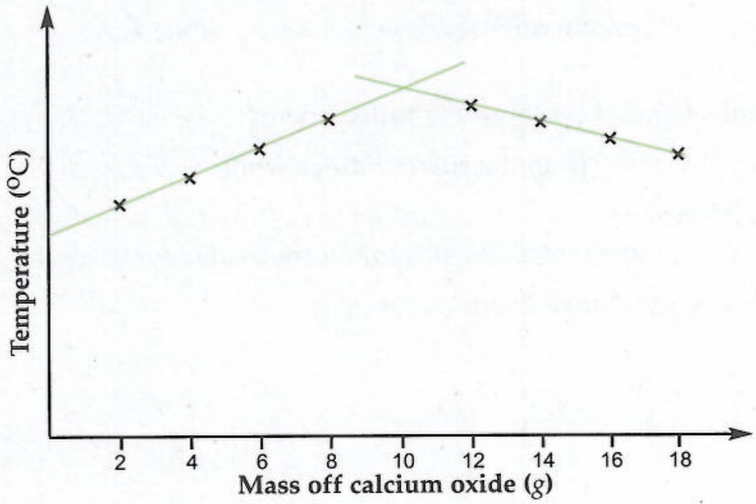
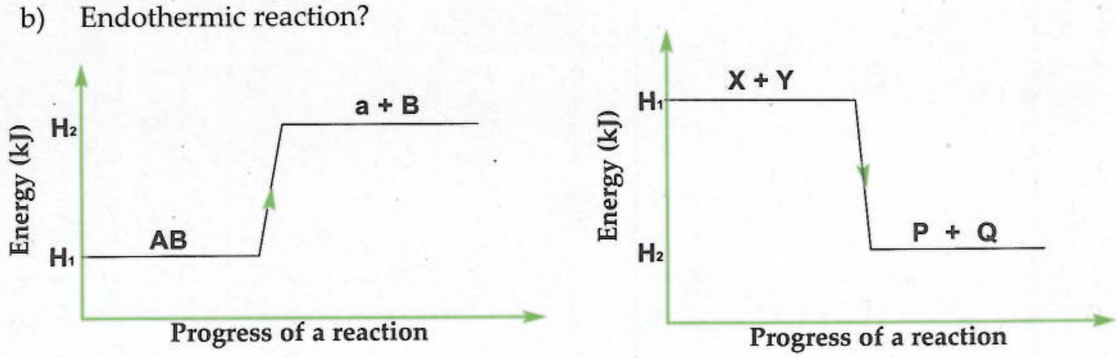


Fig 5.20: Mass of calcium oxide added

   a) Is the reaction exothermic or endothermic?
   b) What mass of calcium oxide is added to use up all the acid?
   c) Why does the temperature go down after the mass of calcium oxide has been added?

3. Using the bond energy data in *Table 5.1*:
   a) Calculate the enthalpy of combustion of ethanol (CH_3CH_2OH)
   b) Draw the energy level diagram to represent the energy process.

4. Which of the following graphs represent:
   a) Exothermic reaction?
   b) Endothermic reaction?



(*a*) *Fig 5.21: The graphs of heat changes* (*b*)

5. Hydrogen gas (H_2) and chlorine (Cl_2) gas react to form hydrogen chloride (HCl) according to the equation:

$$H_2 + Cl_2 \longrightarrow 2HCl$$

If the total energy required to break H – H and Cl – Cl bonds is 678kj and total energy released is 862 kj.

a) Is the reaction exothermic or endothermic?

b) work out the value of **ΔH** of the heat energy change.

Unit 6

Alkanols

Success Criteria

By the end of this unit you should be able to:

1. Identify the functional group of alkanols
2. Draw the structure of the first ten unbranched primary alkanols
3. Name the first ten unbranched alkanols using IUPAC rules
4. Write the molecular formulae of alkanols given the number of carbon atoms
5. Write the condensed formulae of the first ten alkanols
6. Deduce the general formulae of alkanols
7. Write and name the branched chain alkanols
8. Classify alkanols as primary, secondary and tertiary
9. Describe indigenous and improved ways of preparing alkanols
10. Explain physical properties of alkanols such as polarity, solubility, melting point, density, volatility and viscosity.
11. Describe chemical properties of alkanols such as reactivity with metals, dehydration and oxidation
12. Explain uses of ethanol

Key words:

In this unit you will find these key terms and concepts:

Alkanol, hydroxyl group, ethanol, fermentation, IUPAC, nomenclature, primary, secondary and tertiary alcohols, oxidation, esterifcation

Ensure that you understand and learn how to apply them both for your academic and real life situations.

Organic chemistry is *the study of compounds containing carbon*. Carbon atoms have very unique characteristics: i.e. *can join up with other carbon atoms to form various sizes of molecular chains*. These chains may be straight, branched or ringed. At Junior Secondary level, you saw that compounds that contain carbon atoms are called *organic compounds*. You also learnt that there are families of compounds called *homologous series*: i.e. members with similar properties. Examples of homologous series are **alkanes** and **alkenes**. In this course you will study more homologous series of organic compounds, namely: **alkanols**, **alkanoic acids**, **alkanals**, **alkanones** and **alkanoates**. In this chapter, you will focus on **alkanols** (also called **alcohols**).

Functional group

Alkanols are a *homologous series of organic compounds composed of a chain of carbon atoms bonded to a group of hydrogen atoms*, but with **–OH** present in the chain. The **–OH** is called the **hydroxyl group**. It acts as the **functional group** in the compound. The compound in *Figure 6.1* is an example of an alkanol.

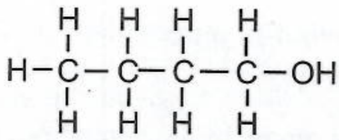


Figure 6.1 An example of alkanol

Nomenclature

If you may recal your Junior Secondary course, you will remember that you name alkanols using the **International Union of Pure and Applied Chemistry (IUPAC) system**. Just as alkanes, *alkanols are named following the number of carbon atoms contained in the molecule*. We shall use similar prefixes as in alkanes to name the alkanols except we must add **the suffix *–anol*** to *denote the presence of the functional group* **–OH**. *Activity 1* will help you recall the prefixes and learn how to name the first ten alkanols.

How can you name alkanols?

Which alkanes can you remember?

Activity 1

Naming alkanols

Complete *Table 6.1* by adding the prefixes depending on the number of carbon atoms and then name each compound with the suffix *–anol*. The first one has been done for you.

Number of carbon atoms	Prefix	Suffix	Name of compound
1	Meth-	*–anol*	*Methanol*
2		*–anol*	
3		*–anol*	
4		*–anol*	
5		*–anol*	
6		*–anol*	
7		*–anol*	
8		*–anol*	
9		*–anol*	
10		*–anol*	

Table 6.1: Naming Alkanols

Note: The name of the alkanol includes a *prefix* and the *suffix –anol*. For example, if the number of carbon atoms is *five*, the *prefix is pent-*, when we add this to the *suffix –anol*, the name of the compound is **pentanol**.

Structure and molecular formulae of Alkanols

You will recall that *carbon has a valency of 4*, hence it forms four bonds (*Figure 6.2*) with other atoms such as hydrogen, oxygen, the halogens, etc.

Figure 6.2: Bonds of a carbon atom

In addition, we saw that *carbon atoms can bond to each other in a continuous chain forming a skeleton* to which other atoms such as hydrogen, oxygen and the halogens can attach (*Fig. 6.3*).

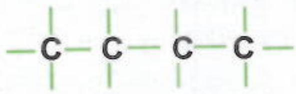


Figure 6.3: Chain of carbon atoms

What is a structural formula? What does it represent

The resulting molecule or compound is represented using a *structural* or *molecular formula*. The structural formula *shows the manner in which the atoms bond together which gives the shape of the molecule*. If you represent the proportions in which the different atoms in the compound bond up, then a *molecular formula* is formed. Take, for instance, the compound *ethanol* in which *two carbon atoms are bonded to five hydrogen atoms* and *a hydroxyl group*; the structure of the molecule would be as in *Figure 6.4*:

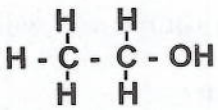


Figure 6.4 A structure of ethanol

If you write the formula by summing up the atoms around each carbon, then you will get a shortened structure of the compound to **CH_3CH_2OH**. This is known as the *condensed formula*.

What is the molecular formula of the compound above? If you counted all the carbon and hydrogen atoms bonded together, you will get the formula **C_2H_5OH**.

Activity 2

Writing structural, condensed and molecular formulae of the first ten unbranched alkanols

Complete *Table 6.2* by adding the structural, condensed and molecular formulae of the compounds named.

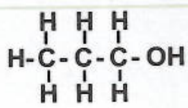
No. of Carbon atoms	Name	Structural formula	Condensed formula	Molecular fomula
1	Methanol			
2	Ethanol			
3	Propanol		$CH_3CH_2CH_2OH$	C_3H_7OH
4	Butanol			
5	Pentanol			

Table 6.2

No. of Carbon atoms	Name	Structural formula	Condensed formula	Molecular fomula
6	Hexanol			
7	Heptanol			
8	Octanol			
9	Nonanol			
10	Decanol			

Table 6.2

Activity 3:

Aim: *Making the structures of ten unbranched alkanols*

Materials: • *pith balls or clay* • *sticks*

Procedure:

1. *Pick three different sizes of pith balls or construct clay balls which have right number of holes to represent valencies of each type of atom.*
2. *Colour the balls black to represent carbon, white to represent oxygen and blue to represent hydrogen atoms, respectively.*
3. *Join up the balls using sticks to form structures of the various alkanols, starting from methanol to decanol.*

Figure 6.5 shows some of the shapes you may have formed:

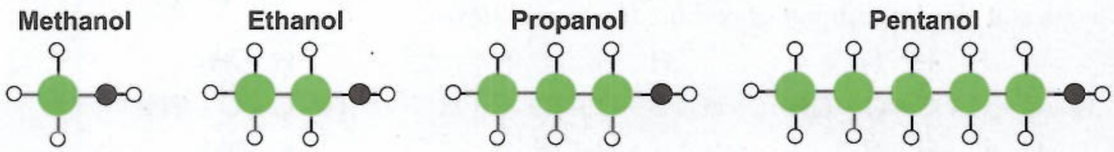


Figure 6.5: Molecular structures of some alkanols

General Formula for Alkanols

You may have noticed that there is a pattern formed in the structure of the molecules as you move from one to the next. Look closely at the condensed formulae of the alkanols in *Activity* 2. You will observe that there is always a difference of $-CH_2$ between successive molecules.

With reference to Activity 2, what can you say about the condensed formulae of alkanols?

You may have noticed that *the number of hydrogen atoms is twice the number of carbon atoms but with one extra hydrogen atom* and the –OH group. If we represented the number of carbon atoms by n, then the number of hydrogen atoms is $2n+1$, the whole relationship may then be presented as $C_n H_{2n+1} OH$. This is the *general formula* for alkanols.

What is the relationship between carbon and hydrogen atoms?

Knowledge of this pattern helps us come up with the formula of any alkanols provided you know the number of carbon atoms as you did with alkanes and alkenes. For example, an alkanol with 6 carbon atoms, has a molecular formula: $C_6H_{2(6)+1}OH$, which comes to $C_6H_{13}OH$.

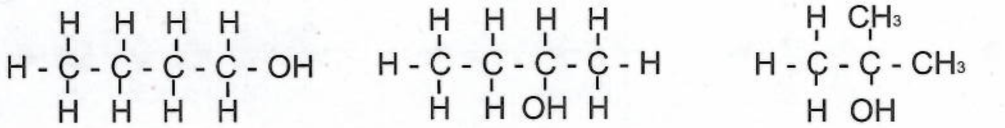
Exercise 1

1. Name the functional group of alkanols.
2. What is the name of the alkanols with
   a) 3 carbon atoms
   b) 7 carbon atoms?
3. Draw the structural formulae of alkanols with
   a) 4 carbon atoms
   b) 7 carbon atoms and write down its molecular formula.
4. Use the general formula of alkanols to determine the molecular formula of an alkanol with
   a) 9 carbon atoms
   b) 15 carbon atoms?

Branching in alkanols

Recall what you studied on the family of alkanes, the alkanol carbon chain may also form branches that enables it to form various isomers of alkanols. Remember that *isomers are molecules which have the same molecular formulae but are different in structure*. Study *Figure 6.6* in which molecules of similar number of carbon atoms are shown.

What is the molecular formula of each compound in Fig. 6.6? Do they have the same name? If not, name them.



A B C

Figure 6.6 Structures of butanol

Note: The *name of an alkanol also shows the position of the hydroxyl group* on the carbon chain.

You may have given the three alkanols systematic names as you learnt with alkanes and alkenes. The *name must show the number of carbon atoms* and *the position of the –OH group*. The following are the names you could have given and their reasons:

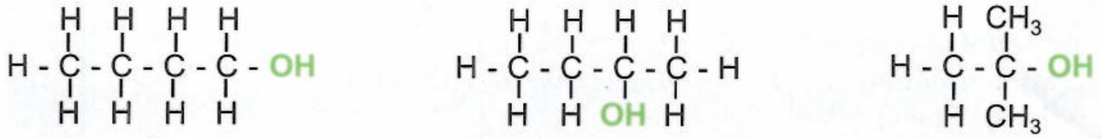
A is *butan-1-ol* since the molecule has 4 carbon atoms and the –OH group is on carbon 1

B is *butan-2-ol* since the molecule has 4 carbon atoms and the –OH group is on carbon 2

C is *2-methylpropan-2-ol* since the molecule has 3 carbon atoms in the longest chain and the –OH group is on carbon 2.

Classification of alkanols

Chemists classify alkanols as *primary*, *secondary* and *tertiary alcohols* based on the *position* *of the hydroxyl group* and *branching of the carbon chain*. Study the structures of the compounds in *Figure 6.7*.



A B C

Figure 6.7: Alkanols classsified as primary, secondary and tertiary alcohols in A, B and C, respectively

You will notice that *the –OH in* **A** *is attached to a carbon which is also attached to two hydrogen atoms*. This is classified as a ***primary alcohol***. In **B** *the –OH is attached to a carbon which is bonded to a methyl group* and *one other hydrogen atom*. This is classified as a ***secondary alcohol***. In **C** *the –OH is bonded to the carbon which is not bonded to any other hydrogen atom* and is classified as ***tertiary alcohol***.

What do you notice about the carbon attached to -OH?

Exercise 2

Classify the following alkanols as primary, secondary or tertially:

1. $CH_3CHOHCH_2CH_3$
2. $CH_3CH_2C(OH)(CH_3){-}CH_3$
3. $OH\,CH_2CH_2CH_3$
4. $CH_3CH_2CH_2CH_2CH_2OH$
5. $CH_2CH_2CH_2C(OH)(CH_3){-}CH_2CH_3$

Preparation of alkanols

Alkanols can be prepared in various ways. The most commonly found and least toxic of alkanols is ***ethanol***. It is *the active substance in beers*. We will use it to demonstrate the preparation and use of these alkanols.

Ethanol *is a colourless, volatile* and *pleasantly smelling liquid that is soluble in water*. It ***cannot conduct electricity*** because it is a non-polar compound. It can be produced both locally and at industrial level. The methods of production include **fermentation of sugars** and **hydration of ethene**.

1. Fermentation of Sugars

It is *the decomposition of complex compounds into simpler compounds by the help of micro-organisms or enzymes in the absence of oxygen*. In the fermentation of sugars, such as **glucose** and **fructose**, yeast is used to provide enzymes which decompose the sugars into **ethanol** and **carbon dioxide**. The micro-organisms feed on the sugar ***anaerobically*** to obtain energy. The equation for the breakdown of the sugars is as follows:

$$C_6H_{12}O_6(g) \longrightarrow 2C_2H_5OH(g) + 2CO_2(g)$$

Glucose — Ethanol — carbon dioxide

To start the reaction, the sugars are thoroughly mixed with yeast in water. After fermentation the mixture is distilled in order to obtain pure ethanol. A good example of *locally made ethanol* is **kachasu**.

Explain how kachasu is prepared locally?

In the traditional setup one mixes sugar and bran with water and leaves it to ferment for 3-4 days to produce ethanol. *Yeast may be added to speed up the process*. After fermentation, *the mixture is distilled to obtain Kachasu* (ethanol). Distillation is done as in *Figure 6.8*.

Note: For the apparatus in *Figure 6.8* there is a pot, a water trough and a receiving vessel at the farther end. Describe the use of the trough and receiving vessel.

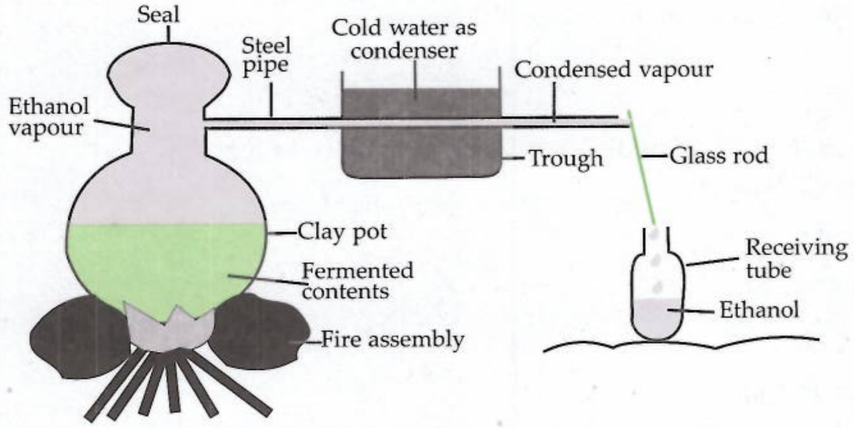


Figure 6.8: Distillation apparatus for locally made ethanol

The mixture of maize bran, sugar and yeast is heated to boiling point. The vapour is passed through a pipe, then a trough of cold water to condense ethanol vapour to liquid ethanol, in order to cool ethanol and collect it in the receiving vessel.

At what temperature is ethanol likely to be collected? Why?

Activity 4

Aim: *Preparation of ethanol*

Materials:

- *Maize bran*
- *condenser*
- *sugar, yeast*
- *collecting bottle*
- *earthen pot*
- *heat source*
- *steel pipe*

Procedure:

1. *Add some bran, sugar and yeast*
2. *Leave it to stand for three to four days; for the mixture to ferment*
3. *Seal the earth pot and connect it to the steel pipe, condenser and collecting bottle as shown is* Figure 6.8.
4. *Heat the mixture gently*

Results/Observations:

Describe what happens from heating the mixture in the pot to the receiving bottle.

2. Hydration of Ethene

Ethanol may also be prepared at industrial level through *hydration of ethene* (*an alkene*). This process is done in the presence of *heat* and *phosphoric acid* as a catalyst.

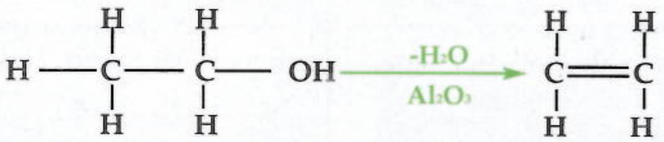
Hydration is *the addition of a water molecule into another molecule*. During hydration the water molecule adds across the *double bond of ethene*. The water itself first splits up into the *hydrogen ion* (H^+) and *hydroxyl ion* (–OH). The H^+ *adds on to one free bond of carbon* while the –OH *adds on to the free bond of another carbon across the double bond* as shown in the equation below:

What happens during the addition reaction of alkenes?

$$CH_2{=}CH_2\,(g) + H_2O(g) \xrightarrow[\text{Heat at 600°C}]{\text{Phosphoric acid}} \underset{\text{H}}{CH_2}{-}\underset{\text{OH}}{CH_2}\,(g)$$

Ethene Steam Ethanol

This reaction releases a lot of heat and it is reversible: i.e. ethanol can be converted back to ethene, as in a dehydration reaction below:



This reaction involves heating an alkanol and passing its vapour over heated aluminium oxide.

Physical Properties of Alkanols

Alkanols differ in size due to the difference in number of carbon atoms in the molecules. This affects their physical properties. ***As the number of carbon atoms increases, the force holding the particles together also increases.*** These forces are called *intermolecular forces*. Comparing the intermolecular forces of alkanols with those of alkanes and alkenes with the same number of carbon atoms, *alkanols have higher intermolecular forces* (IMF). The strength of the intermolecular force *affects melting and boiling points*, *solubility*, *viscosity* and *density* of alkanols.

Activity 5

Group discussion: *Effect of IMF on physical properties of alkanols*

1. *Define the meanings of the terms melting and boiling points, solubility, viscosity and density*
2. *Discuss what happens to the*
   a) *melting and boiling points*
   b) *solubility*
   c) *viscosity*
   d) *density of alkanols when the molecule of the alkanol gets large*

1. Solubility of alkanols

Do *Activity 6* for you to learn on what happens when you add an alkanol to water.

Activity 6

Aim: *Solubility of ethanols*

Materials: • *100ml beaker* • *water* • *ethanol*

Method:

1. *Pour 15 drops of water in a 100ml beaker*
2. *Add 10 drops of ethanol to the beaker. What do you observe?*
3. *Describe your observations when you do the same using methanol, propanol, hexanol and octanol.*

a. When you add alkanol to water, the mixture forms one layer as you may have observed in *Activity 6*. This shows that *alkanols are soluble in water*.

> **Note:** *The size of the molecule affects its properties. An increase in number of carbon atoms in an alkanol reduces its solubility. Therefore, bigger alkanols are less soluble in water than smaller ones.*

b. Solubility in water is used as a *special test for alkanols*. Alkanols are soluble in water because both alkanols and water contain the –OH group. This agrees with the fact that **like dissolves like**.

> **However**, *the carbon chain is not soluble in water. For smaller alkanol molecules, the –OH becomes more powerful than the carbon chain, hence they are more soluble in water than larger ones. So, methanol, ethanol and propanol are more soluble in water because of their short carbon chain. Butanol is moderately soluble in water.*

c. Alkanols containing five or more carbon atoms are ***insoluble in water*** due to high intermolecular forces caused by large sizes of carbon chain.

2. Melting and boiling points

The melting and boiling points of alkanols increase with increase in molecular size. For example, butanol has higher melting and boiling points than methanol. If you compare the melting and boiling points of alkanols with those of alkanes of the same number of carbon atoms, *alkanols will have higher melting and boiling points than alkanes* because in alkanols there are more atoms than in alkane with similar number of carbon atoms. The increase in the number of atoms increases the IMF, making it hard to separate the particles.

3. Viscocity

Viscocity measures the ability of a liquid to flow. Viscous liquids flow less easily. Alkanols are more viscous than water. The *viscocity of alkanols increases with increase in number of carbon atoms*. The viscosity of alkanols also increases with increase in numbers of –OH groups present in a molecule. *The more the number of –OH groups present in a molecule, the more viscous the liquid*. For example, glycerine is made from glycerol with 3-OH groups. This makes it more viscous than other alkanols.

How would you compare the viscousity of ethanol with octanol?

Activity 7

Aim: *Compare viscousity of smaller alkanols with larger alkanols*

Materials: • *100ml beaker* • *water* • *ethanol* • *octanol*

Method:

1. *Pour 15 drops of water in a 100ml beaker*
2. *Add 10 drops of ethanol to the beaker. What do you observe?*
3. *Describe your observations when you do the same using methanol, propanol, hexanol and octanol.*

4. Density

Density is *mass per unit volume*. The density of alkanols increase as the number of carbon atoms increase. For example, butanol is denser than methanol because it has more carbon atoms.

Chemical Properties of Alkanols

At Junior Secondary level you learnt that *the functional group determines the chemical properties of an organic compound*. In alkanols, chemical properties are influenced by the presence of the functional group –OH. The proportion of the functional group to the size of the molecule affects the rate at which alkanols react with other substances. *The proportion is large when the number of carbon atoms is small, but as the carbon atoms increase, the proportion decreases*. This means that ***smaller alkanols react faster*** with other substances than larger alkanols. We shall study several chemical reactions that alkanols undergo.

1. Combustion or burning

To begin with be reminded that ***combustion is the reaction of a substance with oxygen***. Alkanols burn in oxygen with a blue flame. *When the combustion is **complete** they burn with a **blue flame*** and ***carbon dioxide** and **water** are produced*. However, *when the combustion is **incomplete** they burn with a **yellow flame*** and ***carbon monoxide** and **water** are produced*.

$$CH_3CH_2OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

Smaller alkanols like ethanol undergo complete combustion and larger such as octanol undergo incomplete combustion.

Why do larger alkanols give incomplete combustion?

Remember that, an increase in the number of carbon atoms reduces the percentage of OH group relative to the carbon atom part in the molecule.

2. Reaction with sodium

Alkanols react with sodium metal and this produces hydrogen gas. They react with sodium in a similar way to water, because both water and alkanol molecules have the –OH group of atoms. With water, the products are sodium hydroxide and hydrogen. While with alkanols, such as ethanol, the products are sodium ethoxide and hydrogen:

$$2Na_{(s)} + 2CH_3CH_2OH \longrightarrow 2CH_3CH_2ONa + H_2(g)$$

The *reaction of alkanols with sodium is much* **slower** *than with water*. In the reaction with sodium, *only the hydrogen atom joined to oxygen* **is involved**. Hydrogen atoms linked to carbon are *unreactive* as they did in alkanes. The safe way of getting rid of waste sodium in the laboratory is to cut it into small pieces, then add it to ethanol, a little at a time.

Similar reactions may take place with other alkali metals, such as potassium and lithium. The following equations of reactions gives:

$$\underset{\text{Propanol}}{2CH_3CH_2CH_2OH} + \underset{\text{Lithium}}{2Li_{(s)}} \longrightarrow \underset{\text{Lithium propanoxide}}{2CH_3CH_2CH_2OLi} + \underset{\text{Hydrogen gas}}{H_{2(g)}}$$

$$\underset{\text{Ethanol}}{2CH_3CH_2OH} + \underset{\text{Potassium}}{2K_{(s)}} \longrightarrow \underset{\text{Potassium ethoxide}}{2CH_3CH_2OK} + \underset{\text{Hydrogen gas}}{H_{2(g)}}$$

Activity 7

Aim: *Investigating the ease of reaction of alkanols*

Materials:

- *Methanol*
- *ethanol*
- *propanol*
- *octanol*
- *hexanol*
- *deflagrating spoon*
- *burner*
- *sodium metal*
- *beaker*

Procedure:

1. *Heat methanol on a deflagrating spoon*
2. *Observe the ease with which it burns and the flame colour*
3. *Repeat with ethanol, propanol, hexanol and octanol*
4. *Record your observations in a table*
5. *Now put 50ml of each compound in a separate beaker*
6. *Cut 5 equal portions of sodium metal, then add it to each beaker*
7. *Observe what happens and record in a table*

Results:

1. Arrange the compounds in order of reactivity.
2. Write a laboratory report on the activity you have just done.

3. Oxidation (addition of oxygen to a molecule)

Oxidation of alkanols *takes place in the presence of powerful oxidizing agents* such as warm acidified *potassium (IV) dichromate* ($K_2Cr_2O_7$) or *potassium permanganate* ($KMnO_4$). The result is the production of **alkanoic** (*carboxylic*) **acids**. For instance, the oxidation of ethanol with $K_2Cr_2O_7$ will form **ethanoic** (*acetic*) **acid**. Potassium (VI) dichromate *is orange in colour*, but during the oxidation process of ethanol, it changes to a **dark green**. The colour change indicates that a chemical reaction has taken place.

$$CH_3CH_2OH \xrightarrow[H^+]{Cr_2O_7^{2-}} CH_3COOH$$

Oxidation of ethanol takes place in two stages. In the *first stage*, it reacts with the H^+ to form *water* and **ethanal** (an adelhyde). *Second, the ethanal is* **oxidized** *to ethanoic acid.*

1st Step: $CH_3CH_2OH \xrightarrow[H^+]{K_2Cr_2O_7} \underset{\text{ethanal}}{CH_3CHO} + H_2O$

2nd Step: $CH_3CHO \xrightarrow[H^+]{K_2Cr_2O_7} CH_3COOH + H_2O$

ethanoic acid

Note: *The oxygen used in the reaction to convert the adelhyde to an acid comes from the oxidizing agent.*

Oxidation of ethanol may also occur in the presence of *certain micro-organisms* which catalyse the process.

$$C_2H_5OH \xrightarrow{\text{micro-organisms}} CH_3COOH + H_2O$$

ethanoic acid

4. Esterification

Alkanols react with carboxylic acids to form *esters*. This reaction is called *esterification*. Esterification takes place in the presence of *concentrated sulphuric acid* (H_2SO_4). For example, propanol and ethanoic acid can react in the presence of concentrated sulphuric acid to form *propyl-ethanoate* (*an ester*).

$$C_3H_7OH + CH_3COOH \xrightarrow{\text{conc. } H_2SO_4} C_3H_7{-}O{-}CH_2{-}CH_3 + H_2O$$

propyl-ethanoate

In this reaction propanol gives away its $-OH^-$ ion while ethanoic acid releases H^+ ion. Then $-OH^-$ and H^+ join to form a water molecule, which is released from the reaction. The remaining *alkyl group* (**$C_3H_7^+$**) from propanol and the **CH_3COO^-** from ethanoic acid join up into one molecule $CH_3COOC_3H_7$ (*propyl-ethanoate*) is an ester.

When you eat a banana or suck a pear drop or remove nail varnish with a solvent, you sense a *strong* and *foulty smell*, that is an ester. Other esters such as *ethyl-ethanoate* make up the flavour of a fruit. A ripe pineapple contains about 120mg/kg of *ethyl ethanoate*. Fats and vegetable oils, including butter, margarine and corn oils are esters. Terylene is an ester and so are many laminated plastics and surface finishes on kitchen equipment.

An important property of esters is that *they can be split again into the original acid and the alkanol*, this is an example of *hydrolysis*. Water is used to split the molecule. Diluted acid or alkali is used to speed up the reaction, and the mixture is heated.

$$\text{Ester} + \text{water} \longrightarrow \text{Acid} + \text{alkanol.}$$

Example

$$\underset{\text{Ethylethanoate}}{CH_3COOC_2H_5} + H_2O \longrightarrow \underset{\text{Ethanol}}{C_2H_5OH} + \underset{\text{Ethanoic acid}}{CH_3COOH}$$

Exercise 3

1. Explain the differences in IMF between alkanols and alkanes.
2. How do IMFs affect the physical properties of alkanols?

3. Draw the structural formula of nonanol and write down its condensed formula
4. What do you call each of the formulae below:
   a) $CH_3CH_2CH(OH)CH_2CH_3$
   b) $C_7H_{15}OH$
   c)
```
C-C-C-C-C
    |
    OH
```
5. Name each alkanol in question 4 fully.
6. a) What classes of alkanols are the following:
   i)
```
OH-C-C-C
```
   ii)
```
C-C-C-C-C-C
        |
        OH
```
   iii)
```
     CH3
      |
-C-C-C-C-C-
      |
      OH
```
   iv)
```
C-C-C-C-C
        |
        OH
```
   v)
```
      CH3
       |
CH3-C-CH2OH
       |
      CH3
```
   b) Name all the alkanols in (a).
7. Write an equation for the combustion of propanol in air.
8. a) What is esterification
   b) Write down the equation of reaction for the esterification of propanol with butanoic acid. Name the product formed.

Uses of ethanols

Ethanol has a variety of uses both at home and in the industries. The following are some of the uses:

1. Manufacture of alcoholic drinks or beers
2. As a solvent for paints, varnishes, glue, and drugs
3. As the liquid in thermometers
4. As a fuel or sometimes mixed with petrol for vehicles
5. As methylated spirit is used to dress open wound. Methylated spirit is ethanol with small amounts of methanol (*a poisonous substance added to stop people drinking it*).
6. Make food flavourings

However, it has negative health impacts especially when it is taken into the body of a living thing, such as:

1. damages some body organs such as lungs and kidneys
2. may leads to brain impairment
3. in pregnant women it may cause impairment of the foetus (baby).

Unit summary

- Organic compounds are grouped into families called *homologous series*. Alkanols are one kind of homologous series.
- Alkanols contain the *hydroxyl group* (–OH) which acts as a *functional group*.
- As a homologous series, alkanols are linked by the general formula $C_nH_{2n+1}OH$.
- Alkanols are named after alkanes of similar number of carbon atoms.
- The most commonly used alkanol is *ethanol*. It is used as a solvent, fuel, beer and disinfectant.
- Ethanol can be produced by *fermentantion of sugars*. This involves the conversion of sugars into ethanol by the *action of enzymes*, which are found in yeast.
- Ethanol may also be produced by *hydration of ethene*
- Most *physical and chemical properties* of alkanols depend on the size of the molecules. For instance, *smaller alkanols are more soluble in water* than larger alkanols and they also *react faster*.
- Alkanols are soluble in water because they are like water: i.e. *Like dissolves like*.
- Solubility in water is the *test for alkanols*.
- Alkanols *burn in air* to give carbon dioxide and water.
- When the *combustion is complete*, *the colour of the flames is* ***blue***. When the combustion is ***incomplete*** *the flame colour is* ***yellow*** and carbon monoxide and water are produced.
- Alkanols *react with sodium* in a similar way with water.
- *Oxidation* of alkanols form carboxylic acids.

Unit exercise

1. Write the condensed formulae of the following alkanols:
   a) Pent-1-O1
   b) Hexan-3-O1
   c) Nonan-4-O1
2. Ethanol (CH3CH2OH) and ethane (CH3CH3) have different boiling points even though they have the same number of carbon atoms.
   a) Which boils at a higher temperature?
   b) Explaing your answer.
3. a) Ethanol and water can both react with sodium metal. How do the two reactions compare in terms of their:
   i) speed ii) product
   b) Write the chemical equations both reactions.
4. a) Name two processes you can use to prepare ethanol.
   b) Write the chemical equation of each process you mentioned in (a).
5. Explain:
   a) two physical properties of ethanol
   b) two uses of ethanol.
6. a) Complete the equations below.

$$CH_3CH_2OH \xrightarrow[H^+]{K_2Cr_2O_7} \_\_\_\_\_\_\_\_\_\_ + H_2O$$

$$CH_3CHO \xrightarrow[H^+]{K_2Cr_2O_7} \_\_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_$$

   b) This is one process. Why is it done in two steps?

Alkanals and Alkanones

Success Criteria

By the end of this unit you should be able to:

1. Identify the functional groups of alkanals and alkanones
2. Draw and name the structures of the first five alkanals and alkanones
3. Describe the sources and properties of alkanals and alkanones
4. Carry out a test to distinguish alkanals and alkanones
5. Describe the uses of alkanals and alkanones

Key words:

In this unit you will find these key terms and concepts:

Carboxyl group, Aldehyde, Ketone, alkyl group, Nucleophilic addition

Ensure that you understand and learn how to apply them both for your academic and real life situations.

Alkanals and alkanones are two separate homologous series of organic oxycarbons which are closely related. Alkanals are commonly called *aldehydes* while alkanones are called *ketones*. In this unit you will study the differences between aldehydes and ketones, their structural formulae and naming system, their sources, properties and uses.

Functional groups of Alkanals and Alkanones

Just like alkanols, alkanals and alkanones are *compounds of organic oxycarbons* since they contain oxygen in their structure. However, alkanals and alkanones have different functional groups from alkanols. Both alkanals and alkanones have the same functional group called *carbonyl group*. A carbonyl group is the functional group composed of a *carbon atom double - bonded to an oxygen atom* as in *Figure 7.1.*

Figure 7.1: A carbonyl group

The carbon atom of *the carbonyl group in alkanals is bonded to an alkyl group and to a hydrogen atom* as in *Figure 7.2*, where R represents an *alkyl group* or *alkyl radical*. Alkyl groups are hydrocarbons such as *methyl group* (CH_3-), *ethyl group* (CH_3CH_2-) etc.

How can you distinguish alkanals from alkanones?

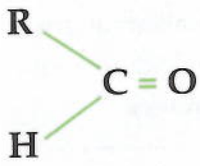
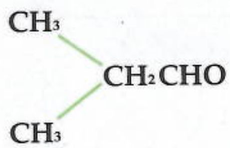


Figure 7.2 An alkanal

From *Figure 7.2*, you can write an alkanal in short as **RCHO**.

Note: Never write as RCOH because it will resemble an alkanol due to the presence of **OH** at the end.

Example



Exercise 1

1. Name and write the molecular formulae of any other four alkyl groups
2. Write the molecular formulae of alkanals with
   a) Methyl group as an alkyl group (R)
   b) Ethyl group as an alkyl group (R)

In alkanones the carbonyl group carbon is bonded to two alkyl groups.

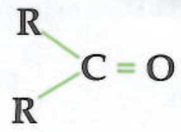
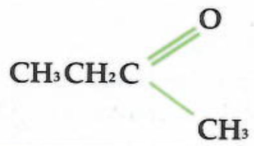


Figure 7.3 An alkanone

Therefore they can be written as **RCOR** where **R** are the alkyl groups. An alkanone can have similar or different alkyl groups.

Example



Nomenclature of Alkanals and Alkanones

Alkanones and alkanals are named differently, but as other organic compounds, both aldehydes and ketones *are named by adding **suffixes** to a **common prefixes***. You may recall that alkanes, alkenes and alkanols are named by adding different suffixes to common prefixes. *Exercise* 2 will help you to remember how to name organic compounds.

Exercise 2

Complete *Table 7.1*.

Molecular formulae	Prefix	Name
CH_4	Meth-	Methane
C_3H_6	Prop-	
		Ethanol
		Nonene
		Butane
$C_8H_{17}OH$	Oct-	
C_5H_{12}		Pentane

Table 7.1

1. Naming Alkanals

When naming alkanals, the suffix *–anal* is added to the organic prefixes. For instance, the alkanal with four carbon atoms (C_3H_7CHO) is called **Butanal**.

Note:

1. *but-* is the prefix used for the *compound with 4 carbon atoms*, while *–anal* is the suffix.
2. To name these aldehydes one must first *know the number of carbon atoms* that are in the chain first.

Table 7.2 gives the first five alkanals. Study it if you have to master the skill of naming them.

Molecular formula	No. of carbon atoms	Prefix	Name
$HCHO$	1	Meth-	Methanal
CH_3CHO	2	Eth-	Ethanal
C_2H_5CHO	3	Prop-	Propanal
C_3H_7CHO	4	But-	Butanal
C_4H_9CHO	5	Pent-	Pentanal

Table 7.2

Some alkanals have special commercial names, for instance *methanal* is called ***formaldehyde***, *ethanal* is called ***acetyldehyde*** and *propanal* is called ***propionaldehyde***.

Exercise 3

1. Name the following aldehydes:
   a) $C_5H_{11}CHO$
   b) C_4H_9CHO
2. Write down the molecular formulae of the following aldehydes
   a) Ethanal
   b) Propanal

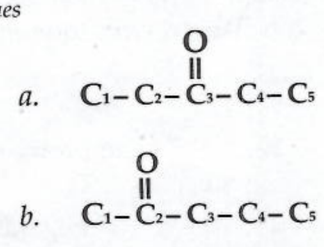
2. Naming Alkanones

Alkanones "*-anone*" suffix. It is added to the common organic prefix. For example, a ketone with four carbons is called *butanone*. *Table 7.3* gives basic names of the first five.

Number of carbons	Molecular formula	Organic Prefix	Name
3	CH_3COCH_3	Prop-	Propanone
4	$C_2H_5COCH_3$	But-	Butanone
5	$C_3H_7COCH_3$	Pent-	Pentanone
6	$C_4H_9COCH_3$	Hex-	Hexanone
7	$C_5H_{11}COCH_3$	Hept-	Heptanone

Table 7.3: names of the first five Alkanones

However, ***the position of the carbonyl group in ketones is not fixed*** as in aldehydes (*always at the end*). In the ketone group, *the carbon atom of the functional group is bonded to two carbon atoms*. Consider the following different skeletal structures of *pentanone* ($C_2H_5COCH_3$).



Skeletal structure of Pentanone

Note: Both are skeletal formulae of a straight chain pentanone, but they are *different* in *respect of the position of the carbonyl group*. By numbering the carbon atoms, the functional group in structure "*a*" is on carbon 3 while in "*b*" is on carbon 2.

The IUPAC system of naming ketones, therefore, *includes the position of the functional group*. Using IUPAC system, structure "*a*" is named as **3 – pentanone** or **petan – 3 – one** where *3 represents the position of the carbonyl group*. Structure "*b*" is then named as **2 – pentanone** or **pentan – 2 – one**.

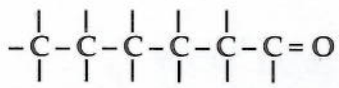
Exercise 4

Draw all possible straight chain skeletal formulae for hexanone and name them.

Structural, condensed and skeletal formulae of Alkanals and Alkanones

As other organic compounds, alkanals and alkanones can be presented as structural, condensed and skeletal formulae.

Consider an alkanal with *6 carbon atoms* whose formula is $C_5H_{11}CHO$. This molecule can be represented as:



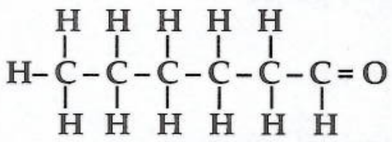
The structure may be **condensed** as follows:

$$CH_3CH_2CH_2CH_2CH_2CHO$$

or

$$CH_3(CH_2)_4CHO$$

Structurally, the molecule $C_5H_{11}CHO$ where all bonds are shown can be drawn as:



If you may recall that the skeletal formula *excludes all hydrogen atoms* <u>except</u> *those that are part of the functional group*. From the structural formula, **the skeletal formula** is:

C – C – C – C – C – C = O

Note: The *last carbon* forms a carbonyl group.

Alkanones have a carbonyl group is bonded to the alkyl groups, hence *they have more than one structure of straight molecule*. For instance, **heptanone** can be presented as:

a.	$C_5H_{11}COCH_3$	which is	2-heptanone
b.	$C_4H_9COC_2H_5$	which is	2-heptanone
c.	$C_3H_7COC_3H_7$;,	which is	4-heptanone.

Note: In all these molecular formulae, there are 7 carbon atoms, 14 hydrogen atoms and 1 oxygen atom.

1. Condensed formulae of ketones

The above molecules can be written as follows:

a) $CH_3CH_2CH_2CH_2CH_2COCH_3$ **or** $CH_3(CH_2)_4COCH_3$

b) $CH_3CH_2CH_2CH_2COCH_2CH_3$ **or** $CH_3(CH_2)_3COCH_2CH_3$

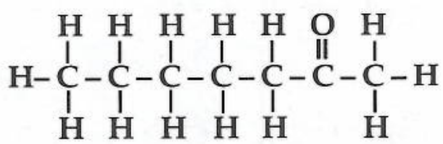
c) $CH_3CH_2CH_2COCH_2CH_2CH_3$ **or** $CH_3(CH_2)_2CO(CH_2)_2CH_3$

These structures represent condensed formulae of 2-heptanone, 3-heptanone and 4-heptanone, respectively.

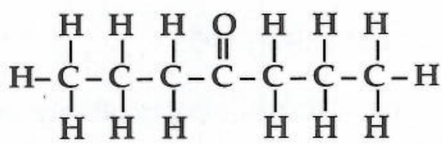
2. Structural formulae of ketones

Structures of keptones can be drawn in the same way as other organic compounds. Consider the examples below:

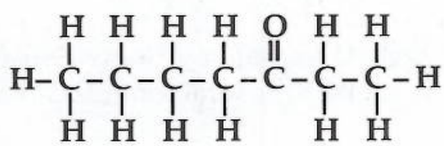
a. $C_5H_{11}COCH_3$



b. $C_3H_7COC_3H_7$



c. $C_4H_9COC_2H_5$



3. Skeletal formulae of ketones

Ketones can also be presented as carbon skeletal formulae, just like aldehydes; where all the hydrogen atoms are not included as shown below:

a. $C_5H_{11}COCH_3$ is presented as:

$$C-C-C-C-C-\overset{\overset{O}{\|}}{C}-C$$

b. $C_4H_9COC_2H_5$ is presented as:

$$C-C-C-C-\overset{\overset{O}{\|}}{C}-C-C$$

c. $C_3H_7COC_3H_7$ is presented as:

$$C-C-C-\overset{\overset{O}{\|}}{C}-C-C-C$$

Exercise 5

1. Draw the structures of the following alkanals:
   a) Methanal
   b) Heptanal.
2. Write the condensed formulae of the following compounds:
   a) Dacanal
   a) Butanone
   a) 2 – heptanone
   a) Petan – 2 – one.
3. Draw the skeletal formulae of the following:
   a) Octanal
   a) 4 – octanal.

Sources of Alkanals and Alkanones

Both Alkanals and Alkanones are *produced from the **oxidation of alkanols***. In Unit 6, you saw that alkanols *undergo an **oxidation process** to produce a compound with a carbonyl group and water*. The carbonyl group containing compound can be a ketone or aldehyde *depending on the class of alkanol used*. You may recall that ***primary alkanols*** are those with hydroxyl group on carbon atom which is attached to less than two other carbon atoms, ***secondary alkanols*** are those in which carbon atom with hydroxyl group is bonded to two other carbon atoms while ***tertiary alkanols*** are those in which carbon atom with hydroxyl group is bonded to three other carbon atoms.

1. Preparation of ethanal

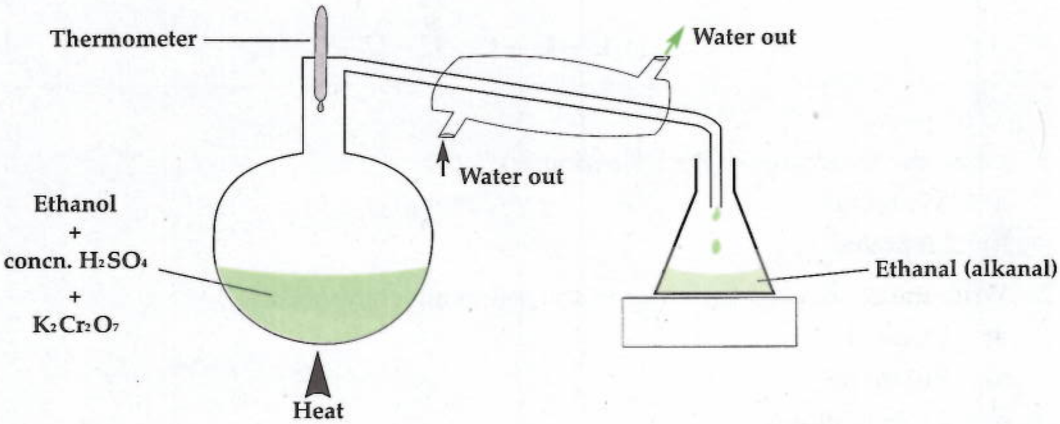
You already seen that *alkanals have a carbonyl group at the end* or *beginning of the molecule*. Now you will see how alkanals are produced from primary alkanols by studying the *preparation of ethanal from ethanol*.

Ethanol is oxidized, by heating it with a *mixture of potassium (IV) dichromate* ($K_2Cr_2O_7$) and *concentrated sulphuric acid* (H_2SO_4). This acidified potassium (IV) dichromate acts as an oxidizing agent. *Ethanol loses two hydrogen atoms* to form ethanal and water. This equation below illustrates this:

$$CH_3CH_2OH \xrightarrow[\text{concn. } H_2SO_4]{\text{Heat, } K_2Cr_2O_7} CH_3CHO + H_2O$$

Ethanol   Ethanal   Water

This reaction takes place in excess ethanol in a simple distillation apparatus. Alkanals have a lower melting and boiling point temperatures. When heating the alkanol with the oxidizing mixture in the boiling tube/flask, alkanals which are more volatile are formed. The alkanal vaporises and rise up the distillation column where it is condensed and collected as distillate in the collecting tube/flask.



2. Preparation of propanone

Propanone is *the smallest ketone* and its common name is ***acetone***. Like aldehydes, ketones are also *prepared from the oxidation of alkanols*. However, *ketones* are made from the oxidation of ***secondary alkanols***. For example, propanone is *prepared from oxidation of secondary propanol* (2-propanol).

The process of making propanone is by *heating secondary propanol in the mixture of acidified potassium (IV) dichromate using simple distillation apparatus*. Propanol loses two hydrogen atoms to produce propanone which vaporizes first from the boiling flask then collected as a distillate. Water is also produced in this process. The equation below describes how propanone is formed.

$$\underset{\displaystyle\ \ \ \ OH}{CH_3\overset{}{C}HCH_3} \xrightarrow[\text{concn. } H_2SO_4]{\text{Heat, } K_2Cr_2O_7} CH_3COCH_3 + H_2O$$

2-Propanol (*secondary alkanol*)   Propanone   Water

Exercise 6

1. Name the organic compound source of the following compounds
   a) Propanal
   b) Heptanone
   c) $C_3H_7COC_3H_7$
   d) $CH_3(CH_2)_4CHO$
2. Draw the structure of a compound formed from the oxidation of 2 – hexanol
3. Differentiate the oxidation process leading to production of an alkanal and alkanones.

Physical properties of Alkanals and Alkanones

Both alkanals and alkanones may be described by a number of physical properties such as *solubility in water*, *melting and boiling points* and *state at room temperature*.

1. Solubility

In *Activity 1* you will investigate the solubility of alkanals and alkanones in water.

Are aldehydes or ketones soluble in water?

Activity 1

Aim: *To investigate the solubility of aldehydes and ketones in water*

Materials:

- *4 test tubes*
- *measuring cylinders*
- *water*
- *propanal*
- *propanone*
- *heptanal*
- *heptanone.*

Procedure:

1. *Put 5cm*^{3} *of water in each of the four test tubes labeled* ***A, B, C*** *and* ***D***
2. *Add 3cm*^{3} *of propanal to test tube* ***A***
3. *Observe the mixture and leave it to settle*
4. *Record your observations in* Table 7.2
5. *Repeat steps 2, 3 and 4 for propanone, heptanal and heptanone in test tubes* ***B, C*** *and* ***D****, respectively.*

Test tube	Observations
A	
B	
C	
D	

Table 7.2: Table of results

Results and Discussion:

1. Which of the test tubes show a single layer?
2. Which test tubes show two layers?

Conclusion:

Propanal and propanone are soluble in water while *heptanal and heptanone are partially soluble* in water. *Small aldehydes and ketones are soluble in water*, but the solubility ***decreases*** *with* ***increasing length*** *of carbon chains*. The longer the chain the lower the solubility. The *solubility of alkanones and alkanals is possible regardless of the absence of* -OH *group*, since they form hydrogen bonds which move with water molecules.

2. Melting and Boiling points

Alkanals and alkanones have high melting and boiling points (*Mpt and Bpt*) as compared to hydrocarbons of similar relative molecular mass (RMM), since they have stronger intermolecular forces (IMF) . For example, butane boils at 0.5°C, Propanal at 49°C while Propanone boils at 56°C. However, ketones have higher melting and boiling points than aldehydes. Passing along the homologous series the molecular size and mass increase, hence the IMF increases, as a result mpt and bpt also increase. *Table 7.3* gives mpts and bpts of the first five alkanals and alkanones.

Alkanals	Mpt	Bpt	Alkanones	Mpt	Bpt
Methanal	-92	-21	Propanone	-95	56
Ethanal	-124	21	Butanone	-87	80
Propanal	-81	49	Pentanone	-78	102
Butanal	-99	46	Hexanone	-55	128
Pentanal	-60	103	Heptanone	-36	151

Table 7.3: Mpts and bpts of the first five alkanals and alkanones

3. State at room temperature

The first member of alkanals, methanal, is gas at room temperature while the rest alkanals are liquids. Meanwhile, all alkanones are liquid at room temperature.

Why is ethanal different from other alkanals? How can you compare the states of alkanals and a

4. Density

Both aldehydes and ketones are less dense than water. They, therefore, float on water. However, the density of these homologous series increases with the size of the molecule. This means that as we go down a homologous series density increases.

Exercise 7

1. Arrange the following alkanals in order of increasing melting and boiling points: Butanal, Decanal, Octanal, Nonanal, Dodecanal.
2. Arrange the following ketones in descending order of densities: Hexanone, Decanone, Propanones, Nonanone, Octanone.
3. Explain why solubility of ketones decreases with the size of the molecule.

Chemical properties of Alkanals and Alkanones

Aldehydes and ketones undergo a variety of chemical reactions that lead to many different products. Both contain a carbonyl group and this makes them undergo similar reactions. The following are some of the chemical reactions which they undergo:

1. Combustion reaction

Both aldehydes and ketones burn in air to produce carbon dioxide and water. For example:

a. **Combustion of propanal**

$$CH_3CH_2CHO_{(l)} + 4O_{2(g)} \longrightarrow 3CO_{2(g)} + 3H_2O_{(l)}$$

b. **Combustion of propanone**

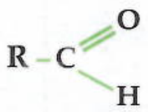
$$CH_3COCH_{3\,(l)} + 4O_{2(g)} \longrightarrow 3CO_{2(g)} + 3H_2O_{(l)}$$

2. Oxidation reaction

Aldehydes are easily oxidized to form an organic acid. In this reaction the aldehyde gains an oxygen atom to form a carboxylic acid The reaction takes place in the presence of oxidizing agents such as acidified potassium (IV) dichromate solution. For example, the oxidation of ethanal occurs as:

$$\underset{\text{Ethanal}}{CH_3CHO} \xrightarrow[\text{concn. } H_2SO_4]{\text{Heat, } K_2Cr_2O_7} \underset{\text{Ethanoic acid}}{CH_3COOH} + \underset{\text{Water}}{H_2O}$$

Remember: Both aldehydes and ketones are produced from the oxidation of alkanols. Further *oxidation is possible in aldehydes to produce carboxylic acid due to the presence of hydrogen atom attached to the carbon that makes double bond with oxygen*.



The *presence of this hydrogen atom makes aldehydes easily oxidize*. Ketones, on the other hand, do not have this particular hydrogen atom, hence cannot get oxidised. This aspect help distinguish ketones from aldehydes.

3. Addition reaction

Alkanals and alkanone *undergo a kind of addition reaction* called ***Nucleophilic addition.*** In this reaction compounds like hydrogen cyanide (HCN), sodium bisulphate ($NaHSO_3$), water (H_2O) and many more are added to ketones and aldehydes. These compounds are called ***nucleophiles***. This reaction is possible due to the presence of the *carbon – oxygen double bond* which breaks when nucleophiles are added. Therefore, *a **nucleophile** is a reacting species which has a tendency to donate electrons or a lone pair of electrons to other reactants*. For example:

a) Addition of hydrogen cyanide (HCN)

When hydrogen cyanide is added to an aldehyde or ketone, compounds known as ***hydroxynitriles*** are produced. To illustrate this let us study the nucleophilic addition reaction of alkanals and alkanones.

i. To alkanals

When hydrogen cyanide is added to ethanal *2–hydroxylpropanenitrile* is produced according to this equation:

Ethanal Cynide 2–hydroxylpropanenitrile

ii. To alkanones

When hydrogen cyanide is added to propanone you produce *2 – hydroxyl – 2 – methylpropanenitrile*, according to the following equation:

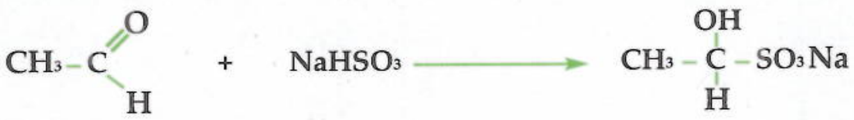
Propanone   Cynide   2–hydroxyl– 2 – methylpropanenitrile

b) Addition of sodium bisulphate ($NaHSO_3$)

Sodium bisulphate is also called *sodium hydrogensulphate*. When it reacts with an aldehyde or ketone, *hydrogensulphite compound* is produced. For example:

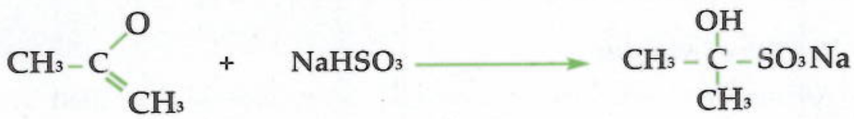
i. To alkanals

When sodium hydrogensulphate is added to ethanal the following reaction takes place:



ii. To alkanones

When sodium hydrogensulphate is added to propanone, the following reaction takes place:



c) Addition of water (H_2O)

When water is added to the *carbon – oxygen double bond*, a hydrated compound with two hydroxyl groups (*diol*) are produced. For example:

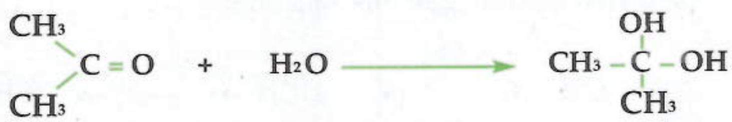
i. To alkanals

When water is added to ethanal, for example, the following is produced:

Ethanal   Water   Ethan-diol

ii. To alkanones

When water is added to ethanone, for example, the following is produced:



Propanone   Water   Propan-diol

Exercise 8

1. Write a balanced complete combustion reaction equation of the following compounds
   a) Butanol
   b) Butanone
   c) Methanal
2. Explain the difference in oxidation reactions between aldehydes and ketones.
3. Write the reaction equation that takes place when hydrogen cyanide is added to 2 – pentanone.

Test for Alkanals and Alkanones

There are different tests that can be used to distinguish alkanals from alkanones. These tests depend on the position of the functional group and their chemical properties.

1. Tests for alkanals

There are a number of tests that are used to identify an alkanal. The following are some of these tests.

a. Brady's Test

This test is used to detect the presence of a *carbonyl functional group* (**– C = O**). This test uses an orange powder called **2,4 dinitrophenylhydrazine**. In short this chemical is called **2,4 – DNPH**.

Method

The following steps are followed when carrying out the Brady's test:

1. Put a match head amount of the orange power 2,4 – DNPH into a clean test tube.
2. Add 2 drops of ethanol to the 2,4 – DNPH in the test tube (*This is done to moisten the powder*).
3. Add 3 drops of concentrated sulphuric acid to the moistened powder in the test tube. (*When this is done a clear yellow liquid is formed.*)
4. Add 1 drop of the liquid under test, i.e. (alkanal) to the test tube.
5. Observe what happens.

Results and Conclusion

When a carbonyl functional group is present in the liquid under test, *a red precipitate is formed*. If a red precipitate is not formed then the liquid tested has no carbonyl functional group. *Both alkanals and alkanones give a positive result* with Brady's test.

b. Fehling's Test

This is another important test for alkanals. Alkanals are easily oxidized to form carboxylic acid when they are treated with acidified dichromate. Hence *this test is used to distinguish alkanals from alkanones*; since alkanones give a negative result.

Method

The following steps are followed when carrying out Fehling's test:

1. Put 15 drops of Fehling's solution into a test tube.
2. Add 15 drops of the liquid under test to the Fehling's solution in the test tube
3. Heat the contents of the test tube gently for a minute

Results and Conclusion

When a test liquid is an alkanal then the blue Fehling's solution turns to a red precipitate. When the test liquid is not an alkanal, then the Fehling's solution remains unchanged.

c. Tollen's Test

This test is also called ***silver mirror's test***. It is also used to distinguish alkanals (aldehydes) from alkanones (ketones). It uses Tollens' reagent which *oxidizes an aldehyde into the corresponding carboxylic acid*. Tollens' reagent is ***a colorless, basic aqueous solution*** containing silver ions coordinated to ammonia $[Ag(NH_3)_2^+]$.

Method

The following steps are followed when carrying out Tollens test.

1. Put 1 ml of Tollen's reagent into a clean oil-free test tube
2. Add 2 drops of the liquid under test to the Tollen's reagent in the test tube
3. Observe the mixture, if there is no reaction heat the contents gently

Results and Conclusion

If the liquid under test is an alkanal then *the silver mirror* or *a black precipitate is formed*. When a liquid under test is not an alkanal then no silver mirror nor black precipitate is formed.

2. Tests for alkanones

Alkanones give positive results with Brady's test. When alkanones are tested using 2,4 DNPH a red precipitate is formed. To carry out this test, the same procedure described under alkanals is followed.

Both Fehling's and Tollen's tests will give negative results for alkanones. This means that these two tests can be used to identify aldehydes from ketones.

Exercise 9

1. Why does the Tollens test not work for hexanone?
2. How can butanone be distinguished from butanal?

Uses of Alkanals and Alkanones

Alkanals and alkanones have a vast number of applications.

1. Alkanals

a. Alkanals such as formaldehyde are used *to preserve dead animals*.

b. They can also be used in manufacturing of synthetic resins like: *bakelite*, *dyes* and *disinfectant materials*.

c. They are used in manufacturing of *organic acid* since alkanals can be oxidized further to produce carboxylic acids.

d. Alkanals are used in the *production of perfumes* and *flavorings*.

e. When alkanals like formaldehydes are treated with fuming nitric acid they are converted into highly *explosive cyclonite*.

f. They are also used for formation of industrial polymers such as *melamine* and *nylon*.

2. Alkanones

a. Acetone, the commonest used ketone. It is used as a *solvent* for resins, lacquers and cellulose and as a *fingernail polish remover*.

b. The ketone, *camphor*, has a pharmaceutical use where it is used in *liniments* and *insecticides*.

c. Ketones are also used in *making perfumes*, *flavorings*, *dyes* and in addition to making of *chloroform* and *plastics*.

d. Alkanones, such as 2 – butanone, are also used as *solvent* and *paint strippers*.

Unit Summary

- Both Alkanals and alkanones have a carbonyl functional group: $R\overset{O}{\overset{\|}{C}}-H$ and $R-\overset{O}{\overset{\|}{C}}-R$, respectively.
- Alkanals are also called *aldehydes* while alkanones are *ketones*.
- The carbonyl group in alkanals *is bonded to less than two alkyl groups* while in alkanones it *is bonded to two alkyl groups*.
- Both aldehydes and ketones are *formed from oxidation of primary* and *secondary alkanols*, respectively.
- Aldehydes are named by adding the suffix *–anal*, while ketones by adding suffix *–anone* to the common organic prefixes.
- Both aldehydes and ketones are *soluble in water*, but their solubility *decreases with size of the molecule*.
- Both aldehydes and ketones have *high melting* and *boiling points* compared to corresponding hydrocarbons because they have relatively high IMF
- Aldehydes can be *oxidized easily* to form carboxylic acids
- Both aldehydes and ketones *undergo combustion* and *nucleophilic additional reactions*
- Brady's test is *used to detect the presence* of carbonyl functional group
- The tests which can be used to distinguish aldehydes from ketones include Fehling's and Tollen's tests which *give positive results with aldehydes* only. While the Brady's test gives *positive results for aldehydes and ketones* because they both contain carbonyl functional group.
- Alkanals are used for *preservation of dead animals* and also for *manufacturing carboxylic acids*, *perfume*, *flavorings* and many more items, while alkanones are used as *solvents* and has *pharmaceutical and industrial use*.

Unit Exercise

1. a) What name do you call the functional group found in both aldehydes and ketones?
   b) Explain how the functional group mentioned in 1(a) differs in aldehydes and ketones.
2. Mention two uses of each of the following:
   a) Acetone
   b) formaldehyde
3. Draw the structure of any ketone with 5 carbon atoms.
4. Complete the following equations:
   a) $CH_3CH_2CHO \xrightarrow[\text{concn. } H_2SO_4]{\text{Heat, } K_2Cr_2O_7}$
   b) $CH_3CH_2CHO + H_2O \longrightarrow$
   c) $CH_3CH_2CH_2COCH_2 + HCN \longrightarrow$
5. Name the following molecules:
   a) $CH_3CH_2CH_2COCH_2CH_2CH_2CH_3$
   b) $CH_3CH_2CH_2CH_2CHO$
   c) HCHO
6. Draw the skeletal formulae for all possible straight molecules for heptanone and name them.
7. Using butanone, write a balanced chemical reaction to describe its combustion reaction with oxygen.
8. Using oxidation, explain how you would differentiate aldehydes from ketones.
9. Suppose you are given a bottle labelled: "*A compound with a carbonyl functional group.*" Describe an experiment you would carry out to classify the compound as an aldehyde or ketone.

Unit 8

Alkanoic Acids

Success Criteria

By the end of this unit you should be able to:

1. Identify the functional group of alkanoic acids
2. Name the first ten unbranched alkanoic acids
3. Draw the structures of first ten unbranched alkanoic acids
4. Write the molecular formula of alkanoic acids given the number of carbon atoms
5. Describe the sources of alkanoic acids
6. Explain the physical properties of alkanoic acids
7. Explain the chemical properties of alkanoic acids
8. Describe the uses of alkanoic acids

Key words:

In this unit you will find these key terms and concepts:

Carboxyl group, Carboxylic acid

Ensure that you understand and learn how to apply them both for your academic and real life situations.

In *Units* 6 and 7 you learnt about some organic compounds such as: *alkanols*, *alkanals* and *alkanones*. All these homologous series are made up of *carbon*, *hydrogen* and *oxygen atoms* hence they are called ***oxycarbons***. In this unit you will study another oxycarbon homologous series called ***alkanoic acids***, also known as ***carboxylic acids*** or ***organic acids***. The other group of acids, commonly used at school is called ***inorganic*** or ***mineral acids***.

At Junior Secondary level you learnt about acids and bases. Now, when we look at the carboxylic acids we will realise that they have properties similar to other acids.

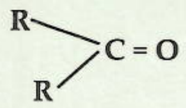
What are some of the properties of acids?

Functional group of carboxylic acids

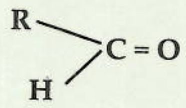
In Form 2, you learnt that *a functional group is an atom or group of atoms contained in a compound that reacts in a characteristic way*. However, alkanes do not have special atom or bond which determines their properties: i.e. they do not have a functional group. *Exercise 1* will enable you to recall the functional groups contained in various organic compounds.

Exercise 1

Identify the functional group that best describe alkanes, alkanols, alkanals and alkanones from the following:

1. 

2. C = C

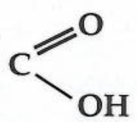
3. 

4. –OH

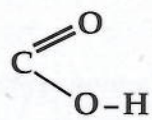
Recall that alkanals and alkanones are carbonyl compounds due to the presence of carbonyl functional group in each. Alkanols, on the other hand, have hydroxyl group (**OH**). Alkanoic acids *have both functional groups of carbonyl and hydroxyl groups* called ***carboxyl group***. This group is written as **–COOH**.

Note: Alkanoic acid are also called carboxylic acids *due to the presence of the carbonyl group* in its functional group.

Structurally, the carboxyl group is drawn as:



or



Usually, this functional group is at the end or beginning of a compound. Consider the following alkanoic acids with 5 carbons.

$CH_3CH_2CH_2CH_2$**COOH** or **HOOC**$CH_2CH_2CH_2CH_3$

A functional group after the carbon chain — **A functional group before the carbon chain**

General formulae of Akanoic acids

Alkanoic acids like other homologous series, have a general formula. The general formula is $C_nH_{2n+1}COOH$, where "n" is *the number of carbon atoms minus one*. The example below illustrates how this general formulae can be used to determine the molecular formulae of the alkanoic acid.

Example: Determine the formulae of an acid with 5 carbon atoms.

Using: $C_nH_{2n+1}COOH$

Where: $n = 5 - 1 = 4$

Then substute n: $C_4H_{2(4)+1}COOH$

$C_4H_{8+1}COOH$

C_4H_9COOH

Note: C_4H_9COOH has 5 carbon atoms including the carbon atom in the functional group (COOH)

Exercise 2

Using the general formula of alkanoic acids, write the molecular formulae of the alkanoic acids with the following numbers of carbon atoms:

1. 1
2. 3
3. 7
4. 9
5. 20

Nomenclature of alkanoic acids

The system used for naming alkanoic acids is common to all organic compounds. In From 2 you learnt how to name alkanes and alkenes, and in previous units you have been learning how you can name alkanols, alkanals and alkanones. In all these activities, the common system of naming is followed, where *the common prefixes and suffixes are used*.

Exercise 3

Name the following compounds:

1. $CH_3CH_2CHCH_2$
2. $CH_3CH_2CH_2CH_2OH$
3. CH_3CH_2CHO
4. $CH_3(CH_2)_4CH_3$
5. CH_3COCH_3

For alkanoic acids, the same naming system is used, where a suffix "*-anoic*" is added to the prefix. For instant, an alkanoic acid with two carbon atoms (CH_3COOH) has *eth-* (as the prefix) and *-anoic* as the suffix. Therefore, it will be *ethanoic acid* (also called *acetic acid*), which is a component in vinegar. Study the following examples:

- $C_7H_{15}COOH$ is Octanoic acid
- $C_8H_{17}COOH$ is Nonanoic acid
- C_4H_9COOH is Pentanoic acid

Note: Octanoic acid has 8 carbon atoms

Exercise 4

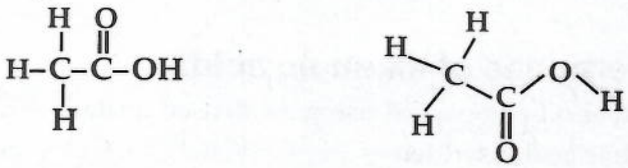
1. Name the following carboxylic acids:
   a) HCOOH
   b) $C_9H_{19}COOH$
   c) C_3H_7COOH
2. Write the formula of the following carboxylic acids
   a) Hexanoic acid
   b) Propanoic acid
   c) Heptanoic acid
   d) Undecanoic acid (*has 11 carbons*)

Structural and skeletal formulae of alkanoic acids

In addition to molecular formular, organic acids can also be presented in different ways. They can be presented as *structural*, *skeletal* and *condensed formulae*. This section will focus on presenting these organic acids in these different formulae.

1. Structural formulae of alkanoic acids

In Form 2, you defined structural formula as the one that *shows how the atoms are arranged and joined in a molecule*. As discussed in alkanols, alkanals and alkanones, alkanoic acids can also be presented as structures. Consider *acetic acid* (CH_3COOH). It has two carbon atoms; *the first carbon is bonded to three different hydrogen atoms*, while the second carbon is bonded to a hydroxyl group and double bonded to oxygen. This is illustrated in *Figure 8.1*.



(*a*) or (*b*)

Figure 8.1: Structural formulae of Ethanoic (Acetic) *acid*

Note: In *Figure 8.1* (*a*) the hydroxyl group (OH) is not expanded, where as in (*b*) the hydroxyl group is expanded (**–O–H**)

Exercise 5

Draw the structural formulae of:

1. Methanoic acid
2. Propanoic acid
3. Butanoic acid.

1. Skeletal formulae of alkanioc acids

You may recall that *skeletal formulae only show the chain of carbons without hydrogen atoms*. However, for organic compounds the hydrogen atom in the functional group is also shown. This means that, the skeletal formulae for alkanoic acids will have **COOH** shown as the *functional group*. *Figure* 8.2 shows the skeletal formula of *pentanoic acid* (C_4H_9COOH).

Note: The numbering of carbon atoms

$$C_5-C_4-C_3-C_2-C_1OOH$$

Figure 8.2: A skeletal structure of Pentanoic acid

Note: The carbon atoms have been numbered from 1 to 5. This helps you to know the number of carbon atoms. *Take note of the numbering system*.

How would you compare the numbering of carbon atoms in Fig. 8.2 and 8.3? Why is the other incorrect?

$$C_1-C_2-C_3-C_4-C_5OOH$$

Figure 8.3: Wrong numbering of carbon atoms in Pentanoic acid

Exercise 6

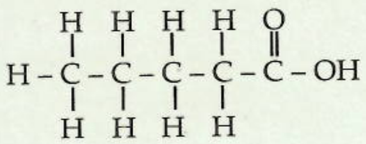
1. Draw the skeletal structures of the organic acids drawn in *Exercise* 5.
2. Name the following carboxylic acids:
   a) C – C – C – C – C – COOH
   b) C – C – C – C – C – C – C – COOH.

3. Condensed formulae of alkanoic acids

Alkanoic acids can also be presented using condensed formula. The condensed formula of pentanoic acid is written as **$CH_3CH_2CH_2CH_2COOH$** or **$CH_3(CH_2)_3COOH$**.

Exercise 7

Write the condensed formula of the following carboxylic acids:

1. Heptanoic acids.
2. An acid with 6 carbon atoms.
3. 

Sources of alkanoic acids

What other substances would you suggest that do have a sour taste?

You have tasted fruits such as mangoes, grapes and oranges. All these have a sour taste. In *Form* 2 you learnt that *acids have a sour taste*; this means that *all foods with a sour taste contain acid*. Organic *acids are found naturally in both plants and animals*. There are different natural sources of different organic acids. For example:

1. *Citric acid* is found in all citrus fruits such as *oranges*, *lemons* and *tangerines*.
2. *Lactic acid* is found in *sour milk*.
3. *Methanoic acid* is found in *insects stings*, such as: *ants*, *bees* and *nettles*.
4. *Carbonic acid* is found in fizzy drinks such as: *lemonade* and *soft drinks*.
5. *Tartaric acid* is found in *grapes* and *baking powder*.

You will find more information in *Table 8.1*.

	Name	Molecular structure	Major source
1	Methanoic (*Formic*) acid	$HCOOH$	ants, bees and nestles
2	Ethanoic (*acetic*) acid	CH_3COOH	vinegar
3	Butanoic (*butyric*) acid	$CH_3CH_2CH_2COOH$	butter
4	Hexanoic acid	$CH_3(CH_2)_4COOH$	cocnut oil, goat's milk
5	Octadecanoic (steeric) acid	$CH_3(CH_2)_{16}COOH$	palm oil, beef fat, cow's milk
6	Citric (2-hydroxypropan-1, 2, 3-tricarboxylic) acid	CH_2COOH | $HO{-}C{-}COOH$ | CH_2COOH	citus fruits, such as: lemon, oranges and lime
7	Lactic (2-hydroxypropanoic) acid	$CH_3{-}CH(OH)COOH$	breast milk, cow's milk
8	Tartaric (2, 3-dihydroxybutanedioc) acid	$OHCHCOOH$ | $OHCHCOOH$	grapes
9	Benzoic (phenylmethanoic) acid	C_6H_5COOH	gums, food additive preservative
10	Oxalic (1, 2-ethanoic) acid	$HOOCH-CHCOOH$	sweet potatoes, tomatoes

Table 8.1: Names and sources of some acids

Exercise 8

Mention the substance which is the source of the following alkanoic acid:

1. methanoic acid
2. butanoic acid
3. tartaric acid.

Apart from these organic acids being found naturally, they can also be made by oxidizing alkanols as already discussed in *Unit 6*. This process is illustrated in the following equation.

$$\underset{\text{Ethanol}}{CH_3CH_2OH} \longrightarrow \underset{\text{Ethanal}}{CH_3CHO} \longrightarrow \underset{\text{Ethanoic acid}}{CH_3COOH}$$

Note: Production of alkanoic acids from oxidation of alkanols involves an *intermediate product*, *alkanal*. This process uses an acidified potassium (IV) dichromate as an oxidizing agent.

Exercise 9

Explain in detail how butanoic acid can be made from butanol.

Properties of alkanoic acids

Alkanoic acids has a number of physical and chemical properties which we are going to discuss in this section.

1. Physical properties of alkanoic acids

Alkanoic acids have a lot of physical properties. Some of these include the following:

a. Solubility in water

Carry out *Activity 1* in order to enble you see what happens to alkanoic acids when mixed with water.

Are alkanoic acids soluble in water?

Activity 1

Aim: *To investigate solubity of alkanoic acids in water*

Materials:

- *Test tubes*
- *ethanoic acid*
- *methanoic acid*
- *propanoic acid and*
- *water*

Procedure:

1. *Put 15 drops of water in the test tube*
2. *Add 10 drops of ethanoic acid in the test tube*
3. *Shake the mixture and wait for the mixture to settle*
4. *Record your observation in the table of results below*
5. *Repeat steps 2 to 4 using methanoic acid and propanoic acid in different test tubes*

Results/Observations:

Complete the Table 8.2 by filling in your observations

Acid	Observations
Ethanoic acid	
Methanoic acid	
Propanoic acid	

Table 8.2: Table of results

What can you say about solubility of alkanoic acids?

Conclusion:

In *Activity 1*, you may have observed that *alkanoic acids form one layer when mixed with water*. This means that alkanoic acids are soluble in water since they do mix with water.

Alkanoic acids contain an hydroxyl group (OH) which is also found in water, hence its solubility in water. Most organic substances that dissolve in water have an OH group. However, the *solubility of alkanoic acids decrease with size*. Those alkanoic acids with long chains of carbon atoms are less soluble in water, as they form two layers. *Alkanoic acids with up to three (3) carbon atoms in their molecules are more soluble* in water.

b. State at room temperature

Alkanoic acids are mostly *liquid at room temperature*. However, some are solid at room temperature, such as: *oxalic* and *benzoic acid*.

c. Viscosity

You may recall that, *viscosity is a measure of how much a liquid resists flowing*. Think how easy it is to pour water compared to syrup or honey. Water flows much faster than the syrup or honey. *The longer the chain, the more viscous the acid is*. Viscosity increases with increase in IMF. As we go across the homologous series, *the IMF increases as well and so does the viscosity*.

Which is more viscous between water and syrup or honey?

d. Boiling points

The boiling point of alkanoic acids increase gradually as the length of the hydrocarbon chain increases. As the length of hydrocarbons increase, the intermolecular forces (IMF) (van der Wall forces) increase, which contributes to the increase in boiling points. *There is also formation of hydrogen bonds between the alkanoic acid molecules*. This is brought about by the ***partial positive*** and ***negative changes*** on the **H** and **O** atoms, respectively, of the carboxylic group (**-COOH**). These charges attract and form *strong hydrogen bonds*, which require more heat to break as shown in *Figure 8.4* (*where* **R** *represents alkyl group*).

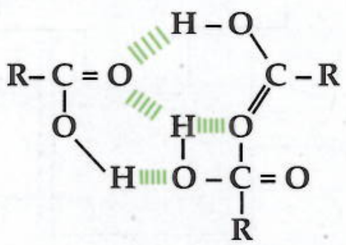


Figure 8.4 Formation of hydrogen bond

Table 8.2 gives boiling points of the first 10 alkanoic acids with their state at room temperature.

Alkanoic acid	Formula	Melting Point (°C)	Boiling Point (°C)	Density g/cm^{-3}	Physical state at room temperature
Methanoic	$HCOOH$	8.4	101	1.220	Soluble
Ethanoic	CH_3COOH	16.6	118	1.049	Soluble
Propanoic	C_2H_5COOH	-20.8	141	0.992	Soluble
Butanoic	C_3H_7COOH	-5.5	164	0.964	Soluble
Pentanoic	C_4H_9COOH	-34.5	187	0.939	Slightly soluble
Hexanoic	$C_5H_{11}COOH$	-4.0	205	0.927	Slightly soluble
Heptanoic	$C_6H_{13}COOH$	-7.5	223	0.918	Slightly soluble
Octanoic	$C_7H_{15}COOH$	16.3	239	0.910	Slightly soluble
Nonanoic	$C_8H_{17}COOH$	12.0	253	0.907	Slightly soluble
Decanoic	$C_9H_{19}COOH$	30.0	269	0.905	Slightly soluble

Table 8.2: Some physical properties of first 10 alkanoic acids

e. Electrical conductivity

Activity 2 will enable you to see whether alkanoic acids conduct electricity or not.

Can alkanoic acids conduct electricity?

Activity 2

Aim: *To investigate electrical conductivity of alkanoic acids*

Materials:

- *Pure ethanoic acid*
- *electrodes*
- *ethanoic acid solution*
- *2 cells, an ammeter*
- *beaker*
- *connecting wire*

Procedure:

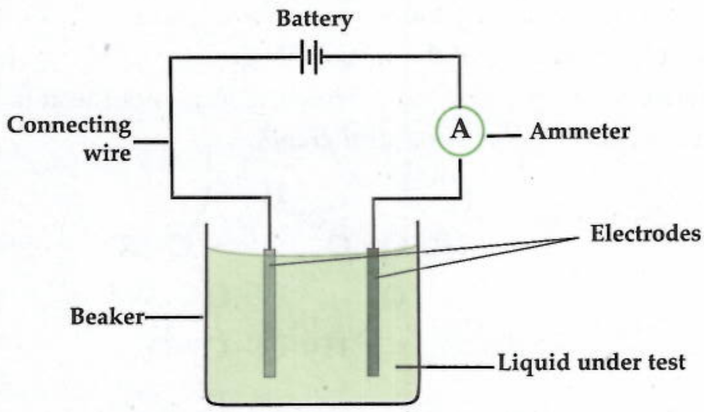


Figure 8.2: Conductivity apparatus for electrolytes

1. *Connect the circuit as in* Figure 8.2.
2. *Put the pure ethanoic acid in the beaker as liquid under test and record the ammeter reading.*
3. *Put the ethanoic acid solution in the beaker as liquid under test and record the ammeter reading.*

Results/Conclusions:

From *Activity 2* you may have noticed that alkanoic acids conduct electricity in aqueous form, but do not do so in their pure state. The reason being that *alkanoic acids dissociate into ions when dissolved in water*, while in pure state there are no ions released. Consider the following dissociation of ethanoic acid in water:

$$CH3COOH_{(aq)} \longrightarrow CH3COO^{-}{}_{(aq)} + H^{+}{}_{(aq)}$$

f. Effects on indicators

Alkanoic acids also have an effect on acid-base indicators. In *Form 2* you learnt about local and commercial indicators. *Activity 3* will help you learn more on the effects of alkanoic acids on acid-base indicators.

Activity 3

Aim: *To investigate the effects of acids on acid–base indicators*

Materials:

- *Test tubes*
- *locally prepared indicator*
- *measuring cylinders*
- *droppers*
- *phenolphthalein indicator*
- *methyl orange indicator*
- *sodium hydroxide (NaOH)*
- *universal indicator*
- *blue litmus paper*
- *ethanoic acid*

Procedure:

METHOD	OBSERVATIONS	CONCLUSION
• Put 5ml of NaOH in a test tube and add 2 drops of phenolphthalein indicator. • To the mixture put 3ml of ethanoic acid		
• Put 5 ml of ethanoic acid in a test tube • Then dip a blue litmus paper indicator		
• Put 5ml of universal indicator in the test tube • Then add 2 drops of the ethanoic acid		
• Put 3ml of Methyl orange indicator into a test tube • Then add few drops of ethanoic acid		
• Put 5ml of local indicator into a test tube • Then add few drops of ethanoic acid		

Table 8.4

Discussion:

From *Activity* 3, you have noticed that alkanoic acids have varying effects on the colours of different indicators. *Indicators are substances whose solutions or state change colour due to changes in* ***pH***; which is, the measure of acidity. This means that the colour of indicators in the activity changes when mixed with the alkanoic acid because it changes their *pH*.

What is the significance of knowing the ph of a substance?

2. Chemical properties of alkanoic acids

Alkanoic acids undergo a number of reactions. In this section you will look at some of these reactions:

a. Combustion reaction

Alkanoic acids burn in air to produce carbon dioxide and water according to the equation:

$$CH_3COOH_{(aq)} + 2O_{2(g)} \longrightarrow 2CO_{2(g)} + 2H_2O_{(l)}$$

b. Esterification reaction

Alkanoic acids react with alkanols to produce ***esters*** in an *esterification reaction*. In this reaction, the alkanoic acid loses its OH while the alkanol loses its H on their functional groups. Then *the organic parts join to form the ester*. Water is then the product of OH and H from alkanoic acid and alkanol, respectively. In this reaction *concentrated sulphuric acid is used as a catalyst*.

$$\underset{\text{Ethanoic acid}}{CH_3COOH_{(aq)}} + \underset{\text{Ethanol}}{CH_3CH_2OH_{(aq)}} \longrightarrow \underset{\text{Ethylethanoate}}{CH_3COOCH_2CH_{3(aq)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

In this reaction, ethanoic acid has reacted with ethanol to produce ***ethylethanoate*** (*ester*) and water. You will learn more of esters in Unit 9.

c. Reaction with bases (Neutralisation reaction)

Neutralisation is *the reaction between an acid and a base* to produce a salt and water. For example, the reaction between *ethanoic acid* and *sodium hydroxide* produces *sodium ethanoate* (*a salt*) and water as shown in this equation:

$$\underset{\text{Ethanoic acid}}{CH_3COOH_{(aq)}} + \underset{\text{Sodium hydroxide}}{NaOH_{(aq)}} \longrightarrow \underset{\text{Sodium ethanoate}}{CH_3COONa_{(aq)}} + \underset{\text{Water}}{H_2O_{(l)}}$$

Note: For an *alkanoic salt formula*, the metal is written last. But when naming the salt, *we start with the metal*.

d. Reaction with carbonates

Acid reacts with metal carbonates to produce salt, carbon dioxide and water. For example, *ethanoic acid reacts with magnesium carbonate* to produce *magnesium ethanoate*, *carbon dioxide* and *water* according to the equation:

$$MgCO_{3(s)} + 2CH_3COOH_{(aq)} \longrightarrow (CH_3COO)_2Mg_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Magnesium Carbonate — Ethanoic acid — Magnesium ethanoate — Carbon dioxide — Water

e. Reaction with metals

Alkanoic acids also *react with alkali metals* and other more *reactive metals to produce alkanoic salt and hydrogen*. In these reactions, the metal displaces hydrogen from the carboxyl group (**–COOH**) of the alkanoic acid. Consider the following examples:

i. Reaction with alkali metal

Ethanoic acid reacts with alkali metals such as *lithium*, *sodium*, *potassium*, e.t.c to form a salt according to the equation:

$$2CH_3COOH_{(aq)} + 2K_{(s)} \longrightarrow 2CH_3COOK_{(aq)} + H_{2(g)}$$

Ethanoic acid — Potassium metal — Potassium ethanoate — Hydrogen

ii. Reaction with other reactive metals

Ethanoic acid also reacts with other metals like *magnesium*. The result is as follows:

$$2CH_3COOH_{(aq)} + Mg_{(s)} \longrightarrow (CH_3COO)_2Mg_{(aq)} + H_{2(g)}$$

Ethanoic acid — Magnesium metal — Magnesium ethanoate — Hydrogen

Uses of alkanoic acids

1. They are used in food preservation because of their effect on bacteria. In this case, the ***pH*** change prevents bacteria from been active and inhibits their growth.
2. They are used in the formation of esters when they react with alkanol according to:

$$CH_3CH_2COOH_{(aq)} + CH_3CH_2OH_{(aq)} \longrightarrow CH_3CH_2COOCH_2CH_{3(aq)} + H_2O_{(l)}$$

Propanoic acid — Ethanol — Ethylpropanoate — Water

3. Alkanoic acids, such as *acetic acids*, are vital in the *production of household vinegar* and it is an *important reagent* for production of *polyvinylacetate* and *cellulose acetate* as well as *synthetic fibers* and fabrics.
4. They are also used as *solvent* for most organic compound.
5. Simple alkanoic acids such as *formic* and *acetic acids* are used for oil and gas well stimulation treatment.
6. Alkanoic acids such as *citric* and *oxalic acids* are used as *rust remover*. They dissolve iron oxides without damaging the base metal.

Unit summary

- The *functional group* of alkanoic acids is called the *carboxyl group*.
- They are also known as *carboxylic acid due to the presences of the carboxyl functional group*.
- They are named by adding the suffix ***– anoic*** to the common organic prefixes.
- Methanoic acid is also called *formic acid* while ethanoic acid is *acetic acid*.
- The general formulae for alkanoic acids is $C_nH_{2n+1}COOH$ where "**n**" is *the number of carbon atoms less one*.
- Natural sources of alkanoic acid include: *citrus fruits*, *grapes*, *baking power*, *sour cow's milk*, *vinegar*, *sting of ant, bee and nettles*, *coke*, *lemonade*, *cheese* and *tea*.
- They are formed from the *oxidation of alkanols* and *alkanals* using acidified potassium (IV) dichromate.
- *First four alkanonic acids are soluble in water* while others are slight soluble. Their *solubility decreases with the size of the carbon chain*.
- As the size of molecule increases, the IMF *increases which affect boiling point*, *viscosity* and other physical properties.
- They *do not conduct electricity* in pure state, **but** in aqueous state.
- They undergo *combustion*, *neutralisation*, *esterification processes* and they react with metals and carbonates.
- Apart from alkanonic acids being used in industries as solvent and for manufacturing different substances they are also *used as food preservatives* and *production of esters*.

Unit exercise

1. Why are alkanoic acids also called carboxylic acids?
2. Draw the structural formular of:
   a) ethanoic acid
   b) hexanoic acid.
3. Mention any three reactions which organic acids undergo in where an organic salt is a product.
4. Write the molecular formula of an Alkanonic acid with 9 carbon atoms.
5. Alkanoic acids have many applications. Explain any three applications.
6. Explain how decanoic acid is formed from decanol.
7. Explain why:
   a) butanoic acid melts and boils at a lower temperature than hexanoic acid.
   b) ethanoic acid (CH_3COOH) is more soluble in water than pentanoic acid ($CH_3CH_2CH_2CH_2COOH$).
   c) alkanoic acids are generally soluble in water.
8. Butanoic acid is added to the same volume of butanol in the presence of concentrated sulphuric acid and warmed.
   a) What name is given to the reaction that takes place upon mixing?
   b) Name the products of this reaction.
   c) Write the equation of the reaction that takes place.
9. Mention any four natural sources of alkanoic acids.
10. Complete and balance the following reaction equations:
   a) $CH_3COOH_{(aq)} + K_{(s)} \longrightarrow$ __________ + ________
   b) $CH_3CH_2COOH_{(aq)} + Mg_{(s)} \longrightarrow$ __________ + ________
   c) $HCOOH_{(aq)} + Na_2CO_3 \longrightarrow$ __________ + ________
   d) __________ + ________ $\longrightarrow CH_3CH_2COOCH_3 + H_2O$

10. Explain an experiment you would carry out to investigate the difference in electric conductivity between concentrated and dilute ethanoic acid.

Unit 9

Alkanoates

Success Criteria

By the end of this unit you should be able to:

1. Identify the functional group of alkanoates
2. Name and draw the structures of alkanoates
3. Describe the sources of alkanoates
4. Describe the properties of alkanoates
5. State uses of alkanoates
6. Describe the process of soap making (*saponification*)

Key words:

In this unit you will find these key terms and concepts:

Functional group, numenclature, condensation, hydrolysis, saponification

Ensure that you understand and learn how to apply them both for your academic and real life situations.

Alkanoates are organic compounds that also form a homologous series. They are commonly called *esters*. Their molecules are made of *a combination of carboxylic acids and alkanols*, just as we saw in esterification of carboxylic acids.

Functional group

Alkanoates contain a functional group RCOOR 'where R and R' are ***alkyl groups***. *This functional group is responsible for their properties*. Its structure is as *Figure 9.1*.

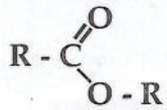


Figure 9.1 The functional group of alkanoates

One of the *simplest esters is ethylethanoate* ($CH_3COOCH_2CH_3$). It is made from the reaction between ethanol and ethanoic acid. To show the bonding in the molecule, the structure of ethylethanoate can be written as in *Figure 9.2*.

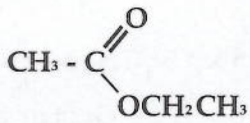


Figure 9.2 Structure of ethylethanoate

Exercise 1

1. What is the functional group of alkanoates?
2. Name the homologous series from which alkanoates are made

Naming the structures of alkanoates

Alkanoates are *the products of **condensation reaction** between carboxylic acids and alkanols called **esterification***. They are *named by regarding them as **alkyl derivatives** of carboxylic acids*. So the name of an alkanoate comes from a **stem of the name of the alcohol** with the suffix ***–yl*** followed by the **stem of the name of the acid** with the suffix ***–oate***.

Therefore, an ester derived from *methanol* and *propanoic acid* is called **methyl propanoate**.

You can work out the name of an alkanoate from its reactants following these steps:

1. *Divide the formula into two parts* by drawing a line across the bridging of the group:

   R_1COO/R_2

2. Name the part that does not carry the **–COO group** an *alkyl group* with the suffix ***–yl***

   R_2 (*Alkyl*)

3. Name the portion that carries the **–COO group** as *carboxylate group* with suffix ***–oate***

   R_1COO (*Alkanoate*)

4. Combine 2 and 3 to give the name of the ester.

   R_1COOR_2 (*Alkylalkanoate*)

Example

What is the name of $CH_3CH_2COOCH_2CH_3$?

Solution

Following the following steps:

Step 1: Divide the formula into two parts CH_3CH_2COO/CH_2CH_3

Step 2: The part that does not have COO is *ethyl* because it contains two carbon atoms:

$-CH_2CH_3$
ethyl

Step 3: The portion with $-COO^-$ is *propanoate* because it contains three carbon atoms:

CH_3CH_2COO-
propanoate

Therefore, when you combine step 2 and 3 the compound is *ethyl propanoate*.

Note:

1. Following steps 1 to 4 you can name any alkanoate that you are given.
2. The name of an alkanoate ***starts with the stem of an alkanol*** and ***ends with the stem of a carboxylic acid***.

Table 9.1 has more names of alkanoates depending on their stems. now you are required to complete the table.

	Structure	Name
1	$HCOOCH_3$	methylmethanoate
2	$HCOOCH_2CH_3$	ethylmethanoate
3	$HCOOCH_2CH_2CH_3$	propylmethanoate
4	CH_3COOCH_3	
5	$CH_3COOCH_2CH_3$	
6	$CH_3COOCH_2CH_2CH_3$	
7	$CH_3COOCH_2CH_2CH_2CH_3$	
8	$CH_3CH_2COOCH_3$	
9	$CH_3CH_2COOCH_2CH_3$	

Table 9.1: Complete the the naming process

Draw structures of alkanoates given alkanoic acid and alkanol

You can draw the structure of any alkanoate if you are given the alkanoic acid and alkanol from which it is formed. You will be able to do this using the naming system illustrated in the first part of this section.

Note:

1. We remove the last –H of –COOH from a carboxylic acid and the –OH from the alkanol.
2. We then combine what remains; *starting with the part of a carboxylic acid*, then *the part of an alkanol*.
3. The –H and –OH combine to form the water molecule.

Example

Write the structures of alkanoates formed from the following:

1. CH_3COOH and CH_3OH
2. CH_3CH_2COOH and CH_3OH
3. $CH_3CH_2CH_2COOH$ and CH_3CH_2OH
4. $CH_3CH_2CH_2CH_2COOH$ and $CH_3CH_3CH_2CH_2OH$

Solutions

The first two have been done for you, now complete the remaing two.

1. $CH_3COO\mathbf{(H)} + CH_3\mathbf{(OH)} \longrightarrow \mathbf{CH_3COOCH_3}$
2. $CH_3CH_2COO\mathbf{(H)} + CH_3\mathbf{(OH)} \longrightarrow \mathbf{CH_3CH_2COOCH_3}$ or $\mathbf{C_2H_5COOCH_3}$
3. $CH_3CH_2CH_2COO\mathbf{(H)} + CH_3CH_2\mathbf{(OH)} \longrightarrow$
4. $CH_3CH_2CH_2CH_2COO\mathbf{(H)} + CH_3CH_2CH_2\mathbf{(OH)} \longrightarrow$

Note: The bracketed atoms are *removed during esterification* and combined to form water molecules.

Deducing reactants of esterification given structure of product

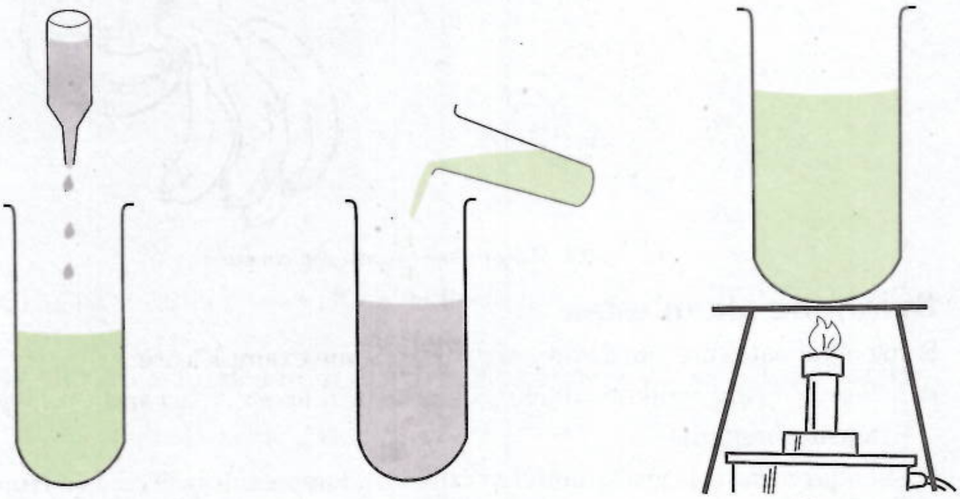
The structure of an alkanoate has a ***carboxylate group*** at the beginning and finishes with an ***alkyl group***. You can, therefore, deduce the structures of the reactants if you are given the structure of the product.

For example, an alkanoate $CH_3CH_2CH_2COOCH_3$ (*methylbutanoate*) is a product of $CH_3CH_2CH_2COOH$ (*butanoic acid*) and CH_3OH (*methanol*). To do this you count the carbon atoms before the two oxygen atoms to identify the carboxylic acid $\overset{4}{C}H_3\overset{3}{C}H_3\overset{2}{C}H_2\overset{1}{C}OOCH_3$. You also count the carbon atoms after the two oxygen atoms $CH_3CH_3CH_2COO\overset{1}{C}H_3$ to identify the alkanol.

Warning: *Acids can corrode the skin and clothing; use protective wear.*

Procedure:

1. *Measure 2cm³ of ethanol using the measuring cylinder and pour it into the test tube.*
2. *Add 3 drops of sulphuric acid to the ethanol*
3. *Add 1cm³ of ethanoic acid to the mixture*



(*a*) (*b*) (*c*)

Fig. 9.5 preparation of ethylethanoate

4. *Heat the mixture as in* Fig. 9.5 (*c*)
5. *After heating pour the mixture into the 150ml beaker containing water*
6. *Smell the mixture in the beaker*

Results / Observations

1. **What smell does the heated mixture produce?**
2. **How many layers are in the beaker?**
3. **What substance have you produced?**

Discussion

The reactions of a carboxylic acid and an alkanol to form alkanoate *takes place in the presence of concentrated sulphuric or hydrochloric acid as a catalyst*. The mixture is heated to make the reaction fast. After heating, the mixture is poured into cold water to cool the alkanoate that has been formed. *The substance you have now formed has a sweet fruity smell*.

The product you have formed from the activity above is an alkanoate.

What alkanoate is produced in Activity 1?

The following chemical equation will help you to understand the *Activity 1*.

$$\underset{\text{ethanoic acid}}{CH_3COOH} + \underset{\text{ethanol}}{CH_3CH_2OH} \longrightarrow \underset{\text{ethylethanoate}}{CH_3COOCH_2CH_3} + \underset{\text{water}}{H_2O}$$

The activity you have just carried out is a ***condensation process*** of a *carboxylic acid* and an *alkanol*. The *product* is an alkanoate called *ethylethanoate* with *water as a by-product*. The condensation process of a carboxylic acid and alcohol with an alkanoate as a product is called ***esterification***.

Exercise 3

1. Name two sources of natural esters.
2. Name the compounds that cause flavour in bananas.
3. Explain why in esterification:
   a) a concentrated acid is added
   b) the mixture is heated.

Properties of alkanoates

Most properties of alkanoates are different from those of other families of organic compounds. Alkanoates have the following properties. They:

1. ***Are colourless volatile liquids***
2. ***Are insoluble in water***

   In *Activity 1*, *you formed two layers* (*Fig. 9.6*); when you poured out the alkanoate, you remained with water. The formation of two layers, therefore, tells us that *the alkanoate is not soluble in water*. However, *solubility of an alkanoate can be improved **if temperature of water is raised***.

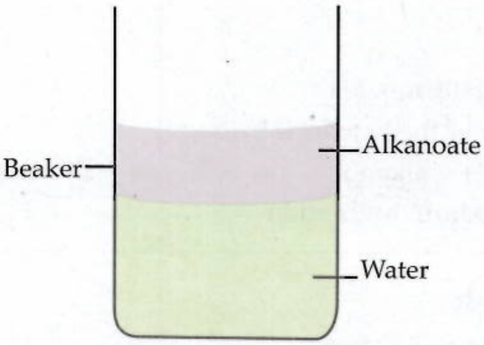


Fig 9.6: Two layers formed when an alkanoate is mixed with water

3. ***Have a good/sweet smell and flavour***

   Alkanoates have a sweet smell. Bananas, oranges and guavas naturally contain alkanoates that give them the good smell and flavour that they have.

1. What is the smell of ethylethanoate?
2. Name 6 fruits that do have a good flavour

4. *Have low melting and boiling points*

Unlike acids and alcohols, *alkanoates do not have free –OH groups*. Hence they *cannot form hydrogen bonds* which make the melting and boiling points of substances high. They therefore, *have lower melting and boiling points* than acids and alkanols.

5. *Combustion*

Alkanoates *burn in air to produce a bright flame* with water and carbon dioxide produced in the process. The equation below illustrates this:

$$CH_3COOH_2CH_3 + 5O_2 \longrightarrow 4CO_2 + 4H_2O$$

6. *Hydrolysis with water and sodium hydroxide*

Alkanoates *react with water in a hydrolysis reaction*. The products of hydrolysis of esters with water *are an alkanol and a carboxylic acid from which they were formed*. So hydrolysis of esters is the *reverse reaction of esterification*. Look at the equation below:

$$\underset{\text{ethylethanoate}}{CH_3COOCH_2CH_3} + \underset{\text{water}}{H_2O} \longrightarrow \underset{\text{ethanoic acid}}{CH_3COOH} + \underset{\text{ethanol}}{CH_3CH_2OH}$$

Esters *are effectively hydrolysed in alkaline solution* to form an *alkaline salt* and *alkanol*. Study the equation below:

$$\underset{\text{ethylethanoate}}{CH_3COOCH_2CH_3} + \underset{\text{sodium hydroxide}}{NaOH} \longrightarrow \underset{\text{sodium ethanoate}}{CH_3COONa} + \underset{\text{ethanol}}{CH_3CH_2OH}$$

Note: This reaction is speeded up by heating the mixture.

The *hydrolysis of esters in an alkaline solution* is called *saponification*; a process used in soap making.

Exercise 4

1. State two properties of alkanoates
2. Give the name of each of the reactions below:
   a) Carboxylic acid + alkanol $\longrightarrow$ alkanoate + water
   b) Alkanoate + sodium hydroxide $\longrightarrow$ sodium alkanoate + alkanol

Uses of alkanoates

The following are the uses of alkanoates:

1. *Solvents*

Some alkanoates are used as *solvents of adhensives*. *Ethylethanoate*, for example, is a solvent for *polysterene*.

2. *Flavourings*

Some alkanoates are *used to give sweets and drinks a fruit flavour*. These fruit flavoured sweets and drinks, such as *yoghurt* and *juices* are very common in shops.

3. ***Fragrances***

Some alkanoates are used in soaps, petroleum jellies and perfumes *to give them a sweet fragrance*.

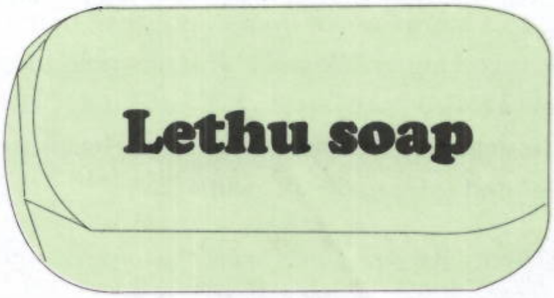


Fig 9.7: Bath soap contains esters

4. ***Fabrics***

Some esters especially polyester can be used for making fabrics (*cloth*), e.g. *terelyne*.

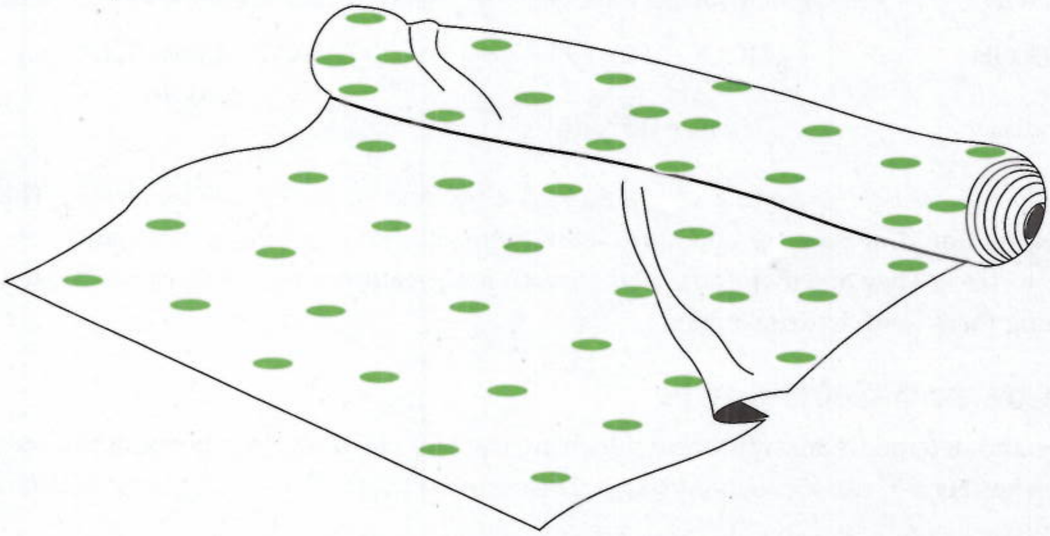


Fig 9.8: Terylene cloth (*one of the alkanoates*)

5. ***Cooking oils***

Cooking oils such as *Mulawe*, Sunflower oil and *Kukoma* contain alkanoates. These alkanoates provide good flavour to the food and make it more appetising.

Fig 9.8: Cooking oil

6. ***Soap making***

Some esters are used for making soap and detergents.

Exercise 5

1. State three uses of esters.
2. What is responsible for the good smell in lemons, apples and grapes?

Fats and oils

Fats and oils are natural esters. They are used as *energy storage compounds* in both plants and animals. They all belong to a large class of *lipids*. As esters, they are made through the same process of condensation; involving an alkanol and large molecules of acids. Most fats are *tri-esters* derived from alkanol called *propane -1, 2, 3-triol* ($OHCH_2 - CH_2OH - CH_2OH$) and a long-chain of a carboxylic acid, e.g. HOOC $(CH_2)_{16}$ CH_3. The alkanol and carboxylic acids are commonly called *glycerol* and *fatty acids*, respectively.

A *glycerol has three* **–OH**, hence it *requires three molecules of a fatty acid* to combine with, to form a molecule of *fat* (*alkanoate*) and *three molecules of water are produced as a* **by-product**. Study the equation below:

$$\begin{array}{c} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array} + \left\{ \begin{array}{c} HOOC(CH_2)_{16}CH_3 \\ HOOC(CH_2)_{16}CH_3 \\ HOOC(CH_2)_{16}CH_3 \end{array} \right\} \longrightarrow \begin{array}{l} CH_2OOC(CH_2)_{16}CH_3 \\ | \\ CHOOC(CH_2)_{16}CH_3 \\ | \\ CH_2OOC(CH_2)_{16}CH_3 \end{array} + 3H_2O$$

glycerol (*an alkanol*) — fatty acids (*carboxylic acids*) — fat (*alkanoate*) — water

Oils *contain a large proportion of unsaturated acids* and *have low melting points*. They are, therefore, liquid at room temperature. However, **fats** *contain saturated acids* and *have higher melting points*. They are, therefore, solid at room temperature. ***Oils can be converted to fats*** by reacting them with hydrogen gas.

Soaps and detergents

Soaps and detergents are substances which are used for cleaning. Soap is one of the detergents which has been in use for so many years. The *main raw materials for making soap are* ***esters***, such as *fats from animals* or *oils extracted from plants*.

Soap making

How is soap manufactured in factories?

When fats or oils *are boiled with an alkaline solution*, such as sodium hydroxide or potassium hydroxide, they *are broken down into glycerol and a sodium salt of a long chained carboxylic acid* which is **soap**. This ***the saponification process***. The word equation for this reaction is:

glycerol stearate (fats) + **sodium hydroxide** $\xrightarrow{\text{heat}}$ **sodium stearate** (soap) + **glycerol**

The chemical equation for this reaction is:

$$\left. \begin{array}{l} CH_2OOC(CH_2)_{16}CH_3 \\ | \\ CH_2OOC(CH_2)_{16}CH_3 \\ | \\ CH_2OOC(CH_2)_{16}CH_3 \end{array} \right\} + 3NaOH \longrightarrow 3CH_3(CH_2)_{16}COONa + \left\{ \begin{array}{c} CH_2OH \\ | \\ CHOH \\ | \\ CH_2OH \end{array} \right.$$

Glycerol stearate (*fats*) — **sodium hydroxide** — **sodium stearate** (*soap*) — **glycerol**

The main product in this reaction is **soap** (*sodium stearate*). After such a reaction, **glycerol** (*a by-product of saponification*) must be separated from the main product.

Soap is separated from glycerol *by adding sodium chloride solution to the mixture*. A layer of soap then forms on the surface which can be skimmed off.

You can, therefore, see that soap making is a simple process that you can even carry out in the laboratory.

Activity 2

Aim: *Investigating soap making*

Materials:

- *10cm³ of concentrated sodium hydroxide solution*
- *2cm³ of cooking oil*
- *measuring cylinder*
- *200 ml beaker*
- *large spoon*
- *funnel*
- *filter paper*
- *conical flask*
- *stirring glassrod*
- *gas burner*

Warning: *Acids and alkali can damage the skin and cause wounds. Wear protective clothing*

Procedure:

1. *Measure up to 2cm³ cooking oil and put it in a beaker*
2. *Measure up to 10cm³ of concentrated sodium hydroxide solution, then add it to the oil*
3. *Heat the mixture gently while stirring with a glass rod*
4. *Add two spoonful of sodium chloride to the mixture*
5. *Add 10cm³ of distilled water to the mixture, then boil it again*
6. *Cool the mixture, then filter it to recover the soap.*

You can follow the procedure above by studying the figure below:

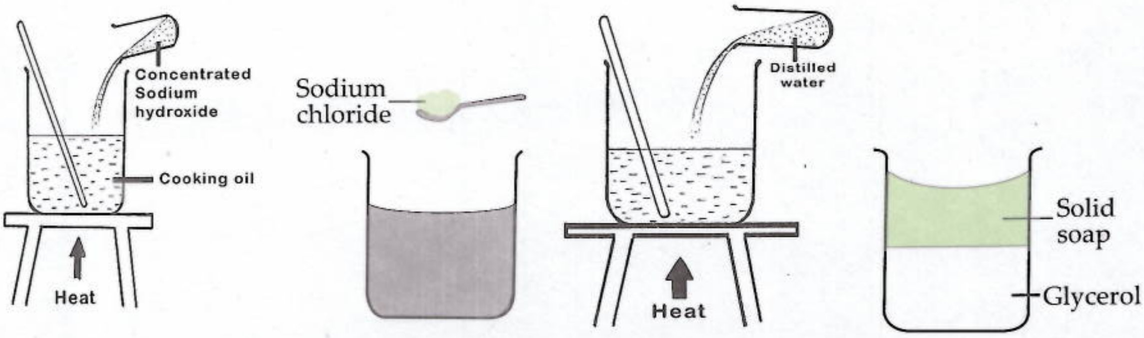


Fig 9.9: Soap making in the laboratory

From the mixture, the soap is left on the filter paper, while the glycerol is the filtrate.

Activity 3

Education visit

Visit one of the soap manufacturing companies, in your area, to see how they make soap on a large scale. Ask the authorities about the raw materials and the procedure they use. Compare what you learnt at the factory with soap making in the laboratory, that you just practiced in Actiivity 2.

Exercise 6

1. What compounds are the reactants that form soap?
2. Why is the addition of sodium chloride solution necessary in soap-making?
3. Complete the following equations:
   a) Fats + sodium hydroxide solution + Heat $\longrightarrow$ – – –
   b) Glycerol l + fatty acids $\longrightarrow$ – – – – – – –

Unit summary

- The functional group of alkanoates is – RCOOR
- The name of an alkanoate *starts with the stem of an alkyl group* and *ends with the stem of carboxylate group*.
- To draw the structure of an alkanoate we *remove the last* –H of a carboxylic acid and *the last* –OH of an alkanol, then combine the two stems.
- You *can deduce the structures of the carboxylic acid and an alkanol* from which an alkanoates is formed by checking how many carbon atoms are before the two oxygen atoms of the alkanoate to give the name of the carboxylic acid. You also count how many carbon atoms are after the oxygen atoms to give the name of an alkanol
- Alkanoates *have a sweet smell*
- There are *natural esters found in plants as oils and in animals as fats*. There are also *synthetic alkanoates* e.g. polyesters
- Alkanoates are *used as fruit flavourings* in sweets and drinks.
- Ethylethanoate is *used as a solvent*.
- Alkanoates are *made through a condensation reaction* of an alkanol and a carboxylic acid
- *Reaction of an alkanoate and alkaline solution* is called hydrolysis. Its products are an alkanol and an alkali salt of a carboxylic acid.
- The *hydrolysis of fats or oils produces a soap* as a main product. This process is called ***saponification***.

Unit Exercise

1. Name the following alkanoates:
   a) $CH_3COOCH_2CH_2CH_2CH_2CH_2CH_3$
   b) $C_3H_7COOCH_3$
   c) $C_5H_{11}COOC_5H_{11}$
2. Give the structures of the following alkanoates:
   a) Butylbutanoate
   b) Hexylmethanoate
   c) Nonylethanoate
   d) Propylpentanoate
3. Give the structures of organic compounds from which the following esters are made:
   a) Ethylbutanoate
   b) Methylmethanoate
   c) $CH_3COOCH_2CH_2CH_3$
   d) $HCOOCH_2CH_2CH_2CH_2CH_3$
4. You are given three solutions of organic compounds in three tubes that are not labelled. But you are told that one of them is an alkanoate. How can you identify the alkanoate in the test tubes given?
5. State two commercial uses of esters.
6. Complete the equations by filling in the gaps.
   a) CH_3COOH + ---- $\longrightarrow$ CH_3COOCH_3 + -----
   b) CH_3CH_2OH + --- $\longrightarrow$ $HCOOCH_2CH_3$ + H_2O
   c) ----- + ------- $\longrightarrow$ $C_7H_{13}COOC_2H_5$ + H_2O
   d) Fats + sodium hydroxide + Heat $\longrightarrow$ ---- + ---
   e) $CH_3COOCH_2CH_3$ + --- $\longrightarrow$ $CH_3COO\,Na$ + ---
7. Give the names and structures of the main products in the following reactions:
   a) CH_3COOH + CH_3OH $\longrightarrow$
   b) $CH_3CH_2CH_2CH_2COOH$ + $CH_3CH_2CH_2CH_2OH$ $\longrightarrow$
   c) HCOOH + Na OH $\longrightarrow$
   d) Fats + sodium hydrodxide $\longrightarrow$
   e) CH_3COOCH_3 + H_2O $\longrightarrow$
   f) Propylpropanoate + water $\longrightarrow$
8. Describe the saponification process.

Identification of unknown compounds

Success Criteria

By the end of this unit you should be able to:

1. deduce the family and structural formula of an unknown organic compound
2. distinguish organic compounds basing on their properties
3. carry out tests for identifying different homologous series

Key words:

In this unit you will find these key terms and concepts:

physical property, chemical property, solubility test, acid test, bromine test, braddys test, flow diagrams

Ensure that you understand and learn how to apply them both for your academic and real life situations.

Remember that organic compounds have at least carbon and hydrogen atoms in their molecules. Almost all of them have functional groups. The organic compounds that have same functional group *belong to the same homologous series*. Each family has similar physical and chemical properties. In this unit you will study *how to deduce the family of organic compounds from their* ***general formula***. You will also learn how to identify different organic compounds by carrying out different tests and using flow diagrams.

Recall: *What did you say is an organic compound?*

Deducing the family of organic compounds

The family of the organic compounds can be deduced using different methods; among which are using the general formula or products of a reaction.

1. *How many families of organic compounds do you know so far?*
2. *What is common amongst the members of each family?*

1. General formula

It is easier to identify the family of organic compounds using the general formula. *The general formula of a family of organic compounds represents all the members of the family and has a functional group*. The functional group of a ***homologous seri orges*** determines the physical and chemical properties of the members.

	General formula	Name of homologous family C_nH_{2n+2}
1	C_nH_{2n+2}	Alkanes
2	C_nH_{2n}	Alkenes
3	$C_nH_{2n+1}OH$	Akanol
4	$C_nH_{2n+1}COOH$	Carboxylic acid
5	$C_nH_{2n+1}CHO$	Aldehydes
6	RCOR	Ketones
7	RCOOR	Esters

Table 10.1: General formula of homologous families

In *Table 10.1* shows that *each general formula has a* ***functional*** *group* **except** *for alkanes*. The subscript "**n**" stands for any *number of carbon atoms*. R and R are the alkyl groups.

2. Products of chemical reactions of some families

The *products of a chemical reaction can be used to deduce* or *identify the family of an organic compound*. Different organic compounds react with different substances to produce different products.

a. Alkanes

Alkanes are generally *un-reactive because they have no functional group*. This can be used to distinguish them from other compounds.

b. Alkanols

Alkanols *react with sodium to produce a salt* and *hydrogen gas*. Look at the equation below:

Alkanol + sodium $\longrightarrow$ salt + hydrogen

$2CH_3CH_2OH + Na \longrightarrow 2CH_3CH_2ONa + H_2$

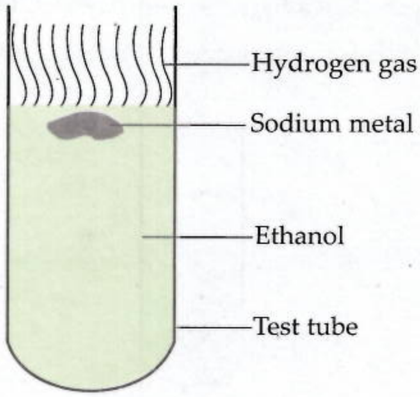


Figure 10.1 Reaction of alkanol with sodium metal

This reaction is possible because of the functional group (-OH) of the alkanols which makes them to behave like water (H_2O).

c. Alkanoic acids (carboxylic acids)

Alkanoic acids react with different substances to produce different products.

i) ***With metals***

Alkanoic acids react with some metals to produce a salt and hydrogen gas because of the presence of ***-OH group*** (-COOH) in them.

carboxylic acid + metal $\longrightarrow$ salt + hydrogen gas

$2HCOOH_{(aq)} + 2Na_{(s)} \longrightarrow 2HCOONa_{(aq)} + H_{2(g)}$

methanoic acid **sodium metal** **sodium methanoate** **hydrogen gas**

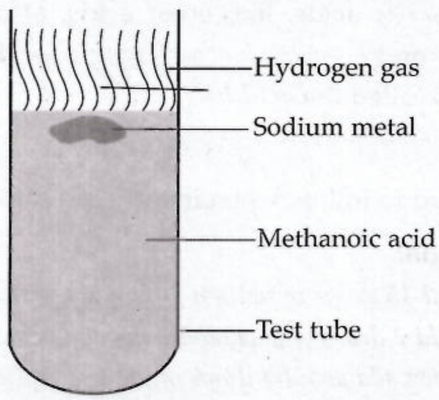


Figure 10.2: Reaction of alkanoic acid with sodium metal

ii) With carbonates

Alkanoic acids *react with carbonates* to produce *a salt*, *carbon dioxide* and *water*.

alkanoic acid + carbonate ⟶ salt + water + carbon dioxide

Look at the following chemical equation:

$$CH_3COOH_{(aq)} + CaCO_{3(s)} \longrightarrow (CH_3COO)_2Ca_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

ethanoic acid **calcium carbonate** **calcium ethanoate (*salt*)** **carbondioxide** **water**

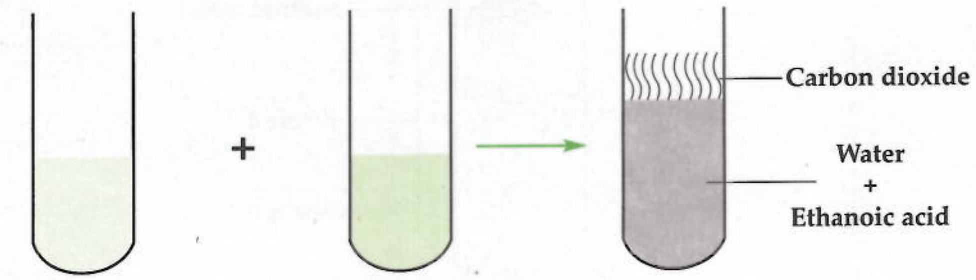


Figure 10.3: Reaction of alkanoic acid with carbonates

iii) With alkanols

Carboxylic acids react with alkanols to produce esters and water. For example,

carboxylic acid + alkanol ⟶ ester + water

$$CH_3COOH_{(aq)} + CH_3CH_2OH_{(aq)} \longrightarrow CH_3COOCH_2CH_{3(aq)} + H_2O_{(l)}$$

ethanoic acid **ethanol** **ethylethanoate** **water**

Note:

1. *The reaction of a carboxylic acid with alkanol produces a sweet fruity smell.*
2. *It is a slow and incomplete reaction.*

Identification of carboxylic acids: Acid test

Carboxylic acids, like other acids, affect the colours of indicators, such as *litmus* and *phenolphthalein*. This property is used to identify carboxylic acids and is called the *acid test*.

Method to follow when identifying carboxylic acids using the Acid test

Method:

1. *Put 15 drops of sodium hydroxide solution in a test tube.*
2. *Add 2 drops of phenolphthalene indicator (a).*
3. *Then add about 6 drops of the test liquid (b).*

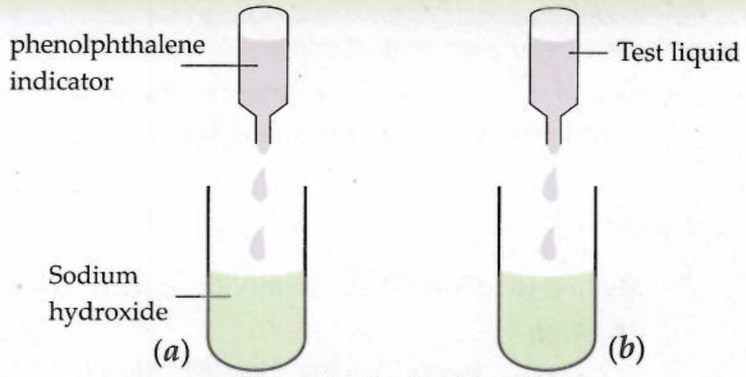


Figure 10.4: Results of the acid test

Expected results:
When *the pink colour in the test tube disappears*, the liquid is a carboxylic acid.

d. Alkenes

Alkenes are very reactive hydrocarbons because of their *carbon double bond carbon* (C=C) *functional group*. They *react instantly with **halogens** to form **haloalkanes***. For example, hexene reacts with bromine to form dibromohexane.

$$\underset{\textbf{hexene}}{CH_3(CH_2)_2CH{=}CHCH_3} + \underset{\textbf{bromine}}{Br_2} \longrightarrow \underset{\textbf{dibromohexane}}{CH_3(CH_2)_2CHBrCHBrCH_3}$$

This reaction of alkenes with bromine is used to identify alkenes. It is called ***bromine test***.

Method to follow when identifying alkenes using the Bromine test

Method:
1. *Put 15 drops of bromine solution, Br_2 in a test tube*
2. *Add 2 drops of the test liquid.*

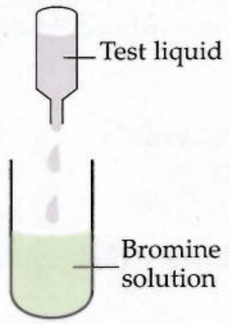


Figure 10.5: Bromine test

Expected result:
If the test liquid is an alkene, *the bromine solution changes its colour* from ***brown*** to ***colourless***.

e. Aldehydes

In Unit 7, you saw that *aldehydes belong to the carbonyl group* because of the presence of –C=O in their structure. They react with Fehling's solution (*blue in colour*) and produce a red precipitate. The *Fehling's test*, therefore, *is used to identify the aldehydes*.

Method to follow when identifying aldehydes using the Fehling's test

Method:

1. *Place 15 drops of Fehling's solution into a test tube (a).*
2. *Add about 15 drops of a test liquid (b).*
3. *Heat for a minute (c).*

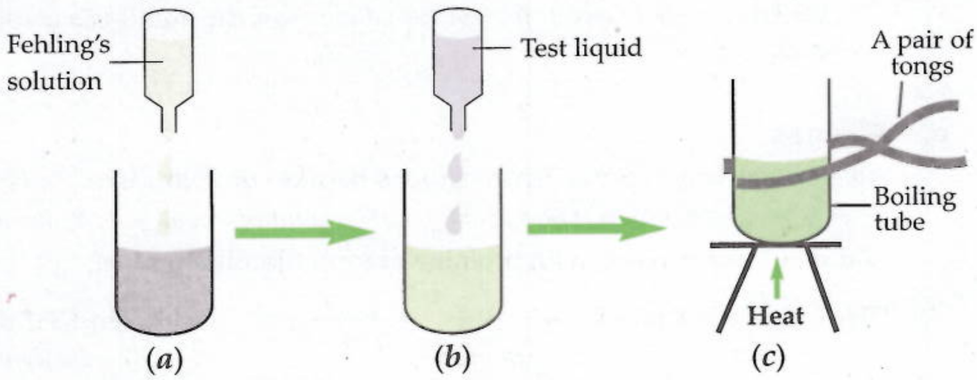


*Figure 10.6: Test tubes **a**, **b** and **c** show the results of Fehling's test*

Expected result:

When the test liquid is an aldehyde, a red precipitate will be produced in the test tube during heating.

f. Ketones

Ketones, like aldehydes, have a carbonyl group (–C = O) in their structure. *But this carbonyl group is not at the end of a carbon skeleton* as is the case with aldehydes. The following are some of them.

Recall: *From Unit 7, do you remember any keytone?*

	Formula	Name
1	CH_3COCH_3	Propanone
2	$CH_3CH_2COCH_3$	Butanone
3	$CH_3CH_2CH_2COCH_3$	Pentanone

Ketones do not react with Fehling's solution. So it cannot be used to identify them.

Method to follow when identifying ketones using the Brady's test

Recap:

In this test, an *orange powder* of 2, 4–DNPH (2, 4–*dinitrophenylhydrazine*) is used. *Both aldehydes and ketones give a positive result*. A red precipitate is produced when either an aldehyde or a ketone is mixed with this powder.

Method:

1. *Place a match-head amount of 2, 4 – DNPH powder in a test tube.*
2. *Add 2 drops of ethanol to moisten the powder in a test tube.*
3. *Add 3 drops of concentrated sulphuric acid. A yellow liquid is formed.*
4. *Add 1 drop of a test liquid.*

Expected result:

When the test liquid is a ketone, a **red precipitate** will be formed.

g. Soluble organic compounds

The common physical property of organic compounds is their solubility in water. When an organic compound is added to water in a test tube; if there is only one layer in it, then the compound is soluble. However, if there are two layers, then it is insoluble.

Method to follow when identifying soluble organic compounds: ***Solubility test***

Method:

1. *Put about 15 drops of water in test tube.*
2. *Add 2 to 3 drops of the test liquid.*

Expected result:

- *When test liquid is **soluble in water**, there will be one layer in the test tube.*
- *When the liquid is **not soluble**, two layers are formed.*

Summary of physical and chemical properties

The chemical and physical properties of organic compounds discussed in this section are summarized in the *Table 10.2*:

Organic compound	Physical properties	Chemical properties
Alkanes	• *insoluble in water* • *small-molecule alkanes are gases at room temperature, while larger molecules are liquids or solids*	• generally unreactive • burn in air
Alkenes	• *insoluble in water* • *small-molecule alkenes are gases at room temperature, while larger molecules are liquids or solids*	• changes red bromine to colourless • burn in air

Table 10.2: Physical and chemical properties of Organic Compounds

Organic compound	Physical properties	Chemical properties
Alkanols	• *soluble in water. The solubility decreases with increasing size of molecules* • *liquid at room temperature*	• reacts with sodium to release hydrogen gas • burn in air
Carboxylic acids	• *soluble in water but solubility decreases with increasing size of molecules*	• turn phenolphthalein in sodium hydroxide from pink to colourless • react with metals to produce a salt and hydrogen gas • react with carbonates to produce a salt, carbon dioxide and water • React with alkanols to produce esters
Alkanals (aldehydes)	• *soluble in water. Large molecule alkanals are insoluble*	• is readily oxidised in Fehling's solution to produce red precipitate
Alkanones (ketones)	• *soluble in water. Large molecule ketones are insoluble*	• is readily oxidized to red precipitate in 2,4 DNPH

Table 10.2: Physical and chemical properities of Organic compounds

Exercise 1

1. State one chemical property of:
   a) alkenes
   b) carboxylic acids.
2. Describe an acid test.
3. Name a test you can do to distinguish the following:
   a) pentane from pentanol
   b) butanone from propanal
   c) hexane from pentene.

Identifying organic compounds using flow diagram

In the previous section you studied physical and chemical properties of different families of organic compounds. These properties can be used to identify them using a *flow diagram*. A flow diagram *is like a guide a chemist follows in order to identify different organic compounds*. This is established by carrying out tests one after the other. The tests carried out are: *solubility*, *acid*, *Fehling's*, *Brady's* and *bromine*, depending on the available families to identify.

How to work out a flow diagram

The names of the compounds to be identified are written on the left side of the flow diagram, as in *Figure 10.7*. You then carry out the following steps to tests one property after the other:

Step 1: *Solubility test*

This is the first test that is carried out. This test will enable you to separate soluble compounds from insoluble ones. The soluble compounds give *one layer when mixed with water* (*positive result*), however, *two layers* means the compound is insoluble (*negative result*). The *soluble compounds are written on the top of the flow diagram*, while the insoluble ones on the bottom.

Step 2: *Acid test*

This test often comes after solubility test. It *identities a carboxylic acid* from the soluble compounds separated in the first step. The *organic compounds that give negative results are written down and identified further*.

- *How do you test for an acid?*
- *What is a positive acid test?*

Step 3: *Fehling's test*

This test *identifies aldehydes*. It can come after the acid test. The *positive result of this test is the formation of the red precipitate* when the aldehyde is oxidized.

Step 4: *Brady's test*

Brady's test can come after Fehling's test. It is used *to identify ketones*.

Step 5: *Bromine test*

The *insoluble liquids* that were separated in the solubility test *can now be identified using bromine test*. Usually, this test is used to *distinguish an alkene from alkane*.

Example 1

You are provided with 6 beakers that are not labelled. Each of the beaker contains one of the following compounds: heptane (*alkane*),hexene (*alkene*), ethanoic acid (*carboxylic acid*), ethanol (*alkanol*), propanal (*aldehyde*) an butanone (*ketone*). Draw the flow diagram that you can use to identify the compounds.

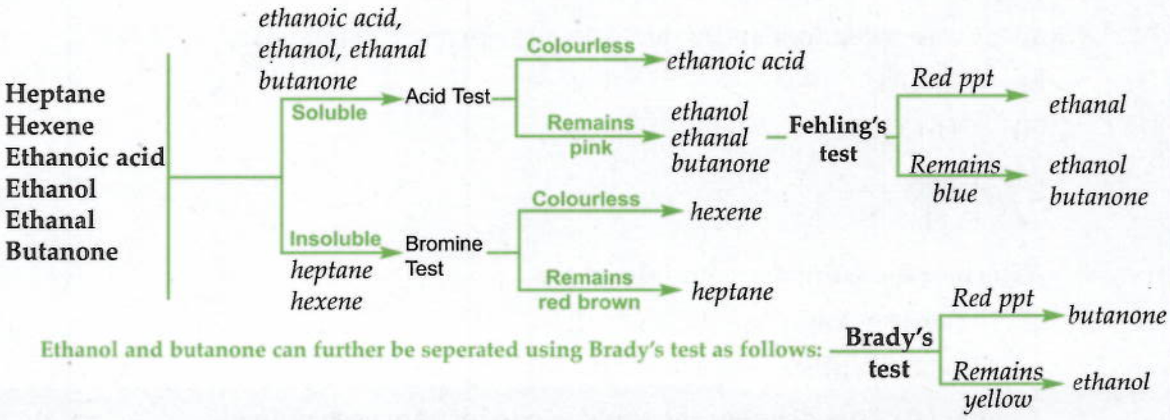


Figure 10.7 A flow diagram used to identify organic compounds

Note:

1. The *kind of test you start with* and the *number of tests* you carry out *depend largely upon what homologous series* of organic compounds you have. In *Example 1*, the first test was solubility because some organic compounds are soluble while others are insoluble.
2. The tests carried out included *acid*, *Fehling's*, *Brady's* and *bromine tests* because the organic compounds to be identified were so many and required all these tests.

Example 2

Your Chemistry teacher gave you three unlabelled beakers; each containing hexane, hexene and ethanoic acid, respectively. Draw the flow diagram you can use to identify them. Study the example give in *Figure 10.8*.

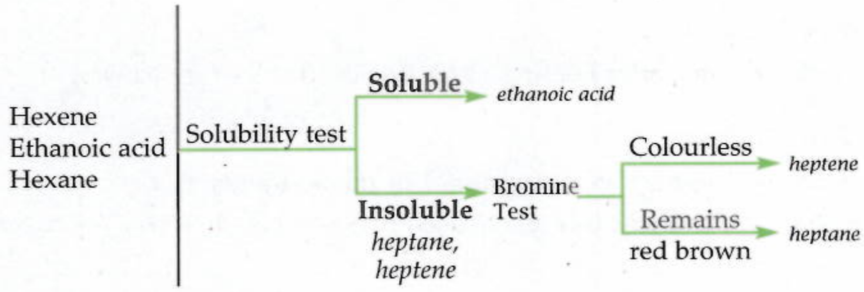


Figure 10.8 A flow diagram

In this second example, there are only two tests that are carried out. The first test is the solubility test. This is because of the few organic compounds given.

Exercise 2

1. A student was asked to identify the following organic compounds:
   - i) CH_3OH
   - ii) CH_3COOH
   - iii) C_5H_{10}
   - iv) C_7H_{16}

   a) Which of these organic compounds are
   - i) Water-soluble
   - ii) Water-isoluble

   b) Construct a flow diagram she could use to identify each of them.
2. Complete the flow diagram below by filling in the spaces:

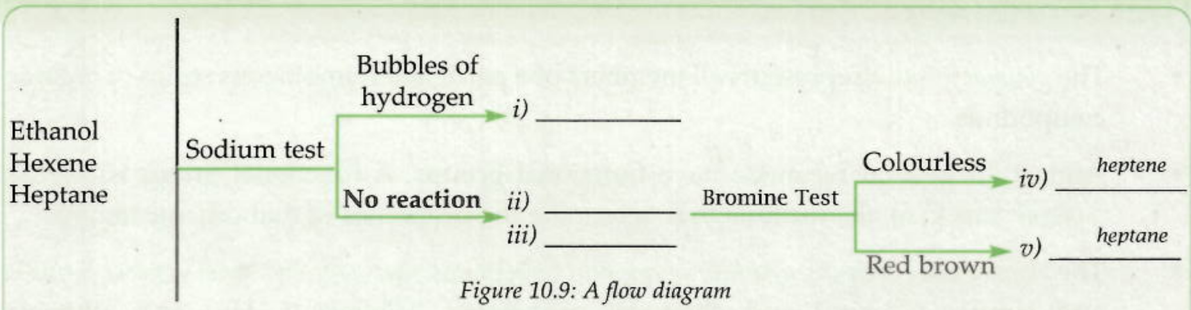


Figure 10.9: A flow diagram

Unit Summary

- The *general formula* represents all members of a particular homologous series of organic compounds.
- Almost all general formulae have functional groups. A functional group is *a set of 'special' atoms* (**n**); the formula that determine the properties of that organic family.
- The *homologous series is a family of organic compounds that have the same general formula* and similar physical and chemical properties. Alkenes (C_nH_{2n}) and alkanols (C_nH_{2n+1})OH are examples of some homologous series.
- Sodium test is *used to separate organic compounds which have* ***–OH group*** *in their molecules from those that do not have*. When sodium reacts with compounds with –OH group, *hydrogen gas is released*.
- *Alkenes react with bromine solution* changing its colour from red-brown to colourless. Bromine test is, therefore, *used to distinguish alkenes from other organic compounds*.
- Fehling's solution, which is blue in colour, *oxidizes the aldehydes producing the red precipitate in the process*. This reaction can be used to identify aldehydes. It is called Fehling's test.
- A flow diagram is *a guide which you can use to identify a number of organic compounds* by carrying out different tests.

Unit exercise

1. Which of the general formulae below belongs to alkenes?
   a) C_nH_{2n+2}
   b) $C_nH_{2n+1}OH$
   c) $C_nH_{2n+1}COOH$
   d) C_nH_{2n}

2. Complete the reactions below:
   a) $CH_3CH = CH_2 + Br_2 \longrightarrow$
   b) $CH_3OH + CH_3COOH \longrightarrow$
   c) $CH_3CH_2OH + Na \longrightarrow$
   d) $CH_3COOH + K \longrightarrow$

3. Describe the procedure for each of the following tests:
   a) Bromine test
   b) Brady's test
   c) Solubility test.

4. Write down the functional groups of the following homologous series:
   a) alkenes
   b) carboxylic acids
   c) alkanols
   d) ketones

5. In the flow diagram below, choose the possible organic compounds that could be represented by the letters M, N O, P and Q.

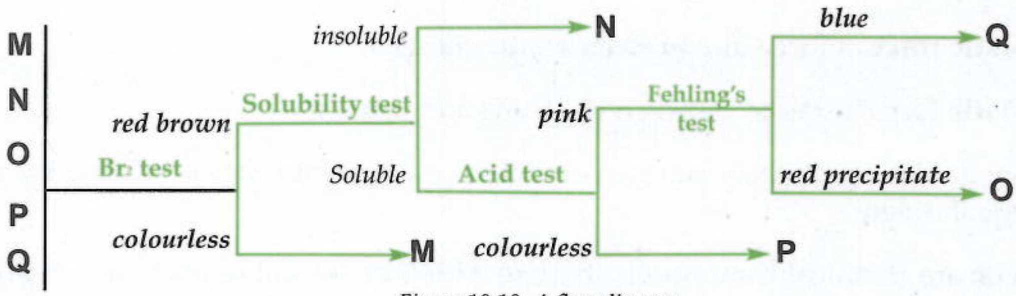


Figure 10.10: A flow diagram

6. Mary was asked by her teacher to construct a flow diagram for identifying propanone, methanol, ethanoic acid and heptane. Construct a possible flow diagram which she drew if she scored full marks for her answer.

Glossary

Acid: a solution with *pH* lower than 7

Acid-base indicator: a substance that can show whether a solution is an acid or base

Atmosphere: air around the earth

Base: a solution with *pH* above 7

Bond: energy that binds two or more atoms together to form a molecule

Catalyst: a substance that helps a chemical reaction to take place

Chemical properties: Characteristics of of a substance in terms of how it reacts with other substaces

Coefficient: a number in front of the variable or a chemical formula

Compounds: are made of more than one kind of atoms

Compressible: able to decrease in volume without changing in mass

Conductivity: ability to conduct electricity or heat

Contact process: a process of making sulphuric acid

Drying agent: a substance that absorbs water from something

Electrolyte: ionic liquid that can conduct electricity

Electron: subatomic particle which has a negative charge

Electronegative: the effect of electrons in the outermost shell

Electrostatic force: a force due to electrostatic charges

Electrostatic force: e xerted between electrons and protons

Equation: a relationship between two quantities on opposite sites separated by an equals sign.

Fertilizers: are chemical compounds that are added to the soil to make it rich in nutrients

Flavours: food or drink additives that chage itastes

Flow diagram: a step by step flow of events illustrated in diagram form

Formula: shows the numbers and kinds of atoms that make up a compound

Fractional distillation: separating the mixture of liquids into separate liquids

Fractional distillation: Separating a liquid mixture into its components

Fragrance: pleasant smell

Haloalkane: the product formed when an alkenes reacts with a halogen

Heat energy: makes atoms of a substance vibrate and a substance increases in temperature

Homologous series: a family of compounds having the same chemical properties

Hydrolysis: a reaction with water

Hydroxyl group: OH group

Intermolecular force: a force that holds molecules together

Lattice: a structure or pattern of arrangement

Mineal acid: an acid made from mineral salts e.g HCl, H_2SO_4

Oleum: concentrated sulphuric acid

Refrigerant: a cooling agent

Relative formula mass: mass of a formula as compared to the mass of carbon-12 atom.

Reversible reaction: a reaction that can return from products to form reactants

Root nodules: small round lumps on the roots

Solubility the rate of dissolving of a substance in a solvent

Synthetic: not natural

Test: an activity carried out to identify a substance using its property

Tetrahedral: a four: sided figure with triangular sides

Transitional temperature: temperature at which a substance can change to another form

Upward delivery: collecting a gas using a beaker or test-tube positioned upside down

Volatile: that can easily change into another form.

Yield: resulting amount of an intended product of a process

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