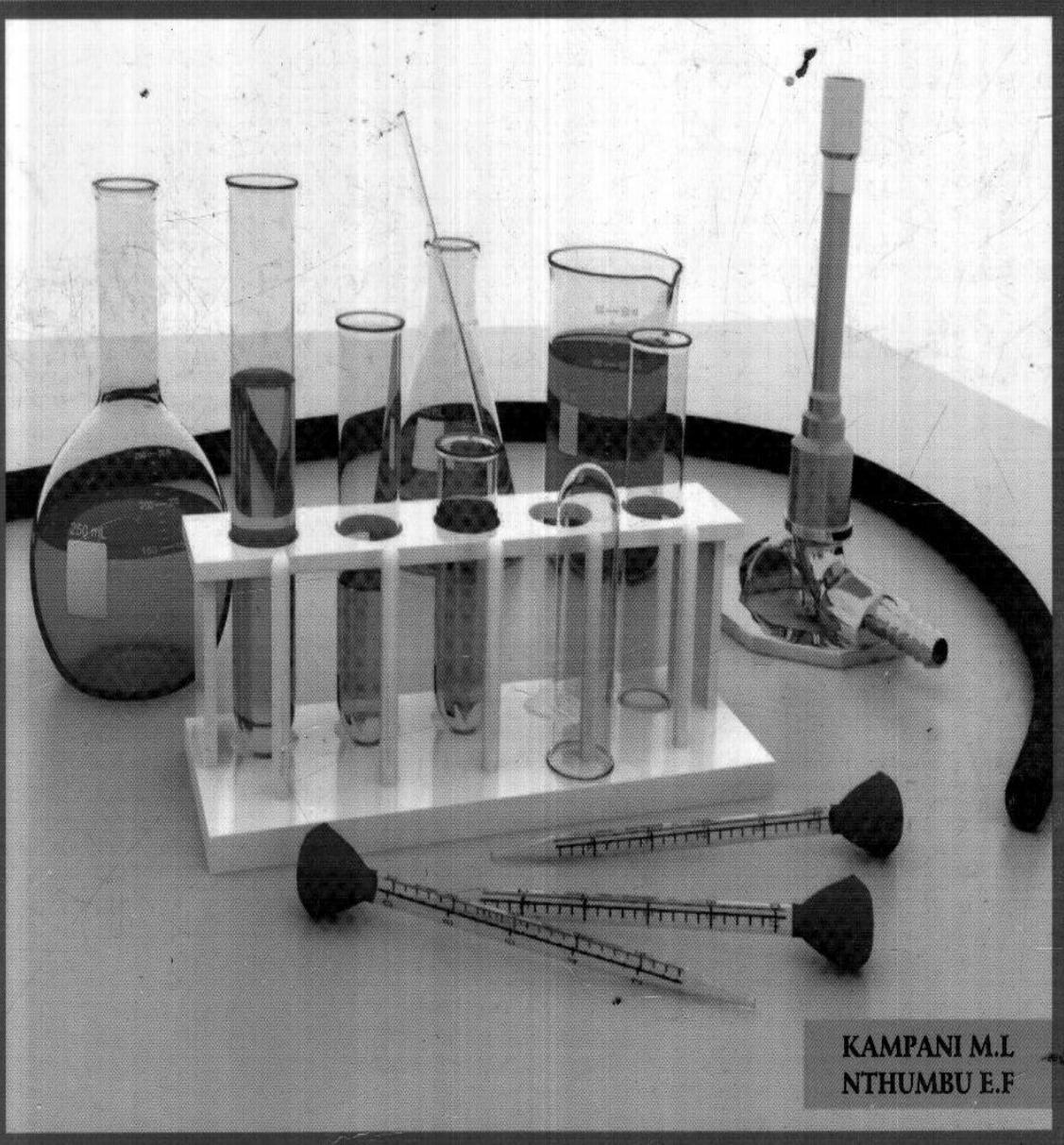


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JHANGO SENIOR CHEMISTRY BOOK 3



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Jhango

Senior Chemistry

Learners' Book

For

Form 3

Kampani M. L. & Nthumbu E. F.

Jhango Publishers Ltd



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Core Element 1: Analytical skills in chemistry

Topic 1: Experimental techniques

Topic 1: Experimental techniques

Experimentation is an integral part of studies in science. Science is both **the process** of gaining knowledge and **the knowledge** gained through scientific investigations. Because of this, you need to actively participate in all investigations in your chemistry lessons.

In this topic you will learn how to carry scientific investigations. You will also learn how to handle chemical wastes in the laboratory to avoid causing harm to our health and the environment.

By the end of this topic you must be able to:

- 1) describe safe and environmentally friendly ways of chemical waste disposal.
- 2) design scientific investigation.
- 3) carry out an investigation to determine purity of substance.

1.1: Ways of disposing of chemical wastes in the laboratory

Chemical wastes

Chemicals of various kinds are kept in the laboratories for use during scientific investigations. Chemicals can also be generated in the laboratory in the process of carrying out experiments during the scientific investigations.

Chemicals generated in the laboratory during experimental activities, chemicals which are in unlabelled bottles and chemicals that are no longer needed are known as **chemicals wastes**. Chemical wastes can be in the form of solid, liquid or solutions which may contain positively and negatively charged ions.

Exercise 1.1

1. In Forms 1 and 2 you carried out a number of experiments in the laboratory. List down examples of chemical wastes that are produced as a result of activities in the laboratory.
2. Classify the chemical wastes into the following groups:
 - a) organic compounds
 - b) positive ions
 - c) negative ions

Disposing of chemical wastes

All chemical wastes generated in the laboratory should be treated as hazardous to health and the environment and, because of that, proper disposal procedures must be followed when getting rid of them. Most chemical wastes generated in the secondary school laboratories are of the types that can be safely disposed of down laboratory sink drains (Figure 1.1) and in landfills. However, before disposing of them in the laboratory sink drains, they must be classified basing on their properties such as corrosiveness, flammability and acids and bases to avoid mixing incompatible wastes, such as mixing alcohols with acids. Some of the procedures for disposing of chemical wastes in the school laboratory are as follows:

A. Chemical wastes that can be disposed of down the laboratory sink drains

The following chemical wastes can be disposed of down the laboratory sink drains (Figure 1.1).

- 1) Acids with a pH of at least above 3 and bases with a pH of at least below 8. This means strong acids and bases need to be neutralised first, before disposing of them in the sink drains.
- 2) Solutions of compounds whose both ions are on the list in Table 1.1.
- 3) Liquid organic compounds listed in Table 1.1. Alcohols, such as ethanol, are flammable, therefore, they must be diluted to concentrations of less than 24% before discharging them into sink drains (see Section 4.6).



Figure 1.1: Laboratory sink drain

Table 1.1: List of chemicals that can go into laboratory sink drain

Positive ions	Negative ions	Organic compounds
Aluminium	Bromide	Ethanoic acids
Ammonium	Carbonate	Acetone
Calcium	Chloride	Pentanols
Hydrogen	Hydroxide	Butanols
Iron	Iodide	Propanols
Lithium	Sulphite	Esters
Sodium	Sulphate	Alkanes
Magnesium	Nitrate	Alkenes
Zinc	Cyanate	Ethanol
Copper	Borate	Methanol
		Sugars

B. Chemical wastes that can be disposed of in landfills

If solid chemical wastes generated in the laboratory are less than 100kg per month, they can be disposed of in landfills. Some liquid wastes can also be solidified or packaged properly and then taken to landfills.

Exercise 1.2

- 1. Explain which ones of the chemical wastes you mentioned in Exercise 1.1 can go into the laboratory drain?**
- 2. Suggest the benefits of proper disposal of chemical wastes to:**
 - a) our health.**
 - b) environment.**

1.2: Designing experiments for scientific investigation

Science follows a sequence of logical steps to generate new ideas, answer questions, and draw conclusions. The logical steps followed in scientific investigations are referred to as **scientific method**. The scientific method is possible in chemistry because of '**cause and effect relationship**' that exists in nature. Scientific investigation is a search for the cause and effect relationships in nature.

Conducting a scientific investigation in chemistry

When you have a problem to be investigated the following are the steps to be followed:

Step 1: Writing the statement of the problem

At this stage, you identify the problem that requires an answer. The problem must be stated in the form of a question. For example, you might wish to investigate if the amount of dissolved salts in water affects the time it takes for water to boil. In that case, your statement of the problem would be written in the following way:

'How does the amount of dissolved salts in water affect the time it takes for the water to boil?'
Note that your problem is that you do not know how the amount of dissolved salts in water affects the boiling point of the water.

Step 2: Formulating a hypothesis

A hypothesis is your guessed answer to the problem.

Your hypothesis should suggest clearly how the variables affect the system you want to investigate. This means that it should suggest the 'cause and effect' relationship. An example of good hypothesis would be: The greater the amount of salts dissolved in water, the longer it takes for the water to boil.

Exercise 1.3

1. What do you think is the importance of writing a hypothesis during scientific investigation?
2. Why would the following not be a good hypothesis? "Water with different amounts of dissolved salts boil at different times?"
3. Suppose you want to know the factors that affect the rate at which sugar dissolves in water.
Suggest:
 - a) the statement of the problem.
 - b) good hypotheses for the investigation.

Step 3: Testing the hypothesis

A hypothesis can be true or false. A test done to check if a hypothesis is true or false is called an experiment. Experiments can be carried out in the following stages:

Stage 1: Designing an experiment

To design an experiment, you take into consideration the hypothesis you made about how the variables bring about the things you observe. At this stage you carry out the following activities:

- a) identify variables;
- b) identify materials to be used;
- c) write the procedure to be followed;
- d) predict the outcome of the experiment.

I. Identifying variables

At first you must identify all the factors that can affect the results of your experiment. These factors are called **variables**. Variables can be time, mass, length, concentration, amount of light, electric current.

During an experiment, you have to change only one variable at a time. Therefore, after identifying the variables, you have to think of how you will change one variable at a time while keeping the other variables constant (unchanged). Variables are categorised as: independent, dependent and control variables.

Independent variables

An independent variable is the variable that you change deliberately during the experiment. There must be only one independent variable during each experiment being performed.

Dependent variables

This is the variable that you will be measuring during the experiment. During each experiment, there should also be only one dependent variable.

Control variables

These are the variables that are kept unchanged during the experiment so that they do not interfere with the results of the experiment.

II. Identifying materials and equipment

After identifying the variables, you have to make a list of the things you will need in order to carry out the experiment and prepare them.

III. Writing procedure

A procedure is a clear description of how the experiment will be carried out. The procedure should have step-by-step directions for conducting experiment. It must state clearly how you will change one variable while keeping the others constant. It must also explain how you will measure the amount of change. Often the procedure includes a sketch of how to set up the experiment.

IV. Predicting

Before conducting an experiment, the experimenter should make a prediction. A prediction is an expected outcome of an experiment. It states in advance the results that will be obtained from the experiment if the hypothesis is true. A prediction, most of the times, takes the form of an 'If ... then ...' statement.

Stage 2: Conducting the experiment and recording data

Usually in chemistry data are numbers. Numbers are used to measure and calculate results. Therefore, in chemistry, experiments are done in order to gather numerical data from which relationships are derived and conclusion are made. As you do experiments, record all numerical measurements made alongside all other observations. These other observations may include the following: amounts of chemicals used, how long something is and time something took.

Stage 3: Recording observations

Observations can be written descriptions of what you noticed during an experiment, or problems encountered. As you do experiments keep careful notes of everything you do, and everything that happens. Observations are valuable when drawing conclusions and useful for locating experimental errors.

Stage 4: Performing calculations

You may need to do calculations from your raw data to obtain the numbers you need to draw your conclusions. For example, suppose you weighed a container. This weight is recorded in your raw data table as 'Weight of Container'. You then added some sugar to the container and then weighed it again. This could be entered as 'Weight of Container + Sugar'. In the calculations section, you do calculations to find out how much sugar was used in the experiment. **(Weight of container + sugar) – (Weight of container) = Weight of sugar used.** Each calculated answer is entered into a table in a 'Results' section. Note that not all experiments need a calculation section.

Stage 5: Summarizing results

This can be in form of a table of processed numerical data, or graphs. It could also be a written statement of what occurred during experiment. It is from calculations using the recorded data

that tables and graphs are made. After studying the tables and the graphs we can see trends that tell us how different variables cause our observations. Based on these trends, we can draw conclusions about the system under study. These conclusions help us confirm or deny our original hypothesis.

Stage 6: Drawing conclusions

Using the trends in your experimental data, and your experimental observations, try to answer your original investigation question. Then check if your hypothesis true or false. A hypothesis is a ‘guess’, therefore, it can be true or false. Do not get worried if your hypothesis is not true.

Other things that can be mentioned in the conclusion:

- If your hypothesis is not correct, what could be the answer to your question?
- Summary of any difficulties or problems you had encountered while doing the experiment.
- Do you need to change procedure and repeat the experiment?
- What could you do differently next time?
- List other things learned as you carried out the experiment.

a) Reporting

After conducting an investigation, you are supposed to write a report so that others can know what you did and what you found. See the structure of a laboratory report in Appendix 1.

Exercise 1.4

1. Suggest the reason why you should change only one variable, while keeping all other variables constant during experimentation?
2. What is the disadvantage of not writing a procedure for an experiment?
3. What is the importance of predicting the results of an experiment?
4. a) Design a scientific investigation to find out whether pure water is less dense than an aqueous salt solution.
b) Carry out the investigation you have designed in (a) above.
c) Write a report of your investigation (see Appendix 1).
d) Present your report to the class for discussion

1.3: Purity of substances

A pure substance is a material that has constant composition and has consistent properties throughout the sample. Pure substances can be elements such as copper, aluminium, oxygen, hydrogen and graphite, or compounds like: water, pure sugar, salt (NaCl), and baking soda (NaHCO_3). The purity of a substance can be checked by measuring melting point, measuring boiling point or using chromatography.

Checking the purity of substances using melting point

When a solid is melting, its particles do not melt all at once at a single temperature. Instead, it melts over a range of temperatures as demonstrated in Figure 1.1. Therefore, experimental melting points are reported as a range such as $42^{\circ}\text{C} - 44^{\circ}\text{C}$. The melting point range of a solid can be found as follows:

1. At the start of melting point range, the sample appears to shrink away from the sides of the beaker or test tube and some liquid begins to appear at the bottom (Figure 1.2(a)). Record the temperature at which this happens.
2. As the temperature rises the liquid phase becomes clearly visible and there is slushy mixture of solid particles and liquid (Figure 1.2 (b)).
3. Finally, the entire solid has melted and a clear liquid remains (Figure 1.2(c)). The temperature at which this happens should be recorded again.
4. Then record the melting point as a range.

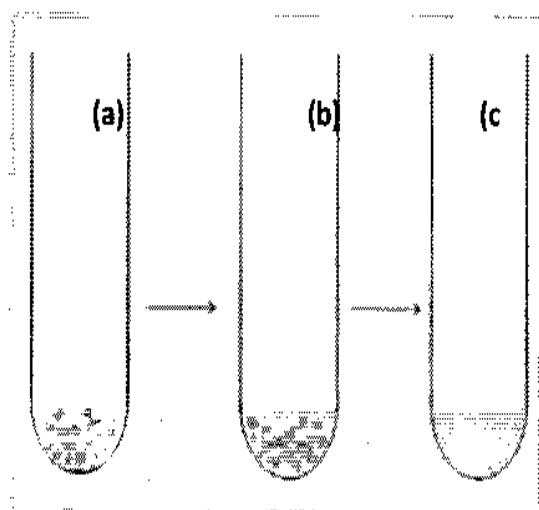


Figure 1.2: Process of melting a solid

Experiment 1.1: Determining the purity of water using melting point

Materials

- Solidified pure (distilled) water (ice)
- Solidified aqueous salt solution
- Solidified river water
- Thermometer
- Warm water in a water bath
- A test tube
- Test tube holder

Procedure

1. Crash the ice into small particles.
2. Place the solid pieces of ice into the test tube to a height of about 3cm.
3. Dip the thermometer into the test tube and place it into the warm water bath.

4. Record the temperature at which some liquid begins to appear in the test tube.
5. Keep the test tube in the warm water bath until all the solid particles have disappeared.
6. Record the temperature at which all the solid particles disappeared.
7. Repeat steps 1 to 6 with the solid salt solution.
8. Repeat steps 1 to 6 with the solid river water.

Discussion

1. What was the melting point range for the:
 - a) pure water?
 - b) salt solution?
2. Between the solidified pure water and the solidified aqueous salt solution, which one had:
 - a) wider melting point range?
 - b) lower melting point range?
3. How do impurities (substances dissolved in water) affect the melting point range of water?
4. Explain whether water from the river is pure or not.

Effect of impurities on the boiling point of a substance

At a given pressure, pure substances boil at a specific temperature. Since atmospheric pressure decreases with increase in altitude, the boiling point of a pure substance at a particular place must be determined before one can use boiling point to check the purity of liquids.

Experiment 1.2: Investigating the effect of impurities on the boiling point of substances

Materials

- three test tubes labelled A, B, C
- sodium chloride
- ethanol
- water trough
- boiling water
- three thermometers

Procedure

1. Place equal volumes of pure ethanol in all the three beakers (A, B and C).
2. Add some sodium chloride to beaker B.
3. Keep adding sodium chloride to test tube C until no more dissolves to produce a saturated solution at that temperature.
4. Put all the three beakers in trough (see Figure 1.3).
5. Insert a thermometer in each beaker. (Make sure the bulb of the thermometer does not touch the wall of the beaker)
6. Carefully, pour hot water into the trough so that the level of the water is higher than the level of the ethanol in the beakers.
7. Stir gently with the thermometer.
8. Record the temperature at which the ethanol in each beaker boils.

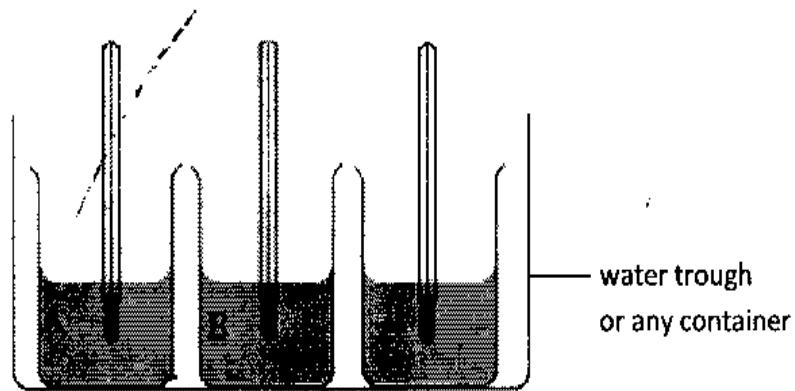


Figure 1.3: Measuring boiling point of ethanol

Discussion

1. How did the boiling points of the ethanol in the three beakers compare?
2. What is the boiling point of pure ethanol in your area?
3. How does the amount of impurities in a substance affect its boiling?
4. How can you use the boiling point of a substance to find out whether the substance is pure or not?

Experiment 1.3: Investigating purity of drinking water

Design an experiment to investigate whether the water you drink in your area is pure or not using boiling point.

Discussion

1. What is the boiling point of pure water in your area?
2. Why is it necessary to determine the boiling point of a pure substance in a particular area before using boiling point to check the purity of a substance?
3. Is the water you drink pure or not? Explain your answer.
4. If it is not pure, how can you obtain pure water from it?

Using chromatography to check the purity of substances

In Form 1, we learnt how to separate coloured components of a mixture such as ink using chromatography. Each of the coloured components written on the chromatography paper is known as a **chromatogram** (Figure 1.4). If a substance is pure, the chromatography produces just one chromatogram. If the substance is not pure, different chromatograms are produced at different distances from the base line.

Experiment 1.4: Investigating if different coloured substances are pure or not.

Materials

- Chromatography (filter) paper
- Beakers (plastic cups)
- Ethanol
- Coloured substances such as ink, dyes and juices from different coloured flowers and leaves of plants.

Procedure

1. Carry out a chromatography of each coloured substance as you did in form one.
2. Record the number of chromatograms produced by each coloured substance.

Discussion

1. Which substance had:
 - a. the highest number of chromatograms;
 - b. the least number of chromatograms?
2. Which ones of the coloured substances tested, if any:
 - a. are pure;
 - b. have the greatest number of coloured components mixed together?

Relative flow (R.f.) values

A particular component will travel the same distance up the chromatography paper by a specific solvent given that the following experimental conditions are kept constant:

- temperature.
- type of chromatography paper.
- solvent concentration.

The distance travelled up the paper by a component divided by the distance travelled up the paper by the solvent is called relative flow value. It is also called retention factor.

$$\text{By definition, relative flow value} = \frac{\text{distance travelled by a compound}}{\text{distance travelled by solvent}}$$

Both distances are measured from the point where the sample was initially spotted on the chromatography paper (Figure 1.4). Unknown samples can be identified by comparing their r.f values with r.f values of known samples.

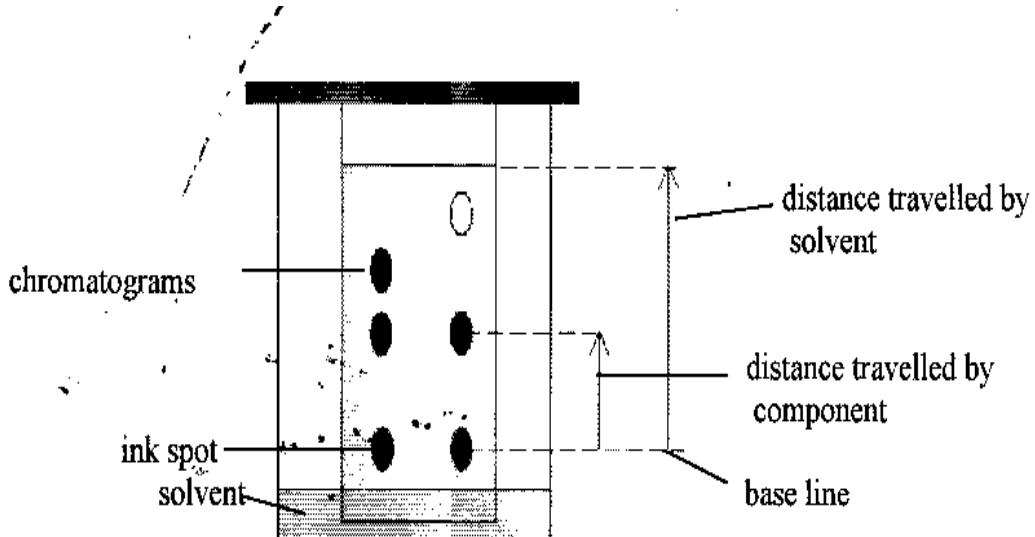


Figure 1.4: Measuring distances travelled by solvent and coloured components

Exercise 1.4

1. If the distance moved by the solvent is 6cm, what is the r.f factor of the component which moved the distance of 2cm from the initial spot of the ink on the chromatography paper (base line) in Figure 1.4?

2. What will be the r.f factor of a component of a substance if:
 - a) it remains on the initial spot, without moving, on the chromatography paper?
 - b) it travels the same distance as the solvent?

Experiment 1.5: Identifying ink with same components using r.f values

Materials

- Four black pens of different brands (labelled 1, 2, 3 and 4)
- Ethanol
- Chromatography paper
- Beaker

Procedure

1. Carry out chromatography with pen 1.
2. Calculate r.f values of its chromatograms.
3. Repeat steps 1 and 2 with the other pens.

Discussion

1. Which other pens have same components as pen 1?
2. Which brand of pens has:
 - a) the greatest number of components?
 - b) least number of components?

1.4: Simple tests for water, ions and gases

Water dissolves a lot of different types of compounds. Different kinds of gases are also produced from different kinds of chemical reactions. No one can tell what substances are dissolved in water or what gases are produced from chemical reactions because usually, they all look colourless. To identify what substances are dissolved in water, or what types of gases are produced during chemical reactions one has to carry out tests.

Tests for ions

Particular ions present in water can be detected by using precipitation reactions. A precipitation reaction is a reaction in which ions in a solution react with each other to form a new substance that is insoluble. The insoluble substance that forms in a solution during chemical reaction is called a precipitate.

A. Testing for cations

Some common cations that can be tested in water include: Pb^{2+} , Al^{3+} , Zn^{2+} , Ca^{2+} , Fe^{2+} , Fe^{3+} , and Cu^{2+} . These cations can be identified using precipitation reactions with aqueous solution of sodium hydroxide or an aqueous solution of ammonia.

Testing for cations using sodium hydroxide/ aqueous ammonia

The following steps can be followed when testing for cations:

1. Place about 2cm^3 of the test liquid into a test tube.
2. Add a few drops of aqueous sodium hydroxide/aqueous ammonia.
3. Shake the test tube to mix.

Conclusion

- a) If a coloured precipitate is formed then stop and find out what the cation is using Table 1.1 or Table 1.2.
- b) If a white precipitate is formed then continue to add more aqueous sodium hydroxide/ aqueous ammonia to it to find out whether the precipitate will dissolve or not. Identify it using Tables 1.1 or Table 1.2.

Table 1.1: Test for cations using aqueous sodium hydroxide

Cation	Results	Further results
Pb^{2+} , Al^{3+} , Zn^{2+}	White precipitate forms	Precipitate dissolves as more NaOH is added to the solution
Ca^{2+}	White precipitate forms	Precipitate does not dissolve as more NaOH is added to the solution
Fe^{2+}	Green precipitate forms	Not applicable
Fe^{3+}	Red-brown precipitate forms	Not applicable
Cu^{2+}	Light-green precipitate forms	Not applicable

Table 1.2: Test for cations using ammonia

Cation	Results	Further results
Pb ²⁺ , Al ³⁺	White precipitate forms	Precipitate does not dissolve as more NH ₃ is added
Zn ²⁺	White precipitate forms	Precipitate dissolves as more NH ₃ is added
Ca ²⁺	No visible reaction occurs	Not applicable
Fe ²⁺	Green precipitate forms	Not applicable
Fe ³⁺	Red-brown precipitate forms	Not applicable
Cu ²⁺	Light-green precipitate forms	Not applicable

Experiment 1.6

You are provided with water samples containing the following cations: Al³⁺, Zn²⁺, Ca²⁺, Fe²⁺, Fe³⁺ and Cu²⁺, in bottles labelled 1, 2, 3, 4, 5 and 6, but not in any order.

- i. Design a scientific investigation that will be done to identify the ions in each bottle using ammonia test.
- ii. Carry out the investigation you have designed in (i) above.
- iii. Write a laboratory report.

B. Testing for anions

Anions are negatively charged ions. Just as with cations, anions can also be detected by precipitation reactions.

Testing for sulphate ions (SO₄²⁻)

If barium chloride is added to a sample of water containing sulphate ions, barium sulphate is formed. Barium sulphate is insoluble in water, and will be seen as a white precipitate. The equation for the reaction is as follows: BaCl₂(aq) + SO₄²⁻(aq) → 2Cl⁻(aq) + BaSO₄(s).

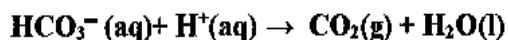
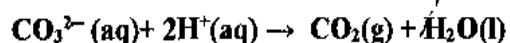
Testing for halide ions using silver nitrate

Halides are ions of group VII elements. They can be detected in solutions by using silver nitrate. If silver nitrate is added to a sample of water containing halide ions, silver halide is precipitated because silver halides are insoluble in water. The precipitates are coloured as follows:

- Silver chloride precipitate is white.
- Silver bromide precipitate is cream.
- Silver iodide precipitate is yellow.

Testing for carbonates (CO_3^{2-}) or hydrogen carbonate (HCO_3^-)

Carbonates or hydrogen carbonates (bicarbonates) produce carbon dioxide when they react with acids:



The presence of carbonate or bicarbonate ions in water can be detected using lime water. Lime water is an aqueous solution of calcium hydroxide ($\text{Ca}(\text{OH})_2$). Lime water can be prepared as follows:

- a) Put 1.5g (about 1 teaspoon) of calcium hydroxide powder in a 1litre bottle.
- b) Fill the bottle with distilled water or tap water.
- c) Shake vigorously for 1-2 minutes.
- d) Let the bottle stand for about 24 hours.
- e) Being careful not to stir up the sediment, pour the clear solution off the top of the bottle through a clean filter paper.
- f) This solution is the lime water. Store it.

If carbon dioxide gas comes into contact with calcium hydroxide in the lime water, calcium carbonate and water are formed as follows: $\text{Ca}(\text{OH})_2(\text{aq}) + \text{CO}_2(\text{g}) \rightarrow \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$. Calcium carbonate is insoluble in water and it appears as a milky or a cloudy precipitate in the water. To determine if a solution contains carbonate or bicarbonate ions:

1. Add an acid, such as dilute hydrochloric acid or dilute sulphuric acid, to a sample of the solution to be tested.
2. If a gas is produced, collect it using a test pipette (Figure 1.5).
3. Pass the gas through the lime water.

Conclusion

If the lime water turns milky (becomes cloudy) then the gas is carbon dioxide, and the solution contains carbonates or bicarbonates.

Testing for nitrates (NO_3^-)

The presence of nitrates in a solution can be detected by using either of the following tests:

A. Using litmus paper

To test if a liquid contains nitrates using litmus paper:

1. Boil the sample with sodium hydroxide solution and aluminium foil.
2. Hold a blue litmus paper at the mouth of the test tube.

Conclusion

If the gas turns the blue litmus paper red, then the gas is ammonia (NH_3), and the sample contains nitrates.

B. Using the brown ring test

To test if a liquid contains nitrates using a brown ring test:

1. Add iron (II) sulphate to the sample.
2. Add concentrated sulphuric acid to the sample.

Conclusion

If a brown ring forms where the liquids meet then the sample contains nitrate ions.

Exercise 1.6

Suppose you are given water samples containing the following anions in bottles labelled A, B, C, D and E, not in any order: NO_3^- , Cl^- , I^- , Br^- and SO_4^{2-} . Construct a flow diagram that you can follow when carrying out tests to identify the anions present in each bottle.

Tests for gases

The presence of different gases can be detected by carrying out various tests.

Testing for sulphur dioxide

SO_2 changes the colour of acidified potassium dichromate (IV) solution from orange to green. To test for sulphur dioxide:

1. Collect the suspected gas with a teat pipette.
2. Bubble the gas into an acidified potassium dichromate (IV) solution.

Conclusion

If the solution turns from orange to green, then the gas is sulphur dioxide.

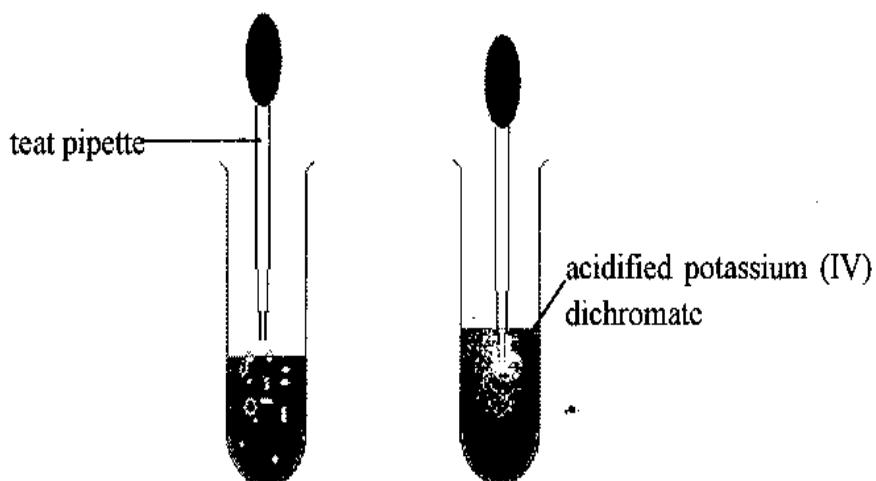


Figure 1.5: Collecting gas with teat pipette

Testing for carbon dioxide

To test if a given gas is a carbon dioxide follow the following procedure:



1. Collect the suspected gas with a teat pipette.
2. Bubble it through lime water.

Conclusion

If the lime water turns milky then the gas is carbon dioxide.

Caution! CaCO_3 precipitate dissolves in excess carbon dioxide to form calcium hydrogen carbonate ($\text{Ca}(\text{HCO}_3)_2$) according to the following equation;
 $\text{CaCO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{HCO}_3)_2(\text{aq})$. Therefore, do not add too much CO_2 .

Testing for oxygen (O_2)

1. Take a glowing wooden splint and insert it into a test tube containing a gas sample.

2. Conclusion

If the wooden splint re-lights, then the gas is oxygen.

Testing for chlorine

Caution! Chlorine gas a poisonous gas with a sharp, chocking smell. Any activity involving handling chlorine gas should be done in a fume hood while wearing a gas mask.

A. Using litmus paper

1. Wet a blue litmus paper with water.
2. Insert the damp blue litmus paper into the suspected gas.

3. Conclusion

If the blue litmus turns red and then white, then the gas is chlorine.

B. Using silver nitrate

1. Put a drop of aqueous silver nitrate on the end of a glass rod.
2. Dip it into the suspected gas.

3. Conclusion

If a white precipitate forms at the end of the rod then the gas is chlorine.

Testing for hydrogen gas

A lighted wooden splint makes a popping sound in a test tube of hydrogen. The reaction that takes place is as follows: $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$. This reaction gives out a lot of heat. To test for the hydrogen gas:

1. Put a glowing splint into a test tube containing the sample gas.

2. Conclusion

A popping sound indicates the presence of hydrogen gas.

Testing for ammonia gas

Ammonia makes damp red litmus paper turn blue. To test for ammonia gas:

1. Soak a strip of red litmus paper in distilled water.
2. Bring the damp red litmus paper on the mouth of test tube of a gas suspected to be ammonia.

3. Conclusion

If the red litmus paper turns blue then the gas is ammonia.

Exercise 1.7

Suppose you are given the following gases in gas bottles labelled 1, 2, 3, 4, 5 and 6, not in any order: ammonia (NH_3), CO_2 , H_2 , Cl_2 , O_2 and SO_2 . Construct a flow diagram that you can use to identify the gas contained in each gas bottle.

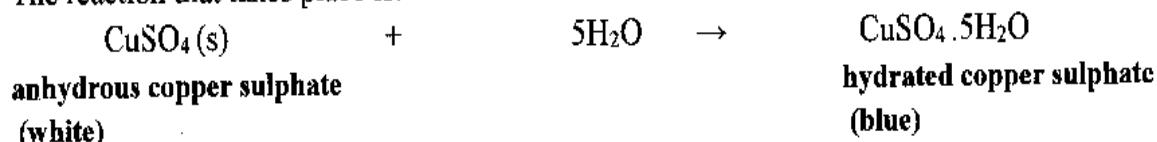
Tests for water

All liquids do not necessarily contain water. Some solids also contain water which is not observable by naked eyes. Tests should, therefore, be used to identify water. Most water tests use anhydrous copper sulphate and anhydrous cobalt chloride. The prefix 'an' means the absence of and 'hydrated' means water. Therefore, the term 'anhydrous' means without water. It is synonymous with dehydrated.

Testing for water using anhydrous copper sulphate

Anhydrous copper sulphate is a white powder. Its formula is CuSO_4 . When anhydrous copper sulphate comes in contact with water it turns from white to blue.

The reaction that takes place is:



A. Testing for water in a liquid using anhydrous copper sulphate

To test if a liquid contains water:

1. Add a few drops of the test liquid to a few crystals of anhydrous copper sulphate.

2. Conclusion

If the anhydrous copper sulphate turns blue then the liquid is water or it contains water.

Experiment 1.6: Testing for the presence of water in liquids using anhydrous copper sulphate

Carry out an experiment to test if the following liquids contain water; cooking oil, lemon juice, milk, paraffin, vinegar.

Discussion

Which liquids have water?

B. Testing for water in a solid using anhydrous copper sulphate

To test if a solid such as tomato, potato, apple, contains water, add a few crystals of the anhydrous copper sulphate to the solid. If the anhydrous copper sulphate turns blue then the solid contains water.

Experiment 1.7: Testing for the presence of water in solids using anhydrous copper sulphate

Carry out an experiment to test if the following contain water; potato, cassava, milk, avocado pear, mango, pawpaw using anhydrous copper sulphate.

Discussion

Which solids have water?

Experiment 1.8: Testing for water using anhydrous cobalt chloride

Materials

- Anhydrous cobalt chloride paper
- Water

Procedure

1. Soak the anhydrous cobalt chloride paper in water
2. Record the colour changes

Discussion

- 1) What is the colour of anhydrous cobalt chloride paper before soaking it into water?
- 2) Describe the colour changes that took place on the cobalt chloride paper when it was soaked in water?
- 3) Explain how you can use the anhydrous cobalt chloride to test for water in different substances.

Topic review

1. Define the term 'chemical waste'.
2. Name chemical wastes that can go into the laboratory drain.
3. What would be the problem of disposing of acids of pH 2 and pH 1 into laboratory drain?
4. What can you do with a base of pH 14 before disposing of it into the laboratory drain?
5. How does the melting point range of a pure substance differ from that of an impure substance?
6. Suppose you want to find the effect of impurities on the boiling point of water.
 - a) Outline all the factors that would affect the experiment.
 - b) Mention:
 - i. independent variable.
 - ii. dependent variable and
 - iii. control variable.
 - c) What would be your hypothesis?
 - d) What would be your prediction?
7. The melting point of sodium chloride is 801°C . The melting points of three samples of sodium chloride were determined. The results were as follows:

sample 1	$800\text{-}801^{\circ}\text{C}$
sample 2	$799\text{-}802^{\circ}\text{C}$
sample 3	801°C

Which sample(s) of sodium chloride is/are pure? Explain your answer.
8. Suggest the importance of experiments in chemistry.

Core Element 2: Inorganic compounds

Topic 2: Nitrogen, sulphur and phosphorus

Topic 2: Nitrogen, sulphur and phosphorus

Different kinds of chemical elements are used different ways in our everyday life. Some elements are also used by plants in different ways such as for growth, seed and fruit production and chlorophyll formation. In this topic we will look at the properties and uses of the following elements: nitrogen, sulphur and phosphorus.

By the end of this topic you must 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 uses of sulphur and its compounds.
- 6) describe sources and properties of phosphorus.
- 7) describe uses of phosphorus and its compounds.

2.1: Properties and uses of nitrogen and its compounds

Nitrogen is an element that is found in period 2 and group V in the periodic table. It exists as a diatomic molecule N_2 .

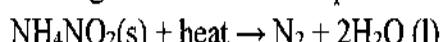
Sources of nitrogen

Nitrogen occurs naturally as uncombined gas (N_2) in air. About 78% of the volume of air is nitrogen. Nitrogen can be obtained from air by fractional distillation. In the laboratory, nitrogen can be prepared by following these steps:

Step 1: heating a mixture of ammonium chloride and sodium nitrite as follows:



Step 2: heating ammonium nitrite produced in step 1 to decompose it:



Exercise 2.1

In Form 2 we discussed how to separate air into its components. Describe, in details, how nitrogen is obtained from air.

Properties of nitrogen

a) Physical properties

The physical properties of nitrogen are as follows:

1. It is colourless gas, liquid or solid without smell.
2. It is a gas at room temperature.
3. It is a non-metal.
4. It does not burn.
5. It is neither acidic nor basic.

b) Chemical properties

1. Inert character of nitrogen

Nitrogen is inert (or does not react) under normal conditions. This is because the nitrogen atoms in the nitrogen molecule, N_2 , are joined together by strong triple covalent bonds as follows $\text{N}\equiv\text{N}$. These bonds cannot be broken easily.

2. Reactivity of nitrogen under specific conditions

Under specific conditions nitrogen reacts. For example:

- a. At very high temperatures, nitrogen reacts with some metals such as calcium and magnesium to form nitrides as follows:
 - $3\text{Mg}(\text{s}) + \text{N}_2(\text{g}) \rightarrow \text{Mg}_3\text{N}_2(\text{s})$
 - $3\text{Ca}(\text{s}) + \text{N}_2(\text{s}) \rightarrow \text{Ca}_3\text{N}_2(\text{s})$
- b. At very high temperatures and pressures such as those found in combustion reactions in power plants or internal combustion engines, nitrogen reacts with oxygen to form nitrogen monoxide, NO , and nitrogen dioxide, NO_2 .

Exercise 2.2

1. Air is composed of 78% nitrogen and 20% oxygen. Explain why nitrogen does not react with oxygen under normal temperature and pressure.
2. Explain why nitrogen is able to react at very high temperature and/or pressure.

Uses of nitrogen, N_2

Nitrogen has a number of applications in our everyday life. Some of the uses of nitrogen are as follows:

1. It is used to preserve freshness of foods. When the foods are packaged in nitrogen rancidity and other forms of oxidative damage of the food is delayed and reduced.
2. It is used in incandescent light bulbs as an inexpensive alternative to argon to prevent the tungsten filament from burning up.
3. It is used in some aircraft fuel systems to reduce fire hazards.
4. It is used to cover liquid explosives as a safety measure. For example, it is used in military aircraft fuel systems to reduce fire hazards.
5. It is used as a refrigerant to preserve blood and reproductive cells.

Exercise 2.3

For each property of nitrogen in column I, write the letter representing the application of nitrogen in column II that is related to that property. Each application can be written once, more than once or not at all.

Column I

1. _____ does not burn
2. _____ is inert
3. _____ neither acidic or basic
4. _____ is colourless gas
5. _____ is gas at room temperature

Column II

- A. to preserve freshness of foods
- B. to reduce fire hazards in aircraft fuel systems
- C. to cover explosive liquids
- D. to preserve reproductive cells
- E. to prevent tungsten from burning up in incandescent tubes.

Nitrogen compounds

Nitrogen is used to produce compounds that are useful to our life such as ammonia (NH_3) and nitric acid (HNO_3).

1. Preparation of ammonia

Ammonia is a compound of nitrogen and hydrogen. It is a colourless gas with a chocking smell, similar to the smell of urine.

Ammonia is produced by a process called the Haber process from the reaction between nitrogen gas and hydrogen gas; $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$. Figure 2.2 summarises the Haber process. The reaction requires very high pressure of about 200 atmospheres and it produces a lot of heat.

The stages in the Haber process

Stage 1: Obtaining hydrogen from natural gas

Hydrogen is obtained by reacting natural gas (methane) with steam (H_2O) as represented by the following equation: $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$ (see Figure 2.2). More hydrogen gas is produced when CO reacts with steam: $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2(\text{g})$.

Stage 2: Obtaining nitrogen from the air

Part of the hydrogen gas produced in Stage 1 is burned in air as follows:

$\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g})$. This reaction removes oxygen from the air, leaving behind nitrogen (Figure 2.2).

Stage 3: Reacting hydrogen and nitrogen to produce ammonia

Finally, the hydrogen gas is reacted with nitrogen gas to produce anhydrous liquid ammonia.
 $3\text{H}_2(\text{g}) + \text{N}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{l})$.

For this reaction to take place, it requires the following conditions:

- a high temperature of 450°C .
- a high pressure of 200 atm.
- iron catalyst.

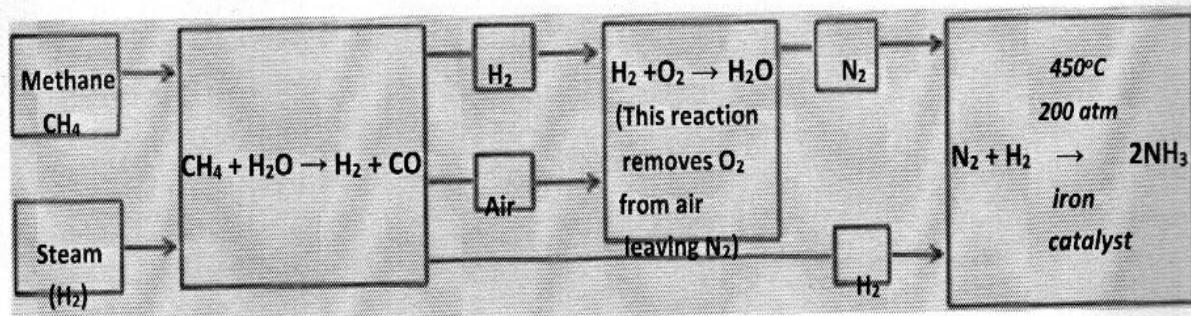
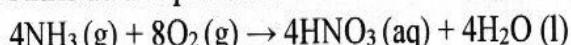


Figure 2.1: Diagrammatic illustration of the Haber process

2. Preparation of nitric acid (HNO₃)

Nitric acid is produced from the reaction of ammonia with oxygen:



Commercial production of nitric acid

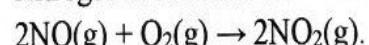
Commercially, nitric acid is produced by a process known as the Ostwald process through the following stages:

Stage 1: Production of nitrogen monoxide

Anhydrous ammonia is reacted with oxygen in the presence of platinum or rhodium catalyst at a high temperature of 227°C and a pressure of 9 atmospheres to produce nitrogen monoxide (NO); $4\text{NH}_3(\text{g}) + 5\text{O}_2 \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$. The reaction produces a lot of heat.

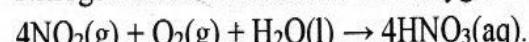
Stage 2: Production of nitrogen dioxide

Nitrogen monoxide is reacted with oxygen in air to produce nitrogen dioxide as follows:



Stage 3: Production of nitric acid

Nitrogen dioxide is reacted with oxygen in air and water to produce nitric acid.

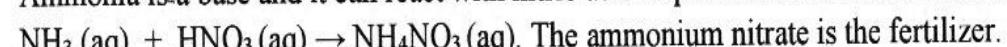


Laboratory preparation of nitric acid

In the laboratory, nitric acid can be produced by reacting equal masses of nitrate salt such as sodium nitrate with concentrated (96%) sulphuric acid (H₂SO₄) and then distilling the mixture at the boiling point of nitric acid which is 83°C: $2\text{NaNO}_3 + \text{H}_2\text{SO}_4 \rightarrow 2\text{HNO}_3 + \text{Na}_2\text{SO}_4$.

Uses of ammonia and nitric acid in the manufacture of ammonium nitrate fertilizers

Ammonia is a base and it can react with nitric acid to produce ammonium nitrate as follows:



Exercise 2.4

1. What name is given to the reaction between a base and an acid?
2. What is the importance of ammonium nitrate fertilizer to our life?

2.2: Properties and uses of sulphur

Sulphur is an element found in group VI and period 3 of the periodic table. It exists as S_8 molecules. The molecules are crown-shaped as shown in Figure 2.3.

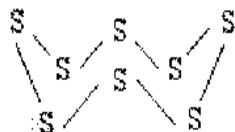


Figure 2.2: A sulphur molecule

Sources of sulphur

In nature, sulphur occurs as a pure element, as sulphide or as sulphate minerals. The major sources of sulphur include the following:

1. Volcanic regions such as Japan, Hawaiian Islands, Indonesia and Kamchatka in Russia.
2. Hydrogen sulphide of the natural gas.
3. Sulphide minerals such as pyrite which is iron sulphide (FeS_2) and galena which is lead sulphide (PbS).

Physical properties of sulphur

The physical properties of sulphur are as follows:

- 1) It is pale yellow in colour.
- 2) It is brittle.
- 3) It is insoluble in water.
- 4) It is a non-metal.
- 5) It does not conduct electricity.
- 6) Its melting point range is $110^{\circ}C$ to $119^{\circ}C$.
- 7) It boils at $444^{\circ}C$.
- 8) Sulphur has no smell.
- 9) Sulphur occurs in different allotropes: as monoclinic sulphur and rhombic sulphur (see Figure 2.5). Below the temperature of $94.5^{\circ}C$ sulphur exists in the form of rhombic sulphur (Figure 2.5 (a)) and at temperatures above $94.5^{\circ}C$, it exists as monoclinic sulphur (Figure 2.5(b)).

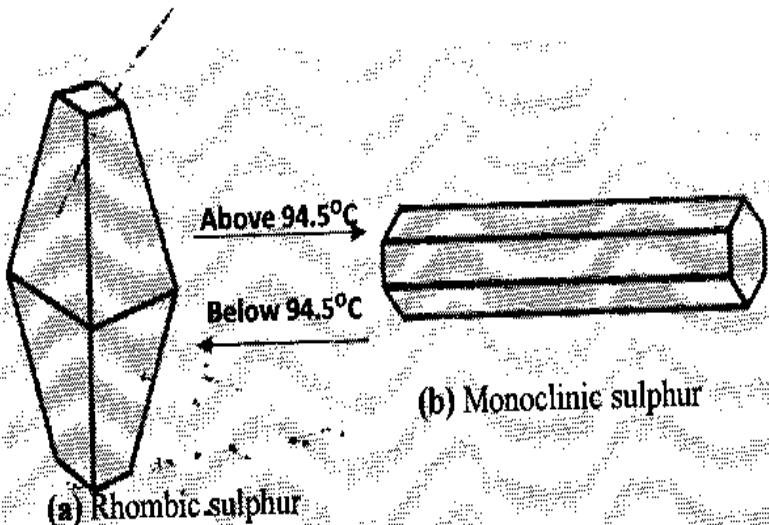


Figure 2.3: Allotropes of sulphur

Exercise 2.5

1. What do the following terms mean?
 - a. allotropes?
 - b. brittle?
2. In what form does sulphur exist at room temperature?
3. Suggest the reason why sulphur does not conduct electricity?
4. At what temperatures would sulphur exist as liquid?

Chemical properties of sulphur

1. Sulphur reacts with metals to form sulphides. For example, sulphur reacts with iron to form iron sulphide: $8\text{Fe(s)} + \text{S}_8\text{(s)} \rightarrow 8\text{FeS(s)}$.
2. Sulphur burns in air to produce sulphur dioxide: $\text{S}_8\text{(s)} + 8\text{O}_2\text{(g)} \rightarrow 8\text{SO}_2\text{(g)}$.
3. Sulphur reacts with bases to produce sulphides and thiosulphates. For example, sulphur reacts with hot potassium hydroxide, KOH, to form potassium sulphide, potassium thiosulphate and water as follows; $\text{S}_8 + 6\text{KOH} \rightarrow 2\text{K}_2\text{S}_3 + \text{K}_2\text{S}_2\text{O}_3 + 3\text{H}_2\text{O}$.

Uses of sulphur

Sulphur has a number of important uses in our life which include the following:

1. to vulcanise rubber. Vulcanisation of sulphur is the process of adding sulphur to rubber in order to toughen it so that it can be used to make such things as rubber tyres.
2. to manufacture products such as matches, pesticides, drugs, dyes, and gun powder.
3. to make sulphuric acid.

Exercise 2.6

Discuss the importance of each use of sulphur stated above in our daily life?

Preparation of sulphuric acid (H_2SO_4)

Sulphuric acid is produced through the process known as the contact process. This process has the following stages:

Stage 1: Production of sulphur dioxide

Sulphur is burnt in oxygen to produce sulphur dioxide (SO_2).

Stage 2: Production of sulphur trioxide

Sulphur dioxide is reacted with oxygen to produce sulphur trioxide (SO_3).

Stage 3: Production of sulphuric acid from sulphur trioxide

After producing sulphur trioxide in Stage 2, there are two routes that can be followed to produce the sulphuric acid:

Route 1: Mixing sulphur trioxide with water

The sulphur trioxide can be mixed with water to produce sulphuric acid, H_2SO_4 . However, this route has the following disadvantages:

- the reaction is very violent so that it is dangerous.
- the acid produced is very thick and too concentrated.

Route 2: Mixing sulphur trioxide with concentrated sulphur dioxide

To produce a good sulphuric acid with required concentration, the following procedure can be followed:

- Sulphur trioxide is mixed with concentrated sulphuric acid to produce a substance known as oleum ($H_2S_2O_7$).
- The oleum is then reacted with water to produce sulphuric acid of the required concentration.

Exercise 2.7

- Write balanced equation for each reaction that takes place in the contact process.

Use of sulphuric acid in the manufacture of inorganic fertilizers

Sulphuric acid is used to manufacture inorganic fertilizers as follows:

1. Ammonium sulphate fertilizer, $(NH_4)_2SO_4$

Sulphuric acid is reacted with ammonia to produce ammonium sulphate.

2. Calcium phosphate fertilizer, $Ca(H_2PO_4)_2$

Calcium phosphate is produced when a phosphate rock ($Ca_3(PO_4)_2$) is reacted with sulphuric acid.

Exercise 2.8

- Write balanced equations for the production of each of the following fertilizers.



2.3: Properties and uses of phosphorus

Phosphorus is an element which is found in group V and period 3 in the periodic table. It exists as P₄ molecules which are tetrahedron shaped as shown in Figure 2.6.

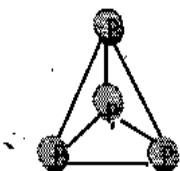


Figure 2.4: A phosphorus molecule

Sources of phosphorus

The main source of phosphorus is phosphate rock whose chemical formula is Ca₃(PO₄)₂. Large deposits of phosphate rock are found in many countries such as the USA, Egypt and South Africa. Other sources include: bone meal, soya husks, and fish bone meal.

Physical properties of phosphorus

- a. It is a non-metal.
- b. It is insoluble in water. But it dissolves in carbon disulphide (CS₂).
- c. Phosphorus melts at 44.1°C and boils at 280°C.
- d. Phosphorus occurs in different allotropes as follows:
 - i. White phosphorus
 - ii. Red phosphorus
 - iii. Violet phosphorus
 - iv. Black phosphorus

Chemical properties of phosphorus

1. White phosphorus spontaneously ignites in air at temperatures slightly above the room temperature to form tetra-phosphorus decaoxide. However, red phosphorus is not spontaneously flammable.
2. White phosphorus glows in the dark when exposed to damp air. This process is called chemiluminescence.
3. White phosphorus, P₄, reacts vigorously with all the halogens at room temperature to form phosphorus trihalides.

Exercise 2.9

1. What is the state of phosphorus at:
 - a) room temperature (25°C)?
 - b) 60°C ?
2. What would be the danger of exposing white phosphorus to air at 30°C ?
3. Write the balanced chemical equation for the following reactions:
 - a) reaction of oxygen with phosphorus.
 - b) reaction of phosphorus with chlorine.

Uses of phosphorus

Phosphorus has a number of uses in our everyday life. Some of the uses are as follows:

a. In toothpaste

Phosphorus is used in toothpastes because it whitens and strengthens teeth.

b. In baking soda

Phosphorus is used to make baking soda.

c. In detergents

Phosphorus is used, in the form of phosphates, as automatic dish detergents and laundry detergents to help soften water and remove soil, oil and grease.

d. In safety matches

Red phosphorus is mixed with powdered glass and is stuck on the sides of boxes of safety matches on which the matches are struck to light them.

Exercise 2.10

1. Discuss the importance of each use of phosphorus mentioned above in our daily life.
2. Get a tube of toothpaste such as Colgate and a packet of baking soda. Write down the compounds of phosphorus used in:
 - a) toothpaste.
 - b) baking soda.

Phosphoric acid, H_3PO_4

One of the compounds of phosphorus is phosphoric acid. Phosphoric acid is produced either by wet process or thermal process:

Wet process

In wet process, phosphoric acid is produced when the phosphate rock ($\text{Ca}_3(\text{PO}_4)_2$) is reacted with concentrated sulphuric acid (93%). The equation for the reaction is as follows:

$\text{Ca}_3(\text{PO}_4)_2(s) + 3\text{H}_2\text{SO}_4(l) \rightarrow 2\text{H}_3\text{PO}_4(l) + \text{CaSO}_4(s)$. The phosphoric acid produced through wet process is less pure. To produce purer phosphoric acid thermal process is used.

Thermal process

In this process phosphoric acid is produced by burning white phosphorus in air at about 1527°C to 2727°C to produce phosphorus pentoxide: $\text{P}_4(\text{l}) + 5\text{O}_2 \rightarrow 2\text{P}_2\text{O}_5(\text{g})$.

Phosphorus pentoxide is then reacted with water (hydrated) to produce the phosphoric acid: $\text{P}_2\text{O}_5(\text{g}) + 3\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{H}_3\text{PO}_4(\text{l})$

Uses of phosphoric acid

Phosphoric acid is used to manufacture inorganic fertilizers and food additives.

a. Manufacturing inorganic fertilizers

- Phosphoric acid is reacted with anhydrous ammonia to produce ammonium hydrogen phosphate fertilizer, $(\text{NH}_4)_2\text{HPO}_4$; $\text{H}_3\text{PO}_4 + 2\text{NH}_3 \rightarrow (\text{NH}_4)_2\text{HPO}_4$.
- When phosphoric acid is reacted with the phosphate rock, calcium dihydrogen phosphate fertilizer is produced: $\text{Ca}_3(\text{PO}_4)_2 + 4\text{H}_3\text{PO}_4 \rightarrow 3\text{Ca}(\text{H}_2\text{PO}_4)_2$.

b. As food additive

Phosphoric acid is used to acidify foods and soft drinks such as Coca-Cola to enhance flavour and preserve freshness by slowing the growth of moulds and bacteria which otherwise multiply in the sugary solutions.

Exercise 2.11

1. Discuss the importance of phosphoric acid in our life.
2. Read the ingredients lists on cans or bottles of different non-alcoholic beverages and then:
 - a) Write down the names of beverages that have phosphoric acid.
 - b) Describe the tests of the soft beverages that have phosphoric acid.

Topic review

1. What is the major source of nitrogen?
2. Why is nitrogen inert under normal conditions?
3. Mention two compounds of nitrogen that are used to make inorganic fertilizer.
4. Mention two allotropes of sulphur.
5. In what form will sulphur exist at 97°C?
6. Describe the Haber process stage by stage.
7. In what state will phosphorus exist at 50°C?
8. Mention two uses of phosphorus.
9. Describe the two processes that are used to make phosphoric acid.
10. Name phosphorus compound that is used in baking soda.
11. Name any soft drink that contains phosphoric acid.
12. For what reasons is phosphoric acid added to food and soft drinks?
13. Mention any two types of inorganic fertilizers that are made from phosphoric acid.
14. White phosphorus is usually kept under oil. Can you suggest the reason why this is so?

Core Element 3: Chemical composition of matter

Topic 3: Chemical bonding and properties of matter

Topic 3: Chemical bonding and properties of matter

In Form 2, we looked at different types of chemical bonds. We defined a chemical bond as the force of attraction that holds together particles of matter. The forces of attraction that hold the particles in different substances determine physical properties of the substances. In this topic, we will look at the physical properties of different types of substances.

By the end of this topic you must be able to:

- 1 describe the properties of ionic and covalent compounds.
- 2 explain the structural differences between ionic and covalent compounds.
- 3 differentiate between polar and non-polar covalent bonds
- 4 describe the effects of intermolecular forces on the properties of covalent compounds.
- 5 define the term allotropy.
- 6 relate the properties of allotropes of carbon to their uses.
- 7 explain similarities between diamond and silicon dioxide
- 8 describe the uses of metals in relation to their properties.
- 9 explain physical properties of alloys.
- 10 state some of the uses of alloys.

3.1: Properties of ionic and covalent compounds

Ionic compounds and covalent compounds have some physical properties which can be used to distinguish them from each other. Some of these physical properties include: solubility, melting and boiling points, volatility and electrical conductivity.

Exercise 3.1

In Form 2, we learnt about ionic and covalent compounds.

1. What do the following terms mean:
 - a) ionic compound
 - b) covalent compound
2. Give one example and, for each example given, draw a dot and cross diagram for each of the following:
 - a) ionic compound
 - b) covalent compound
3. What is the structural difference between ionic and covalent compounds?

Experiment 3.1: Investigating the solubility of ionic and covalent compounds in water

Materials

- Sugar
- NaCl
- candle wax
- KI
- water
- 3 beakers
- a stirring rod.

Procedure

1. Pour some water into one of the three beakers.
2. Place some sugar into the beaker with water.
3. Stir the mixture.
4. Observe and record what happens to solid
5. Repeat steps 1 to 4 above with NaCl, candle wax and KI.

Discussion

1. Which of the substances tested are:
 - a) ionic compounds
 - b) covalent compounds
2. Which substances dissolved in water?
3. What is your conclusion about the solubility of ionic and covalent compounds in water?

Experiment 3.2: Investigating electrical conductivity of ionic and covalent compounds in solid and aqueous solution states

Materials

- a. Solid sugar
- b. Solid salt
- c. water
- d. Two cells in a cell holder
- e. A 100 ml beaker
- f. An ammeter
- g. Two carbon electrodes
- h. Four connecting wires
- i. A switch

Procedure

1. Arrange the apparatus as shown in Figure 3.1.
2. Close the switch.
3. Record the ammeter reading.
4. Open the switch
5. Remove the electrodes from the beaker
6. Add some water until all the sugar has dissolved.
7. Replace electrodes.
8. Repeat steps 2 and 3.
9. Repeat steps 1 to 8 with the salt



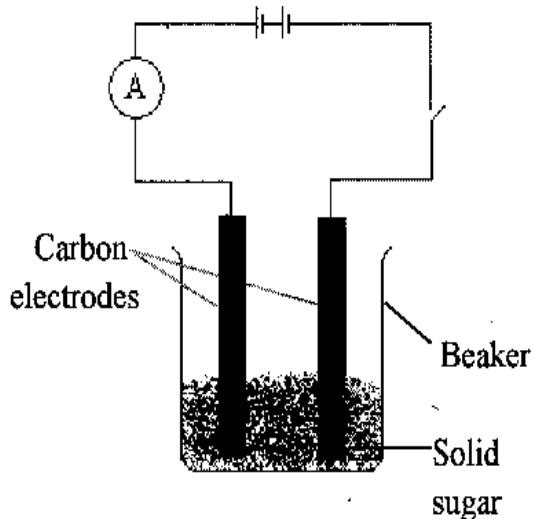


Figure 3.1: Set up of apparatus for testing electrical conductivity of ionic and covalent compounds

Discussion

1. Which one the following compounds: sugar or salt is:
 - a) an ionic compound
 - b) a covalent compound
2. Which of the following substances and states conducted electricity:
 - a) solid sugar
 - b) aqueous solution of sugar
 - c) solid salt
 - d) aqueous solution of salt
3. Explain your observations.
4. What is your conclusion about the electrical conductivities of ionic and covalent compounds?
5. Suggest whether the following conduct electricity or not? Explain your answer in each case:
 - a) ionic compound in molten state.
 - b) covalent compound in molten state.

Experiment 3.3: Investigating melting points of ionic and covalent compounds

Materials

- | | |
|-----------------------|--|
| a. Sugar | f. tripod stand |
| b. candle wax | g. wire gauze |
| c. common salt (NaCl) | h. 4 evaporating basins |
| d. magnesium chloride | i. beam balance |
| e. source of heat | j. Thermometer that can measure up to 2000°C |

Safety precaution!

You are strictly advised to use goggles when performing this experiment.

Procedure

1. Weigh 40g of sugar in the evaporating basin.
2. Arrange apparatus as shown in Figure 3.2.
3. Heat the weighed amount.
4. Measure the melting point of sugar.
5. Repeat steps 1 to 4 with candle wax, magnesium chloride and common salt.

Discussion

1. What is the melting point of each substance?
2. Which one(s) of the substances are:
 - a) ionic compounds
 - b) covalent compounds
3. Which compounds, ionic or covalent, have:
 - a) higher melting points
 - b) lower melting points
4. Which substances, ionic or covalent compounds, would you expect
 - a) to have higher boiling points? Explain your answer.
 - b) to be more volatile (evaporate easily)? Explain your answer.

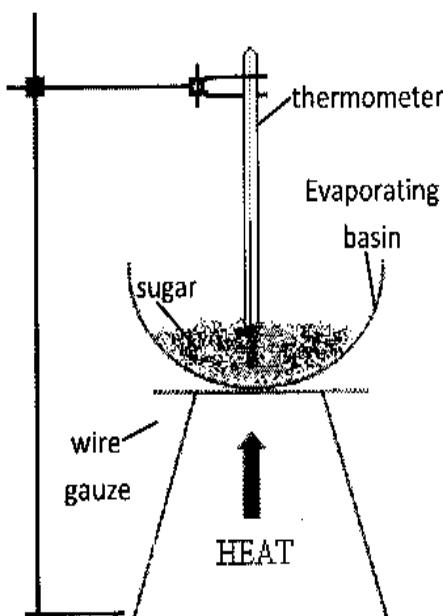


Figure 3.2: Apparatus set for investigating melting points of ionic and covalent compounds

Exercise 3.2

Summarise the following physical properties of ionic and covalent compounds

- a) solubility in water
- b) melting points and boiling points
- c) volatility
- d) electrical conductivity

Structures of ionic and covalent compounds

Particles in both ionic and covalent substances are held together by forces of attraction known as **electrostatic forces**. The electrostatic forces in ionic compounds are different from the electrostatic forces in covalent compound.

Ionic compounds

In ionic compounds, oppositely charged ions are arranged in a regular pattern to form ionic crystal lattices. All the ions in ionic compounds are joined together by electrostatic forces called ionic bonds (Figure 3.3).

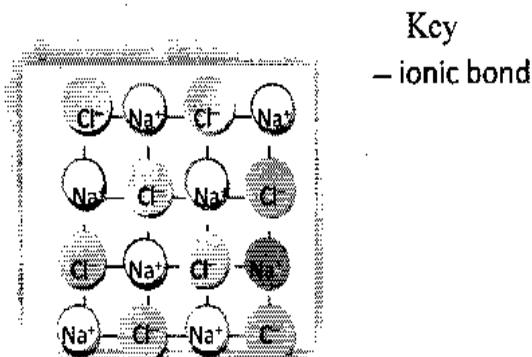


Figure 3.3: Crystal lattice of NaCl

Covalent compounds

In simple covalent compounds such as water and hydrogen chloride atoms are joined together by strong covalent bonds to form molecules. However, the molecules are held together by electrostatic forces called intermolecular forces (Figure 3.4).

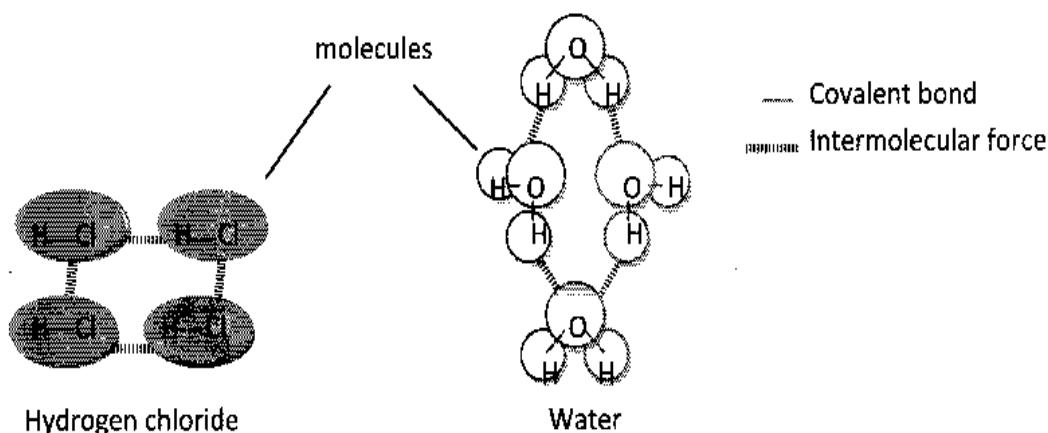


Figure 3.5: Forces in covalent compounds

Experiment 3.4: Comparing the strength of electrostatic forces in covalent and ionic compounds by heating sugar and common salt

Materials

- Sugar
- common salt
- source of heat
- two evaporating basins
- beam balance
- stop clock

Procedure

1. Weigh 40g of sugar in the evaporating basin.
2. Heat the weighed amount using the heat source.
3. Record the time taken for the sugar to melt.
4. Repeat steps 1 to 3 with the common salt.

Discussion

1. Which substance melted faster, sugar or salt?
2. What does the time taken to melt the substance tell you about the strength of electrostatic forces in the substance?
3. Which substance has strong electrostatic forces?
4. What is your conclusion about the strength of electrostatic forces in ionic and covalent compounds?
5. Write a laboratory report of your investigation.

Effects of the structures and bonding on physical properties of ionic and covalent compounds

Melting and boiling of ionic and covalent substances involves breaking electrostatic forces between their particles. To break these forces, energy is required. The type of particles and electrostatic forces between these particles affect the physical properties of the substances such as melting points and electrical conductivity.

Exercise 3.3

1. Name the particles in:
 - a) ionic compounds
 - b) covalent compounds
2. Between ionic and covalent compounds, which compounds are made up of charged particles?
3. From your knowledge of physics, what is electric current?

4. Suggest reasons for the following physical properties of ionic and covalent compounds by referring to strength of electrostatic forces between their particles and type of particles in them.
- Ionic compounds have high melting and boiling points
 - Covalent compounds generally have much lower melting and boiling points than ionic compounds.
 - Ionic compounds do not conduct electricity when in solid state, but they conduct electricity when they are in molten or in aqueous solution states.
 - Covalent compounds do not conduct electricity when in molten state or aqueous solution.

3.2: Bond polarity

Pure covalent bonds and a dative covalent bonds

A covalent bond is formed by two atoms that are sharing a pair of electrons. There are two types of covalent bonds: a pure covalent bond and dative covalent bond (or co-ordinate bond).

When each atom supplies one electron to the bond for the two atoms to share, the bond formed is known as a pure covalent bond.

If both electrons which are shared are donated by one atom, the bond which is formed is called a **dative covalent bond** (or a co-ordinate bond). Dative comes from the Latin word ‘dare’; which means to donate.

Exercise 3.4

- The diagram in Figure 3.6 shows pure covalent and dative covalent bonds in an ammonium ion.
 - How many bonds of the following types are there in the ammonium ion:
 - dative covalent bond(s);
 - pure covalent bond(s)?
 - Name the particles that have formed
 - pure covalent bonds;
 - the dative covalent bonds.
 - What is the difference between dative covalent bond and pure covalent bond?
 - What is a dative covalent bond?



Figure 3.6: Ammonia donating electrons to hydrogen ion

2. A covalent bond is sometimes represented with an arrow pointing away from the particle that has donated the electron pair as shown in the following ammonia-boron trifluoride molecule.



- a) Name the particles that have been bonded together by the covalent bond.
 b) Which particle has donated the electron pair in the dative covalent bond?

Polarity of covalent bonds

A covalent bond can be non-polar, weak polar or strong polar depending on the electronegativity values of the atoms that are bonded together.

Electronegativity

In covalent bonds, the shared electrons are attracted by the nuclei of the bonding atoms. The measure of how strongly an atom pulls the shared electron pair towards itself is called **electronegativity**. Each element has an electronegativity value (see Table 3.1). The electronegativity values are always between 0 and 4.

Table 3.1: Electronegativity values for some elements

Element	Electronegativity
Hydrogen (H)	2.1
Helium (He)	0.0
Lithium (Li)	1.0
Beryllium (Be)	1.5
Boron (B)	2.0
Carbon (C)	2.5
Nitrogen (N)	3.0
Oxygen (O)	3.5
Fluorine (F)	4.0
Neon (Ne)	0.0
Sodium (Na)	0.9

Magnesium (Mg)	1.2
Aluminium (Al)	1.5
Silicon (Si)	1.8
Phosphorus (P)	2.1
Sulphur (S)	2.5
Chlorine (Cl)	3.0
Argon (Ar)	0.0
Potassium (K)	0.8
Calcium (Ca)	1.0
Bromine (Br)	2.8

Exercise 3.5

1. Which element has the highest electronegativity value?
2. What is the electronegativity value of noble gases?
3. Explain the electronegativity value of the noble gases.
4. Apart from the noble gases, which other element has the lowest electronegativity value?
5. If carbon is bonded to hydrogen atom, towards which atom will the bonding electrons be closer? Explain your answer.
6. With exception of the noble gases, between metal and non-metal elements
 - a) which elements have higher electronegativity values?
 - b) which elements have greater ability to attract bonding electrons towards themselves.

Polar covalent bonds

When two atoms with different non-metal elements are bonded to each other, the shared electron pair of electrons will be pulled more strongly towards the atom with higher electronegativity value (Figure 3.7). As a result, this atom will have a slight negative charge because it has more electrons around it while the other atom will have a slight positive charge because it has fewer electrons around it. The slight negative charge and the slight positive charge around the more electronegative atom and the less electronegative atom respectively are known as **partial charges**. The slight negative charges and the slight positive charges are represented by the symbols δ^- and δ^+ respectively as shown in Figure 3.7. The symbol ' δ^- ' is read as delta minus and ' δ^+ ' as delta plus.

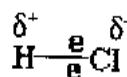


Figure 3.7: Position of bonding electron in atoms of different electronegativities

A covalent bond with a slight positive charge at one end and a slight negative charge at the other end is referred to as a polar covalent bond. The separation of positive and negative charge is known as **bond polarity** and the bond is said to have **electric dipoles** (two poles). Bond polarity results from unequal electron-sharing due to differences in electronegativity values of the atoms involved in the bond.

Non-polar covalent bonds

If two non-metal atoms of the same element are covalently bonded to each other, the electronegativity difference between them is zero and therefore, the bond between them is non-polar covalent bond. In non-polar covalent bond, the bonding electron pair is shared equally between the two atoms (Figure 3.8).

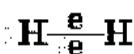


Figure 3.8: Position of bonding electrons between atoms of the same electronegativities

Electronegativity and bonding

The difference between electronegativity values of atoms can be used to tell whether the covalent bond formed between them is non-polar, weak polar or strong polar as shown in Table 3.2.

Table 3.2: Electronegativity difference and bond type

Electronegativity difference	Type of bond
0	Non-polar covalent
0 - 1	Weak polar covalent
1.1 - 2	Strong polar covalent
2.1 and above	Ionic

Exercise 3.6

Consider the following covalent bonds:

- a) C-H
 - b) O-H
 - c) H-Cl
 - d) Cl-Cl
 - e) H-F
 - f) H-Br
1. Calculate the electronegativity difference between the atoms in each covalent bonds.
 2. Arrange the bonds in (1) in order of increasing polarity of bonds.
 3. Which one(s) of these bonds are:
 - i. non-polar?
 - ii. weak polar?
 - iii. strong polar?

Polar and non-polar molecules

Polar molecules

A polar molecule is the molecule that has poles at its ends. This means one end of the molecule has a slight positive charge (δ^+) while the other end has a slight negative charge (δ^-). In polar molecules, electrons in the whole molecule are not evenly distributed in the sense that one end of the molecule has more electrons than the other end (see Figure 3.9).



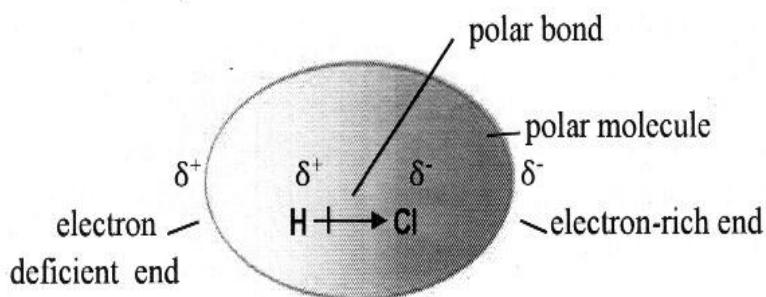


Figure 3.9: Distribution of electrons in a hydrogen chloride molecule

The separation of positive and negative charges in molecules is called **polarization**. Polar molecules are called **dipoles** because they have two poles; a positive pole and a negative pole. All polar molecules have polar bonds. For example, in simple molecules like HCl (Figure 3.9) the bond between the hydrogen and the chlorine atoms is polar and the whole molecule is also polar.

Non-polar molecules

Non-polar molecules are molecules which have no poles. In non-polar molecules electrons are evenly distributed throughout the molecule (see Figure 3.10). All molecules with non-polar covalent bonds like the hydrogen molecule (Figure 3.11) are non-polar molecules.

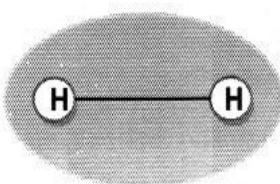


Figure 3.10: Non-polar molecule with non-polar bond

However, there are also some non-polar molecules which have polar covalent bonds. For example, a carbon tetra-fluoride molecule (see Figure 3.11) has polar bonds, but the molecule is a non-polar molecule since it has the same charges all around it.

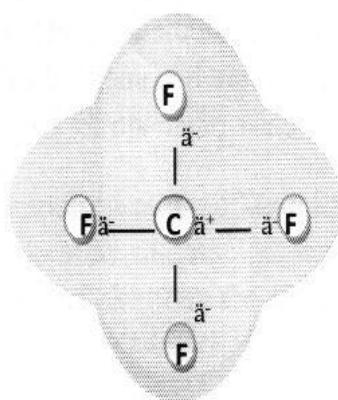


Figure 3.11: Non-polar molecule with polar bonds

3.3: Intermolecular forces (IMF)

Intermolecular forces are electrostatic forces that hold molecules together. They are the forces that enable molecular substances to become liquids or solids. There are two types of intermolecular forces: Van der Waal's forces and hydrogen bonding.

1. Van der Waal's forces

The weakest forces between molecules are called Van der Waal's forces. There are three types of Van der Waal's forces: dipole-dipole forces, dipole-induced dipole forces and induced dipole forces.

A. Dipole-dipole forces

These are forces between dipoles (polar molecules). When one dipole comes close to another dipole, the positive end of one dipole is attracted to the negative end of the other. The two dipoles are then joined together. Figure 3.12 shows two hydrogen chloride molecules joined together by dipole-dipole forces.

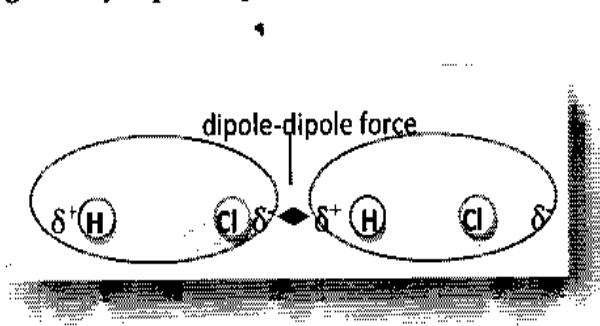


Figure 3.12: Two dipoles joined together by dipole-dipole forces

B. Induced dipole forces

These are forces between non-polar molecules. In non-polar molecules, electrons are evenly distributed (see Figure 3.14). But it happens that, at a particular moment in time the electrons are not be evenly distributed. At this moment the molecule will have temporary partial charges at its ends and temporarily the molecule becomes a dipole (see Figure 3.13). When this happens, the molecules that are next to each other attract each other by very weak forces called induced dipole forces.

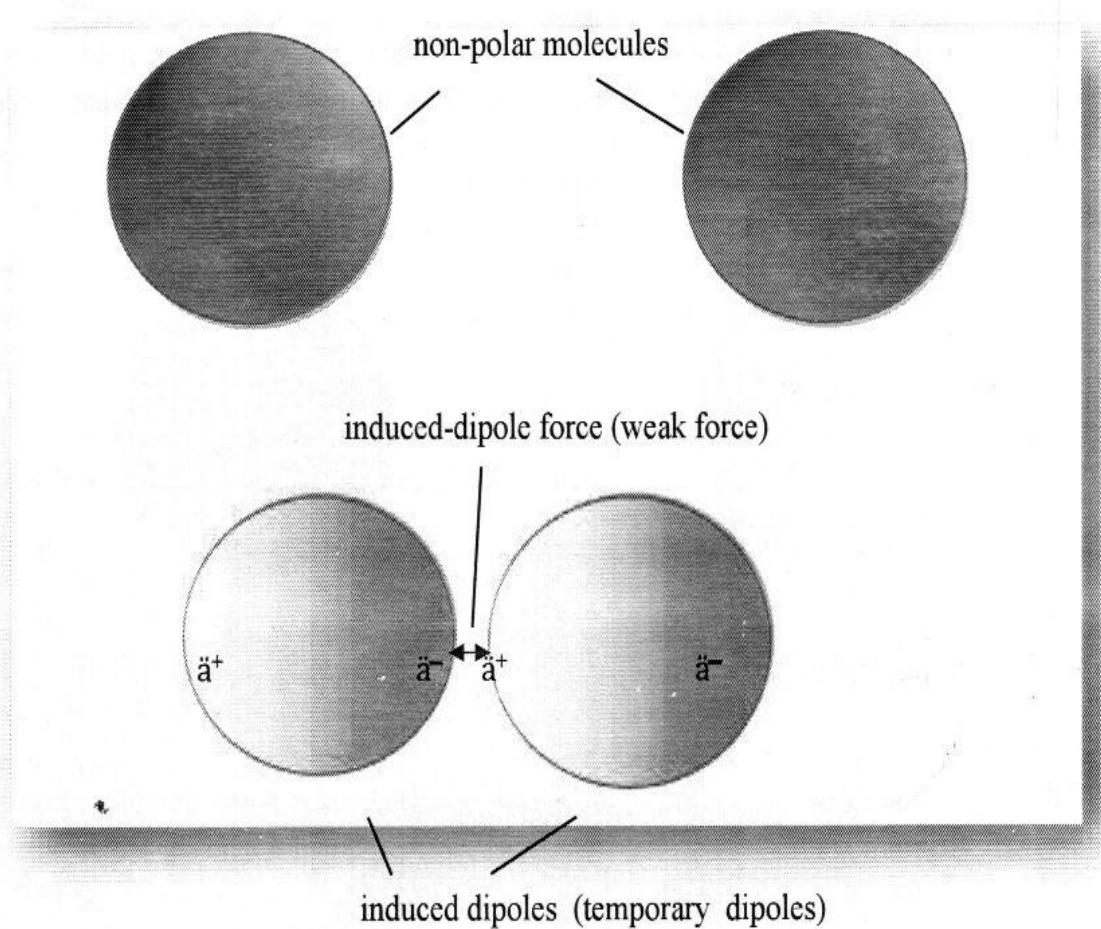


Figure 3.13: Formation induced-dipole forces

Induced dipole forces are also called **London forces**. The induced dipole forces occur between molecules of the substances such as the halogens (e.g. Cl_2), and other non-polar molecules such as hydrogen (H_2) carbon dioxide (CO_2), and nitrogen (N_2).

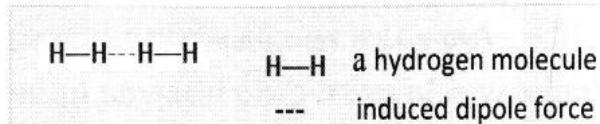


Figure 3.14: An example of molecules held together by induced dipole force

Induced dipole forces are also found in polar molecules in addition to the dipole-dipole forces. The strengths induced dipole forces increase with increase in number of electrons in the molecules. This is because molecules with a lot of electrons are easily polarized and can create stronger temporary dipoles.

C. Dipole-induced dipole forces

When a polar and a non-polar molecule come close to each other, the polar molecule induces a dipole in the non-polar molecule. The two molecules then attract each other and are held together

by a force called dipole-induced dipole force. For instance, the forces that hold chloroform (CHCl_3) molecules and carbon tetrachloride (CCl_4) molecules together are dipole-induced dipole forces (see Figure 3.15)

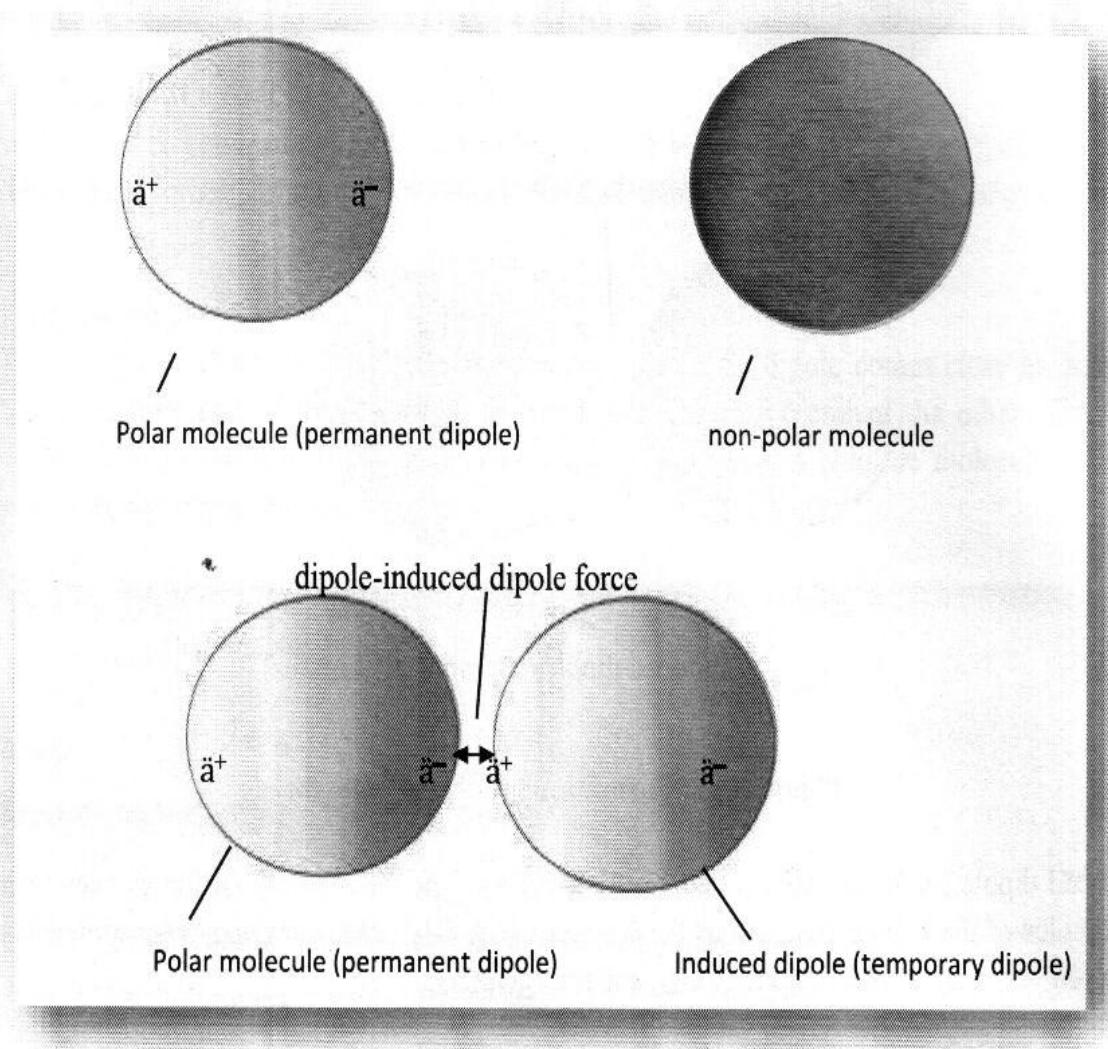


Figure 3.15: Formation dipole induced-dipole forces

2. Hydrogen bonding

Hydrogen bonding is the force which occurs between molecules which contain a hydrogen atom covalently bonded to a highly electronegative atom such as oxygen (O) or fluorine (F). In hydrogen bonding, the highly electronegative atom in one molecule attracts the hydrogen atom on a nearby atom (Figure 3.16). Hydrogen bonding is the strongest of all the intermolecular forces.



Kcy

--- hydrogen bond

— covalent bond

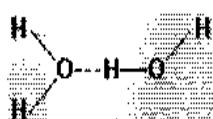


Figure 3.16: Hydrogen bonds

Note that a hydrogen bond is not what holds two hydrogen molecules together. Hydrogen molecules are held together by Van der Waal's forces.

Exercise 3.7

1. In which physical state(s) of matter do intermolecular forces exist between molecules? Explain your answer.
2. Arrange the following intermolecular forces in the order of increasing strength: dipole-dipole, hydrogen bonding, induced dipoles and dipole-induced dipole forces.
3. Name the intermolecular forces that hold together the following molecules:

a. H ₂	e. O ₂	i. H ₂ O and O ₂
b. HBr	f. CH ₄ and CCIF ₃	j. C ₂ H ₅ OH and H ₂ O
c. CH ₄	g. H ₂ O	k. CH ₃ OH and C ₂ H ₆
d. C ₂ H ₅ OH	h. CCIF ₃	l. CH ₄ and C ₂ H ₆
		m.
4. Which intermolecular forces can exist simultaneously between the same molecules?

Effect of intermolecular forces on physical properties of covalent compounds

Strengths of intermolecular forces affect a number of physical properties of covalent compounds. Some of these properties include: solubility in water, melting points and boiling points and volatility.

Solubility

Like dissolves like. This saying means that substances with similar intermolecular forces dissolve in each other. For instance, substances with dipole-dipole forces will dissolve in substances with dipole-dipole forces and substances with induced dipole forces will dissolve in other substances with induced dipole forces. Ionic substances can dissolve in substances which have strong intermolecular forces such as hydrogen bonding because the ionic bonds can easily be disrupted by the formation of the ion-dipole forces.

Melting points and boiling points

Substances with stronger intermolecular forces have higher melting points and boiling points. This is because it requires more heat energy to break the strong intermolecular forces for the substances to melt or boil.

Volatility

Substances with weaker intermolecular forces are more volatile than those with stronger intermolecular forces. This is because weak intermolecular forces require less heat energy to break.

Exercise 3.8

1. Which of the following substances can dissolve in each other? Explain your answer in each case: HBr, CH₄, C₂H₅OH, CCIF₃, H₂O, CH₃OH, C₄H₁₀ and C₂H₆?
2. Explain the following:
 - a. Cooking oil is insoluble in water.
 - b. Grease does not dissolve in water, but it dissolves in paraffin.
 - c. H₂O is a liquid at room temperature, while CH₄ is a gas at room temperature.

3.4: Allotropy and allotropes

Some elements can exist in two or more different forms in the same physical state. This kind of existence is called **allotropy**. The different forms of an element that can exist in the same physical state are called **allotropes**. Allotropy happens when different crystalline solids of an element:

- have different arrangements of atoms and/or molecules, or
- contain different numbers of atoms.

Examples of elements that have allotropes include: carbon, oxygen, phosphorus, tin and sulphur.

Allotropes of Oxygen

Oxygen, as an element, can exist as dioxygen molecules and as trioxygen molecules (commonly known as ozone) (Figure 3.17).

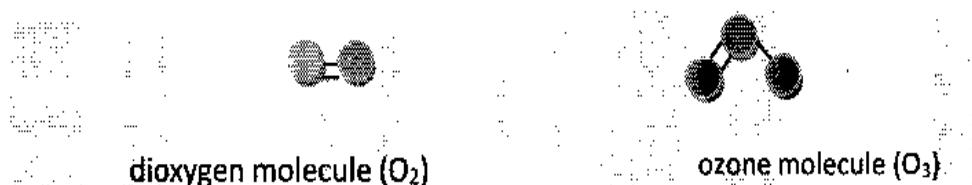


Figure 3.17: Allotropes of oxygen

Allotropes of sulphur

There are a large number of allotropes of sulphur. The most commonly found in nature are rhombic sulphur and monoclinic sulphur (see Figure 3.18). In rhombic sulphur molecules are interlocked. When heated, the S_8 molecules become stacked.

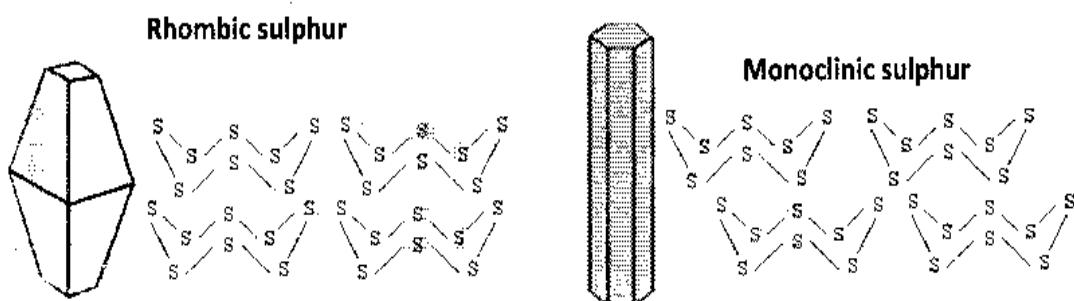


Figure 3.18: Allotropes of sulphur

Exercise 3.9

1. How many atoms are in each of the following molecules:
 - a) dioxygen molecule
 - b) ozone
 - c) monoclinic sulphur
 - d) rhombic sulphur
2. How does allotropy arise in:
 - a) oxygen
 - b) sulphur
3. Describe two ways in which allotropy can arise in elements, using examples of oxygen and sulphur.

3.5: Properties and uses of allotropes of carbon

Carbon has two natural crystalline allotropic forms: graphite and diamond. Each has its own distinct crystal structure and properties.

a) Diamond

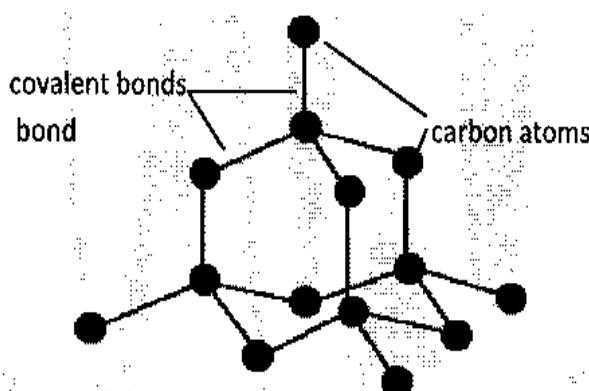
Diamond is one of the allotropes of carbon.



Structure of diamond

In a diamond crystal, each carbon atom is covalently bonded to four other carbon atoms, making a giant macromolecular array, called lattice as shown in Figure 3.19. The covalent bonds on each carbon atom are at an angle of 109° to each other.

All the four valence electrons of each carbon atom are used in covalent bonding; they are said to be localised. Therefore, the movement of electrons in diamond is restricted.



* Figure 3.19: Diagram for the structure of diamond

Exercise 3.10

1. How many valence electrons does each carbon atom have?
2. Explain whether or not there are any free electrons in a diamond structure.
3. Explain the following properties of diamond by referring to its structure as shown in figure 3.19. Diamond:
 - a) is brittle
 - b) does not conduct heat and electricity
 - c) is very hard
 - d) is insoluble in water
 - e) has very high melting point

Uses of diamond in relation to its structure and properties

The structure and properties of diamond make it possible for it to have a number of uses in our daily life. Some of the uses of diamond are as follows:

1. It is used in cutting instruments such as glass cutters, marble cutters and rock drilling equipment.
2. It is used as micro bearings in many small mechanical devices and laboratory instruments.
3. It forms the coating of wear resistant parts which are used to protect the surface of mechanical parts prone to wear and tear.

b) Graphite

Graphite is one of the most common allotropes of carbon. It is the most stable form of carbon under standard conditions.

The structure of graphite

In graphite, each carbon atom uses only three of its four outer energy level electrons to covalently bond to three other carbon atoms in a layer or plane. The other valence electron is contributed to a delocalised system of electrons (Figure 3.21). The delocalised electrons help in bonding the layers of the carbon atoms together. The delocalized electrons are free to move throughout the plane.

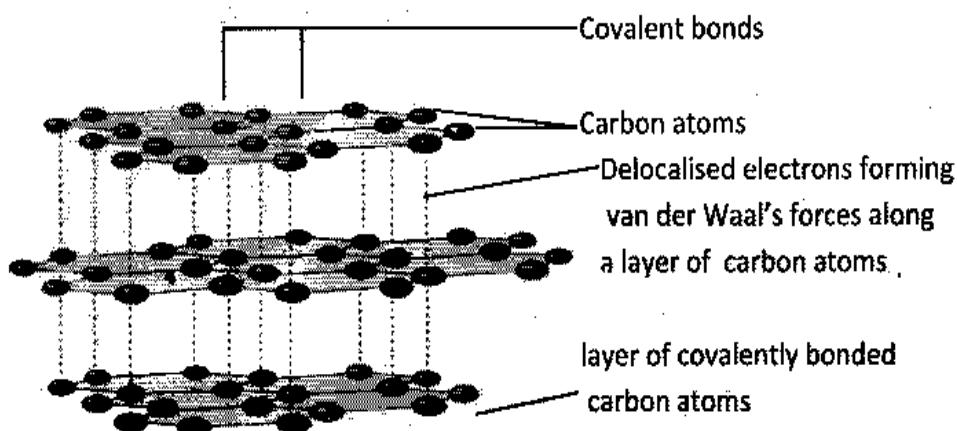


Figure 3.20: The structure of graphite

Exercise 3.11

1. Explain the following physical properties of graphite by referring to its structure as shown in Figure 3.20. Graphite is:
 - a) is a soft and slippery;
 - b) is a good conductor of heat and electricity;
 - c) is brittle;
 - d) is insoluble in water;
 - e) has high melting point.
2. Explain the following properties of graphite by referring to its structure and properties. Graphite is used:
 - a) to make lead pencils;
 - b) as a dry lubricant in machine parts;
 - c) to make crucibles;
 - d) to make carbon electrodes in a dry cells and electric cars.

Similarities between diamond and silicon

Structure of silicon dioxide

Silicon dioxide (SiO_2) is also known as sand. Crystalline silicon has the same structure as diamond (Figure 3.21)

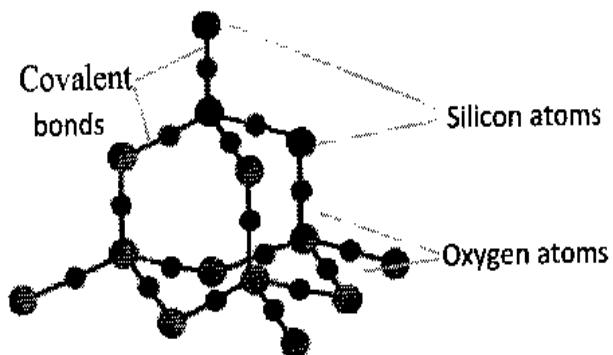


Figure 3.21 The structure of silicon dioxide

Exercise 3.12

1. What is the similarity between the structure of silicon and diamond?
2. What is the difference between the structure of silicon and diamond?
3. How can the structure of diamond be modified to make the structure of silicon?
4. How would you expect the following physical properties to compare between silicon dioxide and diamond?
 - a) hardness?
 - b) heat and electrical conductivity?
 - c) melting and boiling points?

3.6: Properties and uses of metallic substances

Most of the uses of metals are related to their properties such as conductivity of heat and electricity, hardness, ductility, malleability, and lustre.



Experiment 3.5: Investigating heat conductivity of metals

Materials

- A plastic cup (or a beaker)
- Hot water
- Pegs
- Margarine (or Vaseline)
- Stop clock
- Objects: plastic pen case, iron nail, piece of copper wire, glass rod, piece of wood

Procedure

1. Attach each object to the inside of the cup using the pegs as shown in Figure 3.22.
2. Coat each object with a small amount of margarine near the upper end (see Figure 3.22). Make sure the distance from the edge of the cup to the position of margarine is the same on each object.
3. Pour some water into the cup. Make sure the lower part of each object is submerged in the water.
4. Record the time it takes for the margarine on each object to melt

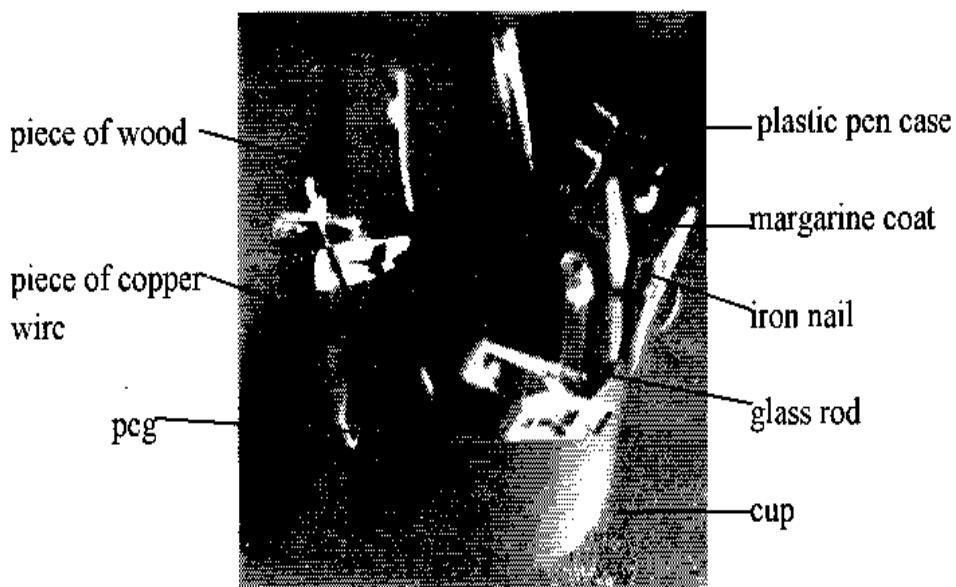


Figure 3.22: Apparatus for demonstrating heat conductivity

Discussion

1. On which object did the margarine melt faster?
2. What is your conclusion on the heat conductivity of metals?

Experiment 3.6: Investigating electrical conductivity of metals

Materials

- Objects: plastic pen case, iron nail, piece of copper wire, glass rod, piece of wood
- A cell in a cell holder
- Four connecting wires
- Electric bulb (1.5 to 3 V)
- A switch

Procedure

1. Connect a switch as shown in Figure 3.23.
2. Connect a piece of wood on the gap XY.
3. Close the switch
4. Record whether the bulb gives light or not.
5. Repeat steps 2 to 3 with the piece of copper wire, plastic pen case, glass rod and the iron nail.

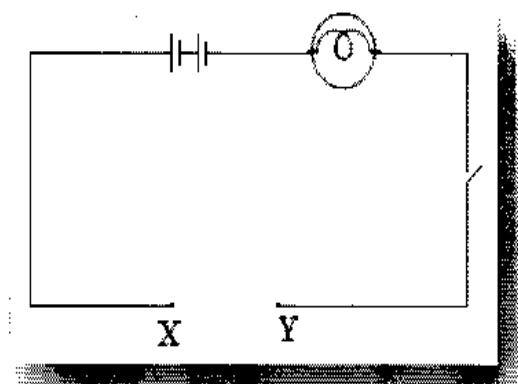


Figure 3.23: Set-up of apparatus for testing electrical conductivity of metals

Discussion

1. After connecting which objects did the bulb give light?
2. What is your conclusion about electrical conductivity of metals?

Experiment 3.7: Investigating lustre (or shiny appearance) of metals

Materials

- An iron nail
- A piece of aluminium wire
- A piece of magnesium ribbon
- A piece of wood
- A piece of plastic material
- Sand paper (or a stone)



Procedure

1. Clean the iron nail with the sand paper or rub it against the stone.
2. Observe its appearance.
3. Repeat steps 1 and 2 with the other materials.

Discussion

1. Which of the materials looked shiny when scratched?
2. What is your conclusion about the appearance of metals when freshly scratched or rubbed?

Experiment 3.8: Investigating malleability of metals (malleability is the ability to be hammered into flat sheets)

Materials

- Small pieces of copper, aluminium, iron nail about 3 inches long and 0.5 cm thick.

Procedure

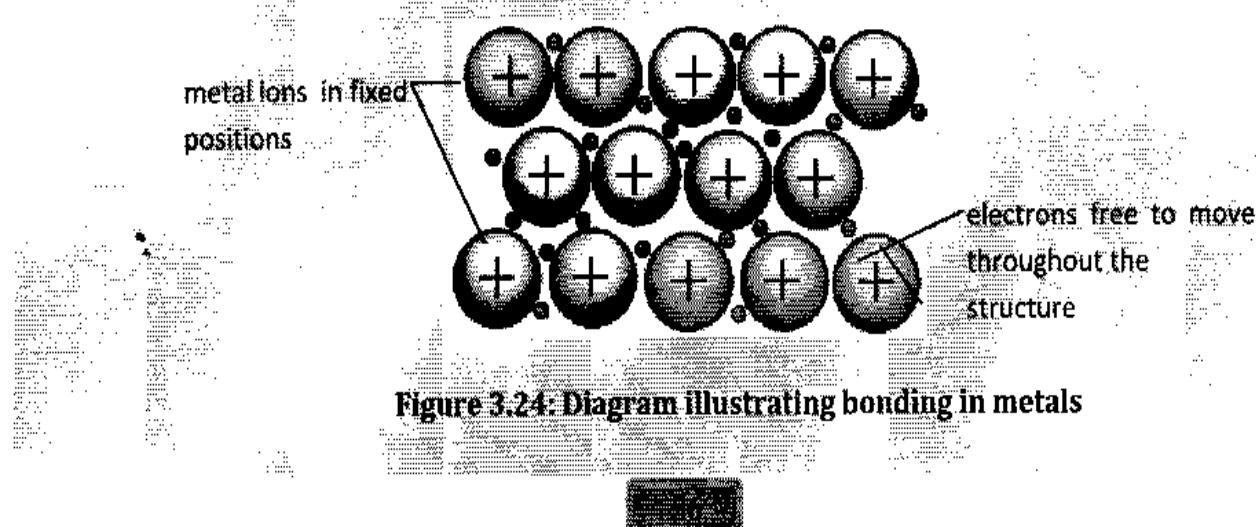
1. Put a piece of copper metal on a block of iron.
2. Hammer it hard 5 to 6 times from the top.
3. Repeat steps 1 and 2 with the other metals.

Discussion

1. What happens to the pieces of metals?
2. Explain whether metals are malleable or not.

Exercise 3.13

1. In Form 2 we discussed the meaning of metallic bond; what does it mean?
2. Explain the following properties of metals by referring to Figure 3.24.
 - a) Electrical conductivity;
 - b) Heat conductivity;
 - c) Hardness;
 - d) Malleability;
 - e) Ductility (ability to be drawn into wires).



Uses of metals

Metals have a number of applications in our everyday life. Some of the uses of metals are as follows: Metals are used to make the following products:

- 1) electrical wires;
- 2) cooking pots;
- 3) sheets such as iron sheets;
- 4) jewellery;
- 5) coins;
- 6) door and window frames.

Exercise 3.14

Relate the uses of metals above to their properties.

3.7: Properties and uses of alloys

Meaning of alloy

Most metals are not as useful in their pure form because they are soft and/or corrosive. For instance, pure copper, lead, iron, aluminium and gold are too soft for many uses. To make a metal useful, its atoms are mixed with atoms of other metals or non-metals to form a mixture called alloy. In alloys atoms of different elements do not chemically bond with one another but are simply mixed together to achieve certain properties (see Figure 3.25).

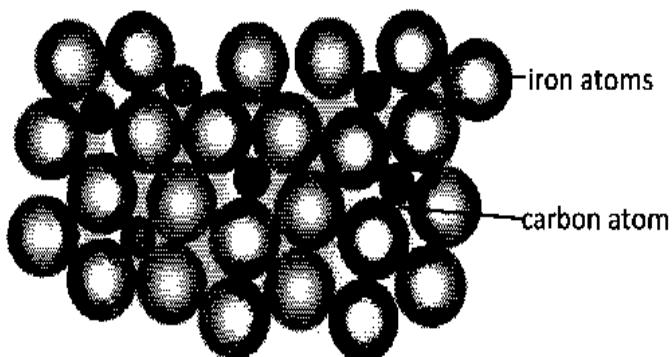


Figure 3.25: An example of an alloy

Exercise 3.15

Table 3.3 shows examples of alloys and their constituent elements. List down other examples of alloys that you knew.

Table 3.3: Alloys and their constituent elements

Alloy	Constituent elements
Stainless steel	Iron and Chromium
Brass	Copper and Zinc

Properties of alloys

Alloys are made to achieve certain properties which pure metals do not have. Some of the properties achieved by alloys include:

- Low density
An example of alloys with low density (which are light) is duralumin. Duralumin is an alloy of aluminium with copper and small amounts of magnesium, silicon and iron.
- Resistance to corrosion
Aluminium-magnesium alloy is highly resistant to corrosion.
Brass is another corrosion resistant alloy.
- Hardness and resistance to corrosion
The alloy that achieves this condition is stainless steel. Stainless steel is alloy of iron with chromium, nickel and carbon.

Uses of alloys

Alloys have wide applications in our everyday life. Some of the uses include manufacturing things shown in **Figure 3.26**:

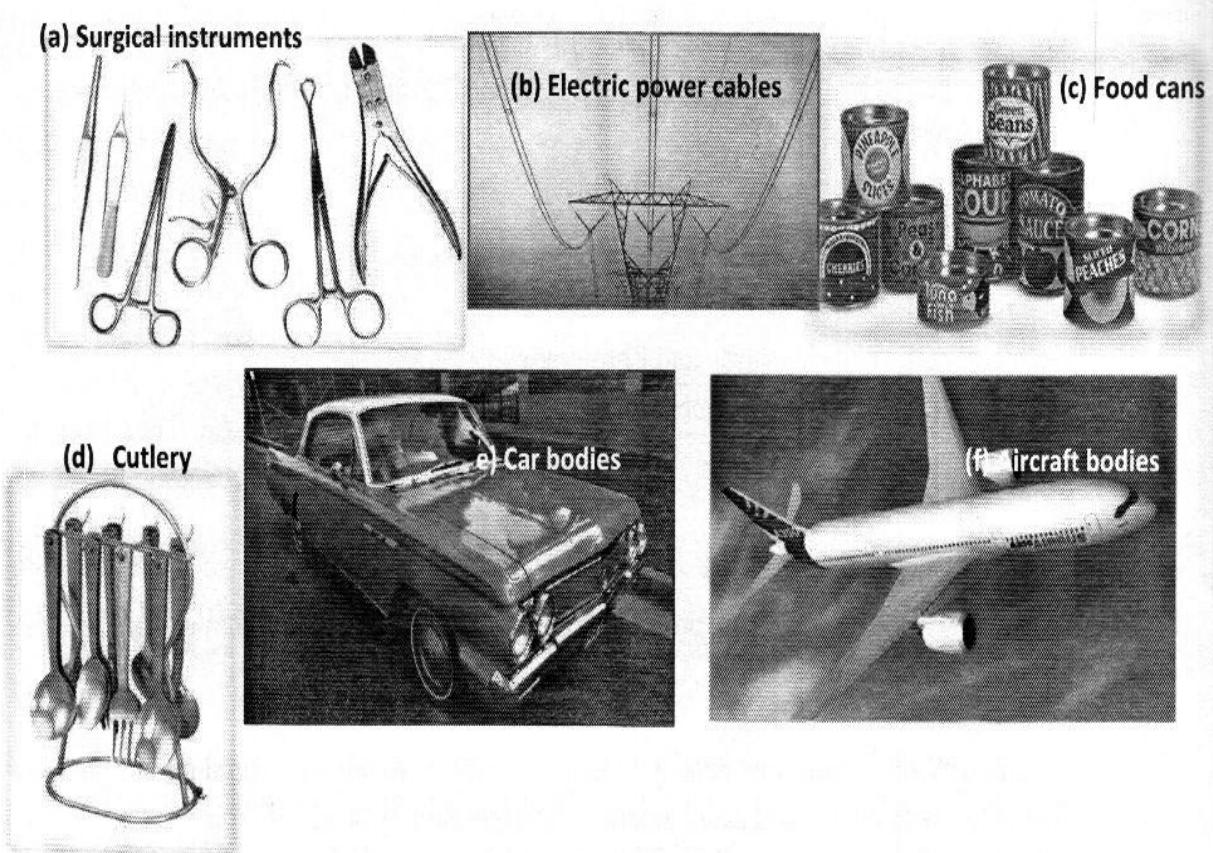


Figure 3.26: Some of the items made using alloys

Exercise 3.16

Explain the following uses of alloys:

- 1) Aluminium-magnesium alloys are used to manufacture food cans.
- 2) Duralumin is used to manufacture
 - a) electric cables.
 - a) aircraft bodies.
- 3) stainless steel is used to manufacture:
 - a) surgical instruments.
 - b) cutlery.
- 4) Steel is used to manufacture car bodies.

Topic review

1. Give three examples of elements or compounds that have simple molecular structures. Why do these have low melting and boiling points?
2. Give three examples of elements or compounds that have giant atomic structures. Why do these have high melting and boiling points?
3. Explain the following properties of metals:
 - a) heat and electrical conductivity
 - b) ductility
 - c) malleability
4. Explain the reason why ionic compounds cannot conduct electricity in solid state, but conduct when in solution or molten states?
5. What is meant by the term 'crystal lattice'?
6. Name the particles that are joined together by electrostatic forces in:
 - a) ionic compounds
 - b) simple covalent compounds
7. Explain the difference between the two types of covalent bonds.
8. State, in terms of differences in electronegativity values, the difference between:
 - a) non-polar covalent bonds
 - b) weak polar covalent bonds
 - c) strong polar covalent bonds
9. Explain the difference between non-polar and polar covalent bonds in terms of how the bonding electrons are shared.
10. What is meant by the term 'dipole'?
11. Describe the following intermolecular forces:
 - a) dipole-dipole
 - b) dipole-induced dipole
 - c) induced dipole
 - d) hydrogen bonding
12. What is the similarity between hydrogen bond and dipole-dipole forces?
13. Why is ethanol able to dissolve in water?
14. Name two electrostatic forces that hold molecules of water together.
15. Explain the difference between the allotropes of:
 - a) oxygen
 - b) sulphur
 - c) carbon

Core Element 4: Chemical reactions

Topic 4: Stoichiometry

Topic 5: Heats of reactions

Topic 4: Stoichiometry

During chemical reactions, atoms of substances ‘rearrange’ to form new substances. Stoichiometry is the calculation of quantities of reactants and products involved in a chemical reaction.

In this topic, we are going to use stoichiometry to balance chemical equations, to determine amounts of products/reactants that are produced/needed in given chemical reactions and to find quantity product yielded by reaction.

By the end of this topic you must be able to:

1. write balanced chemical equations.
2. work out the relative formula mass in a compound.
3. define a mole of a substance.
4. convert moles into other units of measurements.
5. determine the 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 solutions.
9. determine the concentration of a solution using titration.
10. determine the yield in a chemical reaction.

4.1: Balancing chemical equations

Chemical formula

In Form 1, we learnt that a chemical formula is an expression which indicates types of elements in a molecule and the ratio in which they combine to make the molecule.

Exercise 4.1

1. In IUPAC system of naming compounds, how do names of molecular compounds differ from the names ionic compounds?
2. Write down the chemical formula for the ionic compounds below. Use the periodic table to find the sizes of the charges on the ions and their valencies:
 - a) magnesium oxide
 - b) aluminium oxide
 - c) nitrogen trihydride
 - d) dinitrogen monoxide
 - e) nitrogen trifluoride
 - f) hydrogen monochloride
 - g) dihydrogen oxide



Chemical equations

A chemical equation is a written symbolic representation of a chemical reaction. Chemical reactions obey the law of conservation of matter which states that **mass is conserved in chemical reactions**. Mass is defined as quantity of matter. Therefore, the law of conservation of mass actually means that **no atoms can be created or destroyed in a chemical reaction**.

Activity 4.1

Caution: This activity needs the teachers' assistance.

Students have to wear goggles.

Materials: Bunsen burner, magnesium ribbon (about 2 cm long), goggles, tongs, sandpaper, and watch-glass

Procedure

1. Clean the magnesium ribbon with sandpaper.
2. Hold it with a pair of tongs.
3. Burn it using the burner and collect the ash so formed in the watch-glass as shown in Figure 1.1.
4. What do you observe?

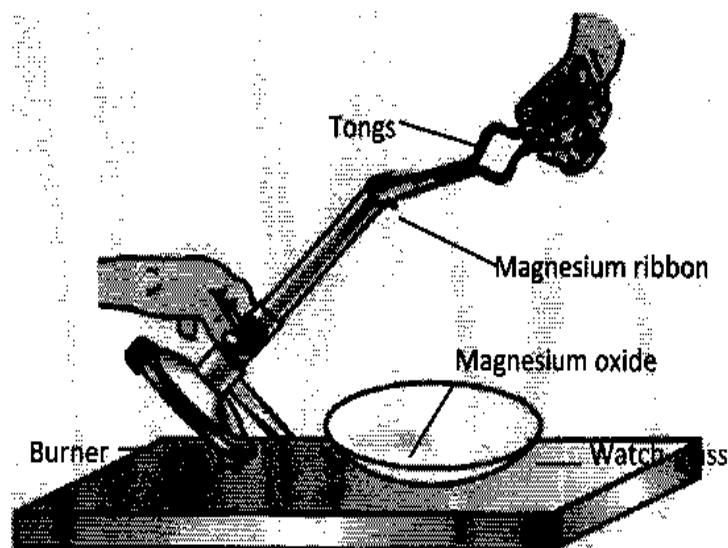
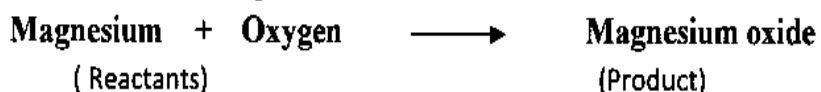


Figure 1.1: Burning magnesium in air

You might have observed, from Activity 4.1, that magnesium burns with a dazzling white flame and changes into white powder. This powder is magnesium oxide. It is formed due to the reaction between magnesium and oxygen present in air. This activity shows that when magnesium burns in oxygen, magnesium oxide is produced. This description of a chemical reaction is quite long. It can be written in a shorter form. The simplest way to do this is to write it in a form of word equation as follows:



Writing a chemical equation

A chemical equation becomes more concise and useful if we use chemical formulae of substances instead of words. If we use chemical formulae of magnesium, oxygen and magnesium oxide, the above word-equation can be written as: $Mg + O_2 \rightarrow MgO$.

Now count the number of atoms of each element on the left-hand side and on the right-hand side of the arrow. Is the number of atoms of each element the same on both sides? If not, then the equation is not balanced and the law of conservation of mass is not obeyed.

Exercise 4.2

Write down chemical equations, using chemical formulae, for the reaction between:

1. sulphur and oxygen to produce sulphur dioxide.
2. aluminium with hydrochloric acid to produce aluminium chloride and hydrogen gas.
3. sodium bicarbonate ($NaHCO_3$) and hydrochloric acid to produce sodium chloride, carbon dioxide and water.
4. ethane (C_2H_6) and oxygen to produce carbon dioxide and water.
5. ammonia (NH_3) and sulphuric acid (H_2SO_4) to produce ammonium sulphate $(NH_4)_2SO_4$.

(Valencies: Al=3, C=4, Cl=1, F=1, H=1, Mg=2, N=3, Na=1, $NH_4^+ = 1$, O=2, S=2, $SO_4^{2-} = 2$)

Balancing chemical equations

To show that the law of conservation of matter is obeyed during a chemical reaction, its chemical equation must show that the number of atoms of each element that are present in the reactants is the same as the number of atoms that are present in the products. **Balancing chemical equation** is the process of balancing atoms of each element on both sides of the chemical equation. Ability to balance chemical equation is a very important skill in chemistry which everyone studying chemistry must acquire. It is for this reason that, after learning how to balance chemical equations in Form 1, it is being repeated here. There are many ways of balancing chemical equations. We are going to discuss two of them: inspection method and algebra method.

Balancing chemical equations by inspection method

In this method, you balance atoms of one element at a time. You should follow the following rules:

- 1) Start by balancing atoms that are found in the most complicated compound.
- 2) Start balancing atoms appear in the lowest number of compounds. You can use fractions at this stage.
- 3) Balance atoms that are present in most compounds at the end (fractions can be used).
- 4) Once all atoms are balanced, remove fractions and find the set of the smallest coefficients.
- 5) Treat polyatomic groups, such as ' SO_4 ' (sulphates), that don't change in the reaction as if they were kind of large 'atoms'; that is to say, instead of balancing atoms in the group individually, balance the group as a whole.

Example 4.1

Balance the equation $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$.

Working out

When there is no coefficient, it means there is only one molecule of the compound, but we will put ones into the equations to differentiate between molecules we have already tried, and those that are not touched yet.

We will start with MgO . So let's put 1 in front of it.



There is one atom of oxygen in on the right-hand side. Since there are two oxygen atoms in one oxygen molecule (O_2) on the left hand-side, to balance the oxygen atoms, we should put $\frac{1}{2}$ in front of O_2 ($\frac{1}{2} \times 2 = 1$):



Removing the fraction

Now there are equal numbers of atoms of each element on both sides of the equation. But to finish balancing, we have to multiply all coefficients by 2 to remove the fraction $\frac{1}{2}$ in front of the O_2 molecule. The equation now becomes:



Example 4.2

Balance the equation $\text{C}_2\text{H}_6\text{O}_2 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$.

Working out

The most complicated molecule is that of ethene glycol ($\text{C}_2\text{H}_6\text{O}_2$).



Balancing carbon atoms

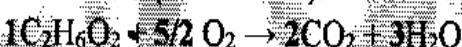


Balancing hydrogen atoms



Balancing oxygen atoms

There are seven oxygen atoms on the right-hand side, four in carbon dioxide and three in water molecules. To balance the equation thus we need seven oxygen atoms on the left-hand side. Two are already present in ethene glycol, so we need five more, therefore, we should put $\frac{5}{2}$ in front of O_2 ($\frac{5}{2} \times 2 = 5$):



Removing the fraction

We should multiply all coefficients by 2 to remove the fraction.



The equation is now balanced.

Exercise 4.3

Balance each equation in Exercise 4.2 by inspection method.

Balancing chemical equations by algebra method

We might have discovered that the last equation in Exercise 4.3 could not be balanced using the inspection method. Thus inspection method has some limitations. A general method that can be used to balance any chemical equation that can be balanced is the **algebra method**. Follow Example 4.3 to understand how this method works.

Example 4.3

Balance the following equation by algebra method.



Working out

Let's assume the balanced equation is $a\text{H}_3\text{BO}_3 \rightarrow b\text{H}_4\text{B}_6\text{O}_{11} + c\text{H}_2\text{O}$, where a , b , and c are the coefficients.

'Balanced' means there is exactly the same number of atoms of each element on each side of the equation. Therefore, using the coefficients a , b and c we can see that we have $a \times 3$ atoms of hydrogen on the left (3 atoms per each H_3BO_3 molecule) and $b \times 4 + c \times 2$ hydrogen atoms on the right (4 atoms per each $b\text{H}_4\text{B}_6\text{O}_{11}$ molecule and 2 atoms per each H_2O molecule). This gives us the following equation: $3a = 4b + 2c$. Similarly we can write equations for elements:

- B: $a = 6b$
- O: $3a = 11b + c$

We can then come up with a total of three equations and three unknowns:

- $3a = 4b + 2c$
- $a = 6b$
- $3a = 11b + c$

All the coefficients must be natural numbers and they must be the smallest ones. To find them, we can assume one of the coefficients to be 1. If $b = 1$, then $a = 6$ and $c = 7$. Therefore, the balanced equation is:



Exercise 4.4

1. Now write the balanced equation for the reaction between ammonia (NH_3) and sulphuric acid (H_2SO_4) to produce ammonium sulphate ($(\text{NH}_4)_2\text{SO}_4$) (Exercise 4.2).
2. Balance the following equations by inspection method.
 - a) $\text{Al} + \text{HCl} \rightarrow \text{AlCl}_3 + \text{H}_2$

- b) $\text{NH}_3 + \text{O}_2 \rightarrow \text{NO} + \text{H}_2\text{O}$
 c) $\text{C}_2\text{H}_5\text{OH} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
 d) $\text{P}_2\text{I}_4 + \text{P}_4 + \text{H}_2\text{O} \rightarrow \text{PH}_3 + \text{H}_3\text{PO}_4$

Writing symbols of physical states

After balancing a chemical equation, it is important to indicate the physical states of the reactants and the products along with their chemical formulae to make equation more informative. The gaseous, liquid, aqueous and solid states of reactants and products are represented by notations (g), (l), (aq), and (s), respectively. The word (aq) is written if the reactant or product is present in solution of water. The equation in Example 4.1 can be written as: $2\text{Mg(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO(s)}$.

4.2: Relative formula masses

Relative formula mass is the sum of the relative atomic masses (A_r) of the elements in a given chemical formula. The symbol for relative formula mass is M_r . If the substance is made up of simple molecules, the relative formula mass may also be called the **relative molecular mass**.

Calculating the relative formula mass

The relative formula mass of a substance is calculated by adding the relative atomic masses (A_r) of all the elements in the chemical formula of the substance.

Example 4.4

Calculate the relative formula mass (M_r) of each of the following substances of water (H_2O)

(A_r : H = 1, O = 16)

Working out

$$\begin{aligned} M_r(\text{H}_2\text{O}) &= 2 \times A_r(\text{H}) + A_r(\text{O}) \\ &= 2 \times 1 + 16 \\ &= 18 \end{aligned}$$

Exercise 4.5

Calculate the relative formula mass (M_r) of the following substances:

- hydrogen gas (H_2)
- sodium chloride (NaCl)
- sulphuric acid H_2SO_4
- magnesium hydroxide Mg(OH)_2
- ammonium sulphate $(\text{NH}_4)_2\text{SO}_4$

(A_r : Na = 23, S = 32, Mg = 24, N = 14, Cl = 35.5)

4.3: The mole concept

The meaning of a mole

A mole is a grouping unit just like a dozen. If there is a dozen of eggs in a carton, then it means there are 12 individual eggs in the carton. The same is said with the mole; if there is a mole of atoms of a substance, then there are 6.02×10^{23} atoms of the substance. A **mole** is, therefore defined as amount of substance that contains 6.02×10^{23} particles of the substance. These particles may be atoms, molecules, formula units or ions. It is symbolised **mol**. The number 6.02×10^{23} is called Avogadro's number (N_A) or Avogadro's constant. This means that for 1 mole of particles of any substance, there is an Avogadro's number of the particles of that substance.

Exercise 4.6

1. Write the Avogadro's constant in standard form.
2. If you have 1 mole of H_2O molecules, how many individual:
 - a) H_2O molecules are there?
 - b) hydrogen atoms are there?
 - c) oxygen atoms are there?
3. Why are atoms, molecules, ions or electrons measured in moles?

Converting moles into other units of measurement

Moles of particles of substances can be converted to other units of measurement such as mass, volumes of gases and number of particles of the substances.

Calculating number of moles and masses of compounds

The relative formula mass, in grams, of any substance (element or compound), contains 1 mole of the particles of that substance. The mass, in grams, of one mole of any substance is called molar mass. **Molar mass = Relative formula mass (in grams)**. The unit of molar mass is gram per mole (g/mol). The molar mass of a substance is important because it can be used to convert:

- a. mass of the substance to number of moles (n) of its particles;
- b. number of moles of the substance to its mass.

a) Converting mass to moles

It is possible to know the number of moles of the particles of a substance, if we know the mass of the substance. To find the number of moles (n) of a substance, given its mass, we divide the mass (in grams) of the substance by its molar mass.

$$\text{Number of moles of a substance} = \frac{\text{mass of substance (g)}}{\text{molar mass (g/mol)}}$$



Example 4.5

How many moles of sodium (Na) are there in 46 g of sodium? (A_r : Na = 23)

Working out

- Mass of Na sample = 46g
- Molar mass of Na = 23g/mol
- Number of moles of Na = $\frac{\text{mass of Na (g)}}{\text{molar mass (g/mol)}}$

$$\bullet \text{Number of moles of Na} = \frac{46\text{g}}{23\text{g/mol}}$$

- Number of moles of Na = 2 moles

Alternatively, using cross multiplication we have:

$$23\text{g Na} = 1\text{mole}$$

$$46\text{g Na} = x$$

$$23\text{g Na} \times x = 46\text{gNa} \times 1\text{mole}$$

$$x = \frac{46\text{gNa} \times 1\text{mole}}{23\text{g Na}}$$

$$x = 2 \text{ moles}$$

b) Converting moles to mass

We have learnt that, given mass of a substance, the number of moles of the substance can be found by dividing mass of the substance (in grams) by the molar mass of the substance.

Therefore, to find the mass of a sample of a substance, given the number of moles, we multiply the number of moles of the sample by its molar mass.

Mass of a substance = number of moles × molar mass.

Exercise 4.7

- I. Consider the molecule of water (H_2O):
 - a) What is the relative formula mass of water?
 - b) How many:
 - i. molecules of water are there in 18g of water?
 - ii. oxygen atoms are in 18g of water?
 - iii. hydrogen atoms are there in 18g of water?

2. How many moles are in each of the following masses of substances:
- 73g of sodium chloride?
 - 13.5g of aluminium?
 - 36g of magnesium metal?
3. What is the mass of each of the following:
- 0.2 moles of HCl
 - 3 moles of NaOH
 - 0.6 moles of calcium carbonate (CaCO_3)
- (Ar: C = 12, O = 16, Na = 23, H = 1, Cl = 35.5, Mg = 24, Al = 27, Ca = 40)

Molar volumes of gases

The volume occupied by any gas depends on its temperature and pressure. At a given temperature and pressure, all gases contain the same number of molecules (see Figures 4.2 and 4.3). The volume occupied by 1 mole of a gas is known as molar volume. At standard temperature (0°C) and pressure (1 atmosphere) (s.t.p), the molar volume of gases is 22.4dm^3 (Figure 4.2), while room temperature (25°C) and pressure (1 atmosphere) (r.t.p) molar volume of gases is 24dm^3 (Figure 4.3). Molar volume can be used to calculate the number of moles of a gas contained in a given volume of the gas at a particular temperature and pressure by using the formula:

$$\text{Number of moles of a gas} = \frac{\text{volume of gas}}{\text{molar volume}}$$

$t = 0^\circ\text{C}$	CO_2	H_2	N_2
$p = 1\text{atm}$			
Volume of each gas	22.4dm^3	22.4dm^3	22.4dm^3
Mass of each gas	44g	2g	28g
Number of molecules in gas	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Figure 4.2: Volume occupied by one mole of molecules of gases at s.t.p.

$t = 25^{\circ}\text{C}$			
$p = 1 \text{ atm}$			
Volume of each gas	24dm^3	24dm^3	24dm^3
Mass of each gas	44g	2g	28g
Number of molecules in each gas	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Figure 4.3: Volume occupied by one mole of molecules of gases at r.t.p.

Exercise 4.8

1. How many moles of gases do the following volumes of gases contain:
 - a) 22.4dm^3 of methane gas at s.t.p?
 - b) 24dm^3 of water vapour at r.t.p?

2. Use molar volumes of gases to answer the following questions:
 - a) How many moles of CO_2 gas are there in 60 litres of CO_2 gas at r.t.p?
 - b) A reaction produces 100 cm^3 of hydrogen gas at r.t.p. How many moles of H_2 is this?
 - c) What volume does 48g of an oxygen gas, O_2 occupy at: (i) r.t.p.? (ii) s.t.p.?
 - d) What is the mass of 56 dm^3 of hydrogen gas (H_2) at s.t.p?

Using the Avogadro's constant to find the number of particles of a substance

We have already seen that one mole of a substance contains an Avogadro's number of the particles of the substance. Therefore, we can use the Avogadro's number to calculate the number of particles in a sample of a substance.

Number of particles of a substance = number of moles (n) \times Avogadro's constant (N).

Example 4.6

How many molecules of water are there in 2.5 moles of water?

Working out

Avogadro's constant = 6.02×10^{23} molecules/mol

Number of moles = 2.5 mol

Number of particles of a substance = number of moles \times Avogadro's constant

$$\begin{aligned} \text{Therefore, number of molecules of water} &= 2.5 \text{ mol} \times 6.02 \times 10^{23} \text{ molecules/mol} \\ &= 1.505 \times 10^{24} \text{ molecules} \end{aligned}$$

Cross multiplying:

$$1 \text{ mole} = 6.02 \times 10^{23} \text{ molecules}$$

$$2.5 \text{ moles} = x$$

$$x \times 1 \text{ mole} = 2.5 \text{ moles} \times 6.02 \times 10^{23} \text{ molecules}$$

$$x = (2.5 \text{ mol} \times 6.02 \times 10^{23} \text{ molecules/mol}) = 1.505 \times 10^{24} \text{ molecules}$$

Exercise 4.9

1. Define a mole in terms of:

- a) molar mass
- b) molar volume of gases

2. How many molecules will 6 litres (dm^3) a gas have at:

- a) s.t.p?
- b) r.t.p?

3. Calculate the number of particles in each of the following substances:

- a) 54g of aluminium metal
- b) 30 dm^3 of chlorine gas at s.t.p.
- c) 60 dm^3 of methane gas at r.t.p
- d) 87.75g of NaCl

(A: Al = 27, Cl = 35.5, Na = 23, C=12, H=1)

4.4: The percentage composition of water in molecular and hydrated ionic compounds

Some compounds contain water molecules, or its constituent elements in their formulae. These compounds are called **hydrates**. Some hydrates are molecular compounds while others are ionic compounds.

Molecular hydrates

Molecular hydrates contain constituent elements of water, which are hydrogen (H) and oxygen (O) in their molecules. An example of molecular hydrates is sugar. The molecule of sugar, $\text{C}_6\text{H}_{12}\text{O}_6$, consists of carbon (C), hydrogen (H), and oxygen (O), with hydrogen to oxygen atom ratio of 2:1 as in water. The molecular formula of sugar can also be written as $\text{C}_6(\text{H}_2\text{O})_6$. Sugar is therefore, considered as a hydrate of carbon, which is called carbohydrate. Other examples of molecular hydrates are ethanol ($\text{C}_2\text{H}_5\text{OH}$), which is considered as a hydrate of ethane. A molecule of water can be eliminated from molecular hydrate by heating or by the action of sulphuric acid.

Hydrated ionic compounds

Some ionic compounds contain water molecules that are bonded to the central metal cation in the compounds using dative bonds, or that have crystallised with the metal complex. The water molecules are combined in definite ratios. Such ionic compounds are called hydrated ionic compounds and the water that is bonded to the compounds is called **water of crystallization**.

An example is cobalt chloride (CoCl_2) which is blue in colour, but turns to red when hydrated. The formula for the hydrated cobalt chloride is

Molecules of water of crystallization contribute to the mass of the compound. Therefore, when calculating the relative formula mass of hydrated compounds, the molecules of the water of crystallization must be taken into account. Percentage of water of crystallization can be calculated using the chemical formula of the hydrate and is called theoretical percentage of water of crystallization.

$$\text{Theoretical \% of water of crystallization} = \frac{\text{mass of water in a hydrate}}{\text{mass of hydrate}} \times 100\%$$

Exercise 4.10

Calculate the theoretical % of water of crystallization in each of the following hydrated compounds:

- magnesium sulphate $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$.
 - hydrated sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
- (A_r: C = 12, Na = 23, Mg = 24, S = 32, O = 16 and H = 1)

Percentage of water in molecular and hydrated ionic compounds

Water in hydrates can be removed through evaporation by heating the hydrate. When water in a hydrate is removed, the material that remains is said to be anhydrous and it is called anhydrate. The process of removing water from a hydrate to form an anhydrate is called dehydration. The percentage of water in a hydrate can be determined experimentally as follows:

- Heating a known mass of the hydrate to remove the water of hydration
- Measuring the mass of the anhydrate remaining
- The difference between the mass of the hydrate and the mass of the anhydrate is the mass of water that has been lost.

$$\text{The percentage of water in a hydrate} = \frac{\text{mass of water lost}}{\text{original mass of hydrate}} \times 100\%$$

Experiment 4.1: Investigating the percentage of water in sugar by heating

Materials

- Sugar
- Crucible (or an evaporating basin)
- Source of heat
- Beam balance

Procedure

- 1) Copy Table 4.1 for your results.
- 2) Weigh the empty crucible and record the mass.
- 3) With the crucible still on the balance, add sugar until the mass increases by approximately 10g.
- 4) Record the mass of sugar.
- 5) Heat the sugar in the crucible until a dry, black solid (carbon) is formed.
- 6) Weigh the crucible + the black substance.
- 7) Heat and reweigh several times until the mass is constant.
- 8) Record the mass of crucible + carbon.
- 9) Calculate the mass of carbon.
- 10) Calculate the mass of water lost.

Table 4.1: Table of results

Item	Mass (g)
Empty crucible	
Crucible + sugar	
Sugar	
Crucible + Carbon	
Carbon	
Water lost	

Conclusion

1. What is the colour of carbon?
2. Write balanced chemical equation for dehydration of sugar by heating.
3. Calculate the experimental percentage of water of hydration in sugar.
4. What is the theoretical percentage of water in sugar?
5. Explain the difference, if any, between the theoretical and the experimental percentage of water in sugar.

Experiment 4.2: Investigating the percentage of water in hydrous copper sulphate by heating

Materials

- Blue copper sulphate
- Crucible
- Source of heat
- Beam balance

Procedure

- 1) Draw a table to record your results.
- 2) Weigh the empty crucible and record the mass.
- 3) With the crucible still on the balance, add blue copper sulphate until the mass increases by approximately 10g.
- 4) Record the mass of blue copper sulphate.
- 5) Heat the copper sulphate in the crucible until a dry, white solid (anhydrous copper sulphate) is formed. ↴
- 6) Weigh the crucible + the white substance.
- 7) Heat and reweigh several times until the mass is constant.
- 8) Record the mass of crucible + anhydrous copper sulphate in the table of results.
- 9) Calculate the mass of anhydrous copper sulphate.
- 10) Calculate the mass of water lost;

(Mass of water lost = mass of blue copper sulphate – mass of white copper sulphate)

Conclusion

- 1) Write balanced chemical equation for the dehydration of blue copper sulphate.
- 2) Calculate the experimental percentage of water in the hydrated copper sulphate.
- 3) Calculate the theoretical percentage of water in hydrated copper sulphate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.
- 4) Explain the difference, if any, between the theoretical and the experimental percentages of water in hydrated copper sulphate.

(Ar: Cu = 63.5, S = 32, O = 16, H = 1)

4.5: Empirical and molecular formulae

Meaning of empirical formula

In Form 1, we learnt that a chemical formula indicates two things: types of elements in a molecule and the ratio in which the elements combine to make the molecule. In some chemical formulae, the ratio of the elements may be in its lowest terms, while in others, it is not. When the ratio of the elements in a chemical formula is in its lowest term, then the chemical formula is called **empirical formula**.



Exercise 4.11

1. Consider the following chemical formulae:
 - a) NaCl
 - b) C₂H₆
 - c) C₃H₈
 - d) H₂O
 - e) C₃H₆
 - i. Write the ratio of the numbers of atoms in each chemical formula.
 - ii. Which of these chemical formulae are empirical formulae? Explain your answer.
2. If the chemical formula of hexane is C₆H₁₄:
 - a) write the ratio of carbon atoms to hydrogen atoms in its lowest term.
 - b) what is the empirical formula of hexane?

How to work out empirical formulae

If compositions of compounds by masses or percentages are known, the empirical formulae of the substances can be worked out. To work out the empirical formula, remember the following:

- a) If the composition is given as percentage, assume that there is 100g of the substance and express the percent compositions as mass compositions in grams.
- b) Convert the masses to moles. This gives us mole ratio of the elements.
- c) Express the mole ratio in whole numbers in the lowest terms. A good technique is to start by dividing all the numbers in the ratio by the smallest number of moles.
- d) Write the empirical formula with whole numbers in the mole ratio as subscripts. It is standard not to write the subscript if the number is 1.

Example 4.7

An experiment shows that 2.4g magnesium combines with 1.6g oxygen to form magnesium oxide. What is the empirical formula of magnesium oxide? (Ar: Mg = 25, O = 16).

Working out

a. Converting masses to moles

Magnesium

$$\begin{aligned}\text{Mass of Mg} &= 2.4\text{g} \\ \text{Molar mass} &= 24\text{g/mol} \\ \text{Moles of Mg} &= \frac{2.4\text{g}}{24\text{g}} \times 1\text{mole} \\ &= 0.1 \text{ mole}\end{aligned}$$

Oxygen

$$\begin{aligned}\text{Mass of Mg} &= 1.6\text{g} \\ \text{Molar mass} &= 16\text{g/mol} \\ \text{Moles of Mg} &= \frac{1.6\text{g}}{16\text{g}} \times 1\text{mole} \\ &= 0.1 \text{ mole}\end{aligned}$$

b. Mole ratio of atoms

Ratio of Mg: O = 0.1:0.1

Dividing by smallest ratio number (or scaling up a factor) to get simplest whole number ratio

$$\frac{0.1}{0.1} \cdot \frac{0.1}{0.1} \text{ (or } 0.1 \times 10 : 0.1 \times 10) = 1.1$$

c. Empirical formula

The simplest whole number ratio of Mg : O is 1:1. Therefore, the empirical formula for magnesium oxide is MgO .

Example 4.8

1.35g of aluminium was heated in oxygen until there was no further gain in weight. The aluminium oxide formed was weighed and its mass was 3.75g. Work out the empirical formula of aluminium oxide. (A_e , Al = 27, O = 16)

Working out

a. Converting masses to moles

Moles of aluminum

Mass of Al heated = 1.35g

Molar mass of Al = 27 g/mol

$$\text{Moles of Al} = \frac{1.35 \text{ g}}{27.9 \text{ g/mol}} \times 1 \text{ mol} = 0.05 \text{ mole}$$

Moles of oxygen

Mass of O_2 = Mass of aluminium oxide – mass of aluminium

-3.75g- 1.35g

- 24 -

Molar mass (O_2) = 32 g/mol

$$\text{Moles of Mo} = 2.47 \times 1 \text{ mole}$$

— THE —

b. Mole ratio of atoms

Ratio of Al:O = 0.05:0.075

Dividing by smallest ratio and multiplying by 2 to get smallest whole number ratio, we get:

$$\begin{array}{r} 0 \quad .05 \quad 0.075 \\ -0.05 \qquad \qquad 0.05 \\ \hline \end{array} = 1 \frac{1}{2}$$

Clearing the fraction

$$= (1 \times 2) : \left(\frac{3}{2} \times 2 \right)$$

c. The empirical formula

The simplest whole number ratio of Al to O atoms is 2 : 3, so the empirical formula of aluminium oxide is Al_2O_3 .

Exercise 4.12

1. An experiment shows that 32g of sulphur combine with 32g of oxygen to form a compound of sulphur and oxygen. Work out the empirical formula of the compound.
2. A hydrocarbon has a mass composition of 85.7% of carbon and 14.3% of hydrogen. Calculate the empirical formula of the hydrocarbon.
3. A compound is 47% carbon, 10.6% hydrogen and 42.0% sulphur. What is its empirical formula?

(A_r: S = 32, O = 16, H = 1, C = 12)

Molecular formulae

If the ratio of elements in a chemical formula is not in its lowest term, then the formula is called **molecular formula**. Molecular formulae show the actual numbers of atoms of each element in a molecule. In order to deduce the molecular formula, we need to know the molecular formula mass of the compound. It is worked out as follows:

- a) Work out the empirical formula of the compound.
- b) Calculate the empirical formula mass.

molecular formula mass

- c) Calculate the ratio: empirical formula mass

- d) Multiply each subscript in the empirical formula by the ratio:

Example 4.9

An alkane has percentage composition of 84.2 % carbon and 15.8% hydrogen. Its molecular formula mass is 114g. What is its molecular formula? (A_r: C = 12, H = 1).

Working out

Assume there is 100g of the substance, and then express the percent compositions as masses in grams. Therefore, the sample has 84.2g carbon and 15.8g hydrogen.

- a) Empirical formula

Moles of carbon

$$\text{Mass of C} = 84.2\text{g}$$

$$\text{Molar mass of C} = 12\text{g/mol}$$

$$\text{Moles of C} = \frac{84.2\text{g}}{12\text{g}} \times 1\text{mol}$$

$$= 7.02 \text{ moles}$$

Moles of hydrogen

$$\begin{aligned}\text{Mass of H} &= 15.8 \text{ g} \\ \text{Molar mass of H} &= 1 \text{ g/mol} \\ \text{Moles of Mg} &= \frac{15.8 \text{ g}}{1 \text{ g}} \times 1 \text{ mole} \\ &= 15.8 \text{ moles}\end{aligned}$$

Ratio of atoms

$$\begin{aligned}\text{C}_4\text{H}_{10} &= 7.02 : 15.8 \\ \frac{7.02}{7.02} : \frac{15.8}{7.02} &= 1 : 2.25 \\ &= 1 : 2\frac{1}{4} \\ &= 1 : \frac{9}{4} \\ 1 : \frac{9}{4} &= (1 \times 4) : (\frac{9}{4} \times 4) \\ &= 4 : 9\end{aligned}$$

The simplest whole number ratio of C to H atoms is 4 to 9. Therefore, the empirical formula of the alkane is C_4H_9 .

b) Empirical formula mass

$$\text{Empirical formula mass} = (12 \frac{\text{g}}{\text{mol}} \times 4 \text{ mol}) + (1 \frac{\text{g}}{\text{mol}} \times 9 \text{ mol}) = 57 \text{ g}$$

c) Dividing the molecular formula mass by the empirical formula mass

$$= \frac{114 \text{ g}}{57 \text{ g}} = 2$$

d) Multiplying the subscripts in the empirical formula mass by 2 to get the molecular formula:



Therefore, the molecular formula of the alkane is C_8H_{18} .

Exercise 4.13

1. 1.2g of magnesium combines with 3.55 g of chlorine. If the molecular formula mass is 95g, work out the molecular formula of the compound.
2. An organic acid has a percentage composition by mass of C = 40%, H = 6.7% and O = 53.3%. Its relative molecular mass is 60. Calculate its:
 - a) empirical formula
 - b) molecular formula(Ar: Mg = 24, Cl = 35.5 C = 12, H = 1, O = 16).



4.6: Concentration of solutions

Meaning of concentration

Concentration of a solution refers to the amount of solute present in a given volume of the solution. The amount of solute can be in moles or grams.

$$\text{Concentration} = \frac{\text{amount of solute}}{\text{volume of solution}}$$

In Form 1, we learnt about different types of solutions such as gas-in-solid, liquid-in-solid and solid-in-liquid solutions. We are going to discuss concentrations of solid-in-liquid and liquid-in-liquid solutions.

Exercise 4.14

Chimwemwe dissolved 30g of sugar ($\text{C}_6\text{H}_{12}\text{O}_6$) in water to make 80 cm^3 of sugar solution (Figure 4.3):

1. What does the term 'solution' mean?
2. In this solution which substance is:
 - a) the solvent.
 - b) solute.
3. How many moles of sugar are there in 30g of sugar?
4. Express 80cm^3 of sugar solution in dm^3 ($1\text{dm}^3 = 1000\text{cm}^3$).
5. Calculate the concentration of this sugar solution in:
 - a) g/cm^3 .
 - b) mole/dm^3 .

($\text{Ar: C} = 12, \text{H} = 1, \text{O} = 16$)

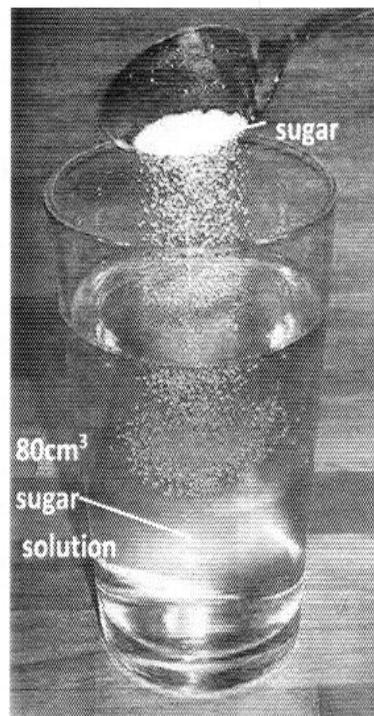


Figure 4.4: Making sugar Solution

Ways of expressing concentration

Concentration of a solution can be expressed as molar concentration (mole/dm^3), mass concentration (g/cm^3), or as percentage.

a) Molar concentration

Molar concentration indicates amount of solute in moles contained in 1dm^3 of solution ($1\text{dm}^3 = 1\text{litre}$). Concentration expressed in moles per dm^3 is also called molarity.

$$\text{Molarity} = \frac{\text{amount of solute in moles}}{\text{volume of solution} (\text{dm}^3)}$$

(Or molarity = n/v , where n is number of moles, and v is volume of solution in dm^3).

Molarity is the commonly used method of expressing concentration. The unit of molarity is mol/l or mol/dm³. The unit mol/dm³ is given the name molar, and it is abbreviated as M. To calculate the molarity of a solute in a solution, we need to know:

1. number of moles of the solute present in the solution.
2. volume of solution in litres (or in dm³).

Exercise 4.15

1. What would be the molarity of a solution where 0.2 moles of a solute are dissolved to make 2.5 dm³ of solution?
2. 100 ml of a solution was made by dissolving 20 g of NaOH in water. What is the molarity of the solution?
3. What is the concentration of a solution containing 4 moles of NaOH in 2 dm³ of solution?
4. What is the concentration of sodium carbonate solution containing 53 g of the salt (Na₂CO₃) per dm³?
5. How many moles of the solute are there in each of the following solutions?
 - a) 50 cm³ of 4M sugar (C₆H₁₂O₆) solution?
 - b) 200 cm³ of 0.5M hydrochloric acid?
 - c) 500 cm³ of 0.1M NaOH solution?

(A.: Na = 23, C = 12, O = 16, H = 1)

b) Mass concentration

Mass concentration indicates the mass of solute dissolved in a given volume of a solution. It is usually expressed in the units of g/cm³. By definition,

$$\text{Mass concentration} = \frac{\text{amount of solute in grams (g)}}{\text{volume of solution (cm}^3\text{)}}$$

To calculate mass concentration of a solution we need to know the following:

- a) mass of the solute in gram.
- b) volume of the solution in cm³.

Exercise 4.16

Calculate mass concentration of the following solutions

- a) 500 cm³ of a solution containing 20 g of sodium hydroxide (NaOH).
- b) 600 cm³ of a solution containing 3.5 g of sodium chloride (NaCl).
- c) 500 cm³ of a solution containing 5.6 g of potassium hydroxide (KOH).
- d) 2 dm³ of solution containing 14.2 g of lithium hydroxide (LiOH).

(A.: Na = 23, H = 1, O = 16, Cl = 35.5, Li = 7)

c) Percentage concentration

There are different ways of expressing percentage concentration: mass percentage, volume percentage, and mass-volume percentage. When reporting concentration in percentage, it is necessary to specify which percentage concentration is being referred to.

i. Mass percentage (mass/mass %)

Mass percentage concentration is mass of solute present in 100g of solution. It indicates the mass of solute dissolved in 100g of solution. For example, if the mass percentage of a NaCl solution is 24%, it means that every 100g of the solution contains 24g NaCl.

$$\text{Mass/mass \%} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

To calculate percent by mass we need the following information:

- a) mass of the solute dissolved;
- b) mass of solution.

Mass percentage concentration is usually used to express concentration of acids and bases. For example, commercially available concentrated hydrochloric acid (HCl) is 37% by mass.

Exercise 4.17

1. 10.00g of BaCl₂ was dissolved in water to make 90.0g of solution. Calculate the mass percent concentration of the solution.
2. What is the mass percent of the following solutions:
 - a) 30g sucrose in 120.0g of water?
 - b) 5.3g ethanol in 85.0g of water?
3. Calculate the mass of NaCl in 400g of 2.5% solution of sodium chloride.

Converting mass percentage to molar concentration

Sometimes we need to convert mass percentage to molar concentration for easy calculations. To be able to do this, we need the following information:

- a) mass percentage concentration of the solution;
- b) molar mass of the solute (g/mol);
- c) density of the solution (g/cm³).

The molar concentration is then calculated as follows:

$$\text{Molarity} = (\text{mass \%}) \times [\text{density (g/cm}^3)] \times (1000\text{cm}^3/\text{dm}^3) \div [\text{molar mass (g/mol)}]$$



Example 4.10

The concentration of HCl is given as 37% by mass. If the density of the solution is 1.185g/cm³, work out the molar concentration of the solution. (A_r: H = 1, Cl = 35.5).

Working out

$$\text{Molarity} = (\text{mass/mass \%}) \times [\text{density (g/cm}^3)] \times (1000\text{cm}^3/\text{dm}^3) \div (\text{molar mass (g/mol)})$$

$$\text{Mass/mass \%} = 37/100 = 0.37$$

$$\text{Density} = 1.185\text{g/cm}^3$$

$$\text{Molar mass} = (1 + 35.5)\text{ g/mol} = 36.5\text{g/mol}$$

$$\begin{aligned}\text{Therefore, molarity} &= (0.37 \times 1.185\text{g/cm}^3 \times 1000\text{cm}^3/\text{dm}^3) \div 36.5\text{g/mol} \\ &= (0.37 \times 1.185\text{g/cm}^3 \times 1000\text{cm}^3/\text{dm}^3) \times 1\text{mol}/36.5\text{g} \\ &= 12.012\text{mol/dm}^3 \\ &= 12.0 \text{ M}\end{aligned}$$

ii. Volume percentage (Volume/volume %)

Volume percentage indicates the volume of solute present in 100cm³ of solution. It is used to express concentration of a solution made by dissolving a liquid in a liquid. For example, a 70% (v/v) solution ethanol can be prepared by dissolving 70cm³ of 100% ethanol in a solution of volume 100cm³.

$$\text{Volume/volume \%} = \frac{\text{mass of solute}}{\text{volume of solute}} \times 100\%$$

Exercise 4.18

Calculate volume percentage of the following solutions

a) 5cm³ of ethanol in 20cm³ of solution.

b) 12cm³ of ethanol in 100cm³ of wine.

iii. Mass-volume percentage (mass/volume %)

Mass-volume percentage indicates the mass of solute present in 100cm³ of solution. For example, if 50cm³ of solution contains 2g of NaCl, then the mass-volume percentage of the solution is 4g/100cm³. Mass-volume % is used to express concentrations of dilute solutions.

$$\text{Mass-volume \%} = \frac{\text{mass of solute}}{\text{volume of solute}} \times 100\%$$

Exercise 4.19

1. Show that mass-volume percentage concentration of 50cm^3 of solution containing 2g of NaCl is $2\text{g}/100\text{cm}^3$.
2. 2.70g of sugar ($\text{C}_6\text{H}_{12}\text{O}_6$) was dissolved in water to make 90.00cm^3 of water. What is the concentration of the solution?
3. What is the concentration of a solution made by dissolving 2.3g of NaCl in 120g of solution?
4. What volume of a $25\text{ g}/100\text{cm}^3$ solution of sugar is needed to contain 40g of sugar?
5. How many grams of baking soda (NaHCO_3) are there in 60cm^3 of $20\text{g}/100\text{cm}^3$ solution?
($\text{A}_r: \text{Na}=23, \text{C}=12, \text{H}=1, \text{O}=16, \text{Cl}=35.5$)

4.7: Preparation of standard solutions

Meaning of standard Solutions

A standard solution is a solution of known concentration. It is prepared by dissolving a known amount of solute to make a known volume of solution. When preparing standard solutions, it is important to remember the following formulae:

- Number of moles = concentration \times volume.
- Mass = molar mass \times number of moles.

Standard solutions by either of the following two ways:

- 1) dissolving a solute;
- 2) diluting a stock solution.

1) Preparing standard solutions by dissolving a solute

This method involves weighing out a calculated mass of the solute and dissolving it to give a particular volume of solution. To prepare a standard solution by dissolving a solute, we need the following information:

- a) the molar mass of the solute;
- b) the concentration of the solution;
- c) the volume of the solution.

The mass of the solute to be dissolved can be calculated as follows:

$$\text{Mass of solute} = \text{concentration (mol/dm}^3\text{)} \times \text{volume (dm}^3\text{)} \times \text{molar mass (g/mol)}.$$

Steps for preparing standard solution by dissolving solute

1. Calculate the mass of the solute to be dissolved.
2. Weigh the calculated mass of the solute in a beaker using a beam balance.
3. Add some distilled water to the beaker to dissolve the solute.
4. Stir with a stirring rod.
5. Using a funnel, transfer the solution into a volumetric flask of the required volume.

- Using distilled water, rinse the stirring rod, beaker and funnel into the volumetric flask.
- Add distilled water to the volumetric flask to bring the solution up to the mark.
- Close the volumetric flask with a stopper and shake it gently to mix the contents.

Exercise 4.20

- How much potassium carbonate (K_2CO_3) can you dissolve to make 150cm^3 of 2M potassium carbonate solution?
- 1.25M sodium hydroxide solution was made by dissolving 20g of sodium hydroxide pellets. What was the volume of the solution?
- Describe how you can prepare 250cm^3 of a solution of sodium bicarbonate ($NaHCO_3$), of concentration 0.2mol/dm^3 .
- How can the calculated mass of the solute be weighed in a beaker?
- Suggest why it is important to do the following when preparing standard solutions by dissolving solute.
 - dissolve solute using distilled water;
 - stirring the solute in the water;
 - rinsing the stirring rod, beaker and funnel.

(A: K = 39, C = 12, O = 16, Na = 23, H = 1)

2) Preparing a standard solution by dilution with specified final volume

Dilution is the process of adding a solvent to a more concentrated stock solution to make it less concentrated. In the laboratory, some solutions are stored at high concentrations and are diluted whenever they are to be used. These solutions are called **stock solutions**. To prepare a standard solution by dilution of stock solutions, we need to have the following information:

- volume of the standard solution to be prepared;
- concentration of the standard solution to be prepared;
- concentration of the stock solution.

When preparing standard solutions by dilution, it is necessary to remember that:

number of moles (n) = concentration (c) × volume (v). Therefore, to calculate the volume of the stock solution to be diluted, the following formula is used: $c_1 v_1 = c_2 v_2$; where:

- c_1 = concentration of the stock solution;
- v_1 = volume of the stock solution to be diluted (ml or dm^3);
- c_2 = concentration of the dilute solution to be prepared;
- v_2 = volume of the dilute solution to be prepared (ml or dm^3).

After rearranging the formula, we get $v_1 = (c_2 \times v_2) \div c_1$.

Steps for preparing standard solution by diluting stock solutions

1. Calculate the volume of the stock solution to be diluted.
2. Measure the calculated volume of the stock solution using a measuring cylinder.
3. Transfer it into an appropriate volumetric flask.
4. Rinse the measuring cylinder with some distilled water and transfer the water into the volumetric flask.
5. Add more distilled water to the volumetric flask up to the mark.
6. Put the stopper and shake gently to mix the contents.

Exercise 4.21

1. Suppose you want to prepare 200ml of 0.1M ethanol, from 5M ethanol. What volume of 5M ethanol are you going to dilute?
2. 100 ml of sulphuric acid was made by diluting 25ml of 8M sulphuric acid. What was the concentration of the diluted solution?
3. Describe how you can prepare 500 cm³ of 0.2 M hydrochloric acid from 2M hydrochloric acid solution.
4. Suggest a reason why ' $c_1 V_1 = c_2 V_2$ '

4.8: Titration

Calculating concentration of a solution using a standard solution

The concentration of a solution, of unknown concentration (*solution of x conc.*), can be determined by reacting it with a standard solution (*std solution*), using a process called titration. To determine the concentration of the solution of x concentration using titration, the following information is required:

- 1) the exact volume of the solution of x conc. that reacted with the std solution.
- 2) the exact volume of the std solution that reacted with the solution of x conc.
- 3) a balanced chemical equation for the reaction between the solution of x conc. and the std solution. From the balanced equation, you need to note the following:
 - a) the coefficient of the solution of x conc.
 - b) the coefficient of the std solution.

This information can now be used calculate the concentration of analyte as follows:

Step1: Calculate the number of moles of the standard solution

The number of moles of the standard solution that reacted with the solution of unknown solution can be calculated as follows:

$$\text{Mol (std solution)} = \text{Vol (std solution)} \times \text{Conc (std solution)}$$

Step 2: Calculate the number of moles of solution of unknown solution

To find the number of moles of the solution of unknown solution:

- i. Write chemical equation for the reaction between the solution and the standard solution.
- ii. Balance the chemical equation for the reaction of the solution with the standard solution.
- iii. Note the coefficients (Coeff.) of the solution and of the standard solution in the balanced equation.
- iv. Calculate number of moles of the solution of unknown concentration

$$\text{Mol (solution of x conc.)} = \frac{\text{Coeff. (solution of x conc.)}}{\text{Coeff. (std solution)}} \times \text{mol std solution}$$

Step 3: Calculating the concentration of the solution of unknown solution

The concentration of the solution of unknown concentration can be calculated as follows:

$$\text{Conc. (solution of x conc.)} = \frac{\text{Mol (solution of x conc.)}}{\text{Vol (solution of x conc.)}}$$

Shortcut

Alternatively, after determining the volume of the standard solution and balancing the chemical equation, the three steps can be combined and the concentration of the solution of unknown concentration can be calculated as follows:

$$\text{Conc. (solution of x conc.)} = \frac{\text{Coeff. (solution of x conc.)}}{\text{Coeff. (std solution)}} \times \frac{\text{Vol (std solution)} \times \text{Conc (std solution)}}{\text{Vol (solution of x conc.)}}$$

Exercise 4.22

1. During titration, it was found that 20.0cm^3 of 0.5M of sodium hydroxide (NaOH) solution reacted with 10.0cm^3 of dilute sulphuric acid (H_2SO_4). Calculate the concentration of the H_2SO_4 .
2. 24.2cm^3 of a solution containing 0.2 mol/dm^3 of hydrochloric acid just neutralised 25.0 cm^3 of a potassium hydroxide solution. What is the concentration of this potassium hydroxide solution?
3. 25 cm^3 of ammonia solution (NH_3) were neutralised by 20 cm^3 of 1M sulphuric acid (H_2SO_4). Calculate the concentration of the ammonia solution.

Activity 4.2

Carry out the following activity and answer the questions that follow:

Materials

- 0.1M NaOH, 0.1M HCl, phenolphthalein indicator, 10cm³ syringe, beaker.

Procedure

1. Put 10 cm³ of 0.1M NaOH into a beaker.
2. Add 4 drops of phenolphthalein indicator.
3. Suck 0.2M HCl into a 10cm³ syringe.
4. Add HCl slowly **drop by drop** to the NaOH in the beaker while swirling the beaker.
5. Stop when there is a colour change of the contents in the beaker.

Discussion

- 1) What happens when an acid is added to a base?
- 2) Why did the colour of the phenolphthalein indicator change?
- 3) What volume of HCl was added to NaOH solution?
- 4) Write a **balanced** chemical equation for the reaction between NaOH and HCl.
- 5) From the balanced chemical equation, what is the ratio of moles of NaOH to the moles of HCl?

Meaning of titration

After carrying out Activity 4.2, we found that the volume of 0.2M hydrochloric acid (HCl) that is required to react with 10cm³ of 0.1M sodium hydroxide solution (NaOH) is 5cm³. The process of determining the volume of one reactant needed to react with a given volume of another reactant is known as **titration**. It is also called **volumetric analysis**. In titration, one solution is added gradually to another solution in such a way as to stop at the point when the reaction is exactly completed. This point is known as **end-point**. The end-point of the reaction is made visible by using an indicator.

Determining the concentration by titration

To determine the concentration of a solution by titration, the titrant is added **drop by drop** from a burette (or a syringe) to a known volume of the analyte in conical flask (or beaker) until the reaction is complete. The aim of carrying out titration is to determine the volume of the titrant that reacts completely with a known volume of the analyte. The volume of the analyte is worked out as follows:

- **For a burette:** volume used = final level – initial level.
- **For a syringe:** volume used = initial volume – final volume.



Experiment 4.3: Determining concentration of hydrochloric acid, of unknown concentration, by titration

Materials

I. Apparatus

- Burette (or syringe)
- Pipette (or measuring cylinder)
- Conical flask (or a beaker)
- Clamp and a clamp stand
- A dropper
- White tile or a piece of paper.

II. Chemicals required for a titration

- Hydrochloric acid
- 0.2M Na₂CO₃ solution (standard solution)
- Phenolphthalein indicator

Procedure

1. Carefully measure 10 cm³ of 0.2M Na₂CO₃ solution using a measuring cylinder or a pipette and place it in the conical flask. Record this volume as volume used in the table of results (Table 4.2)
2. Add four drops of phenolphthalein indicator to the conical flask.
3. Clamp the burette to a stand (see Figure 4.5).
4. Wash the burette with little of the acid.
5. Fill the burette with the hydrochloric acid as shown in Figure 4.5.
6. Record the initial volume of the burette in the table of results (Table 4.2). Make sure to read the bottom of the meniscus.
7. Place the conical flask containing the Na₂CO₃ solution under the burette on a white tile or piece of paper (to make the colour changes more visible).
8. Run the hydrochloric acid from the burette drop by drop into the Na₂CO₃ solution in the conical flask. Swirl the conical flask while adding the Na₂CO₃ solution.
9. Stop when the colour of the solution in the beaker has changed. This is the end point.
10. Record the final level of the acid in the burette as the final volume (Table 4.2).
11. Find the volume of the acid that has been used.

Table 4.2: Results

Reactant	V _i (cm ³)	V _f (cm ³)	Volume used (cm ³) (Volume used = V _f –V _i)
HCl			
Na ₂ CO ₃			

Calculations

1. Write balanced chemical equation for the reaction of Na_2CO_3 and HCl.
2. From the balanced chemical equation, what is the coefficient of each reactant?
3. Calculate the concentration of the analyte.

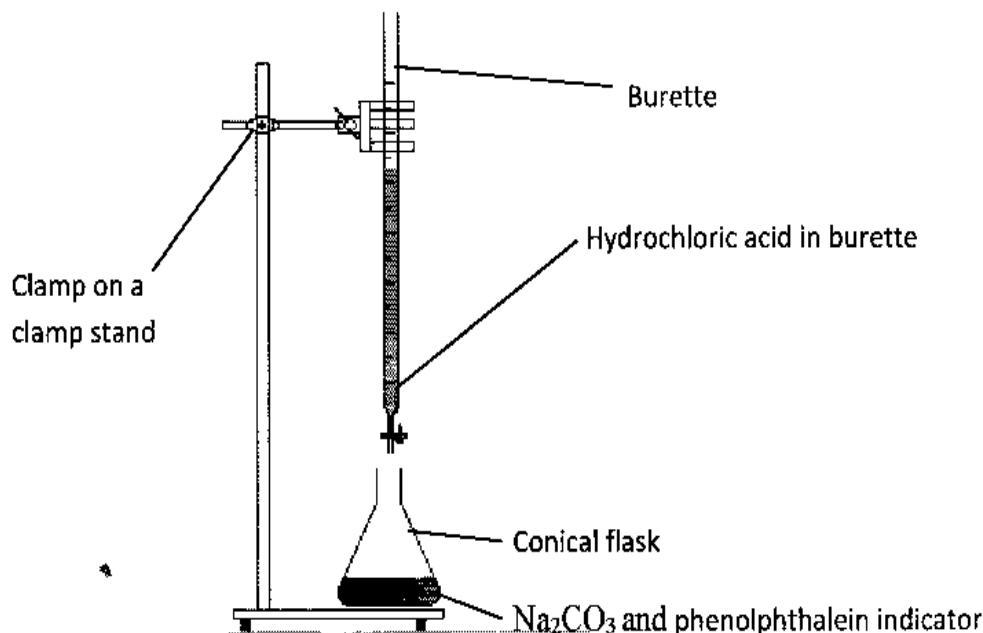


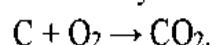
Figure 4.5: Apparatus set-up for titration

Discussion

- 1) What is the concentration of the hydrochloric acid, HCl?
- 2) What problems did you face in this experiment?
- 3) Suggest ways of avoiding the problems that you faced in this experiment.
- 4) Write a laboratory report.

4.9: Percentage yield

Reactions rarely produce predicted amounts of products from the masses of reactants used. For instance, when carbon is burned in oxygen we expect 2 moles of carbon dioxide to be produced from every 1 mole of carbon burned as the following equation indicates:



In terms of masses, if 12g of carbon are used to make CO_2 then the amount of CO_2 expected is 44g. But practically, the amount of CO_2 that is obtained is less than 44g. The problem arises because of competing reactions that happen. For example, some carbon reacts to form carbon monoxide, CO as follows: $\text{C} + \text{O}_2 \rightarrow 2\text{CO}$. The carbon that participates in this side reaction will

not be able to make CO_2 and therefore, the reaction will not yield 100% of the expected CO_2 . To be able to calculate the percentage yield we need the following information:

- A balanced chemical equation for the reaction.
- The limiting reagent.
- The theoretical (or predicted) yield.
- The actual (or experimental) yield.

Determining excess and limiting reagents

When reagents are mixed together to produce products, one reactant is used up before another runs out. The reactant that is totally consumed when the chemical reaction is complete is called **limiting reagent**. The reaction will stop when the limiting reactant is finished. The reactant that remains when the reaction has stopped is called the **excess reagent**. The excess reagent remains because there is nothing which can react with it.

Activity 4.3

Suppose to make 1 cup of tea requires 125 ml of hot water and 3 tablespoons of sugar (see Figure 4.4). If 1 litre of hot water and 18 tablespoons of sugar are available:

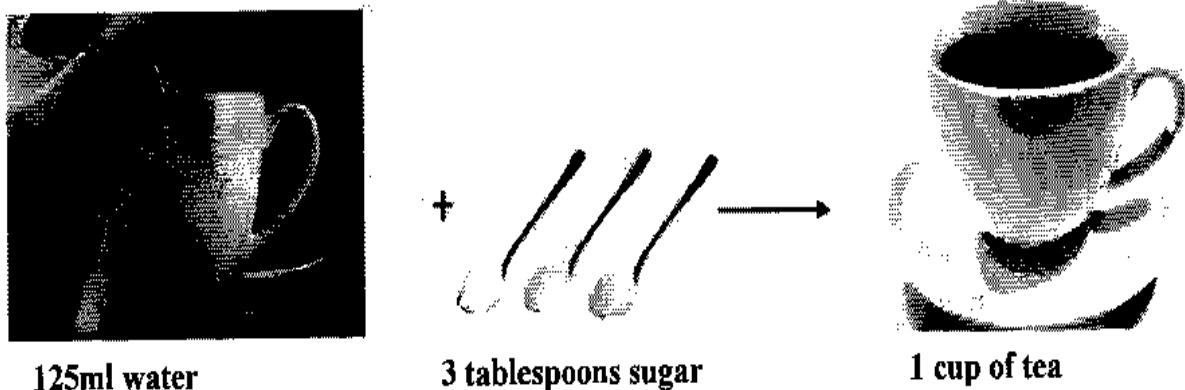


Figure 4.6: Illustration of limiting reagent

- How many cups of tea will you be able to make?
- Which one of the two, the hot water or the sugar, is the limiting reagent? Explain your answer.
- How much of the excess reagent will remain?

Determining limiting and excess reagents in a given chemical equation

In Activity 4.3, we should have discovered some of the hot water will remain when sugar is finished. We can say that, in Activity 4.3, sugar is the limiting reagent, while the hot water is the

excess reagent. For most practical purposes it is important to know how to determine limiting and the excess reagent to avoid wastage of substances. To determine the limiting reagent and excess reagent we need the following information:

- 1) balanced chemical equation for the reaction.
- 2) the amount of each reagent available.

The limiting and the excess reagent can then be calculated either of the following ways:

Method 1: comparing amounts of reactant

This method is useful when there are only two reactants. Go through Example 4.8 carefully.

Example 4.11

Calcium carbonate (CaCO_3) reacts with hydrochloric acid (HCl) to produce calcium chloride (CaCl_2), carbon dioxide (CO_2) and water (H_2O) according to the following equation:

$\text{CaCO}_3 + \text{HCl} \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$. Determine the excess reagent when 10g of CaCO_3 reacts with 50ml of 2M HCl.

(Ar: Ca = 40, C = 12, O = 16, Cl = 35.5)

Working out

Step 1: Balancing the equation

The balanced equation is $\text{CaCO}_3 + 2\text{HCl} \rightarrow \text{CaCl}_2 + \text{CO}_2 + \text{H}_2\text{O}$.

This shows that 2 moles of hydrochloric acid (HCl) is required to react with 1 mole of calcium carbonate (CaCO_3). Therefore, the mole ratio of HCl to CaCO_3 is 2 mol HCl to 1 mol CaCO_3 .

Step 2: Number of moles of each reactant available

Since the reactant amounts are given in grams, they must, first, be converted into moles for easy comparison with the reactant amounts in the balanced chemical equation.

▪ Moles of CaCO_3

Moles = mass/molar mass

Mass = 10g

Molar mass = 100g/mol

$$\text{Therefore, moles } \text{CaCO}_3 = \frac{10\text{g}}{100\text{g}} \times 1\text{mol} = 0.1 \text{ mol}$$

▪ Moles HCl

Moles = molarity x volume (dm^3)

$$\begin{aligned}\text{Volume} &= 50\text{ml} \\ &= \frac{50\text{ml}}{1000\text{ml}} \times 1\text{ml} \\ &= 0.05 \text{ dm}^3\end{aligned}$$

Concentration = 2mol/ dm^3

$$\text{Therefore, moles of HCl} = 2\text{mol}/\text{dm}^3 \times 0.05\text{dm}^3 = 0.1\text{mol}$$

Step 3: Amount of HCl required to react with 0.1 mole of CaCO₃

The amount of HCl required to react with 0.1 moles of CaCO₃ can be calculated using cross multiplication.

$$\frac{2 \text{ mol HCl}}{1 \text{ mole CaCO}_3} = \frac{x}{0.1 \text{ mol CaCO}_3}$$

$$x = \frac{0.1 \text{ mol CaCO}_3 \times 2 \text{ mol HCl}}{1 \text{ mol CaCO}_3}$$

$$= 0.2 \text{ mol HCl}$$

Therefore, 0.1 moles CaCO₃ require 0.2 moles HCl.

Step 4: Determining which reactant is an excess reagent

Moles of HCl required = 0.2

Moles of HCl available = 0.1

There is not enough HCl to react with 0.1 moles of CaCO₃. This means some of the CaCO₃ will be left, when the HCl is finished.

Therefore, CaCO₃ is the excess reagent.

Method 2: comparing product amounts that can be formed from each reagent

In this method, the balanced chemical equation is used calculate the amount of one product which can be formed from each reactant in the amount present. The limiting reactant is the one which can form the smallest amount of product considered. This method can be extended to any number of reactants more easily than the first method. Go through example 4.12 carefully.

Example 4.12

20g of iron (III) oxide (Fe₂O₃) are reacted with 8g of aluminium (Al) according to the following equation: Fe₂O₃(s) + 2Al(s) → 2Fe(l) + Al₂O₃(l). Which reactant is the limiting reagent?

(A, Fe = 160, O = 16)

Working out

Step 1: Calculating number of moles of each reactant available

$$\text{Mol Fe}_2\text{O}_3 = \frac{20\text{g}}{160\text{g/mol}} = 0.125\text{mol}$$

$$\text{Mol Al} = \frac{8\text{g}}{27\text{g/mol}} = 0.296\text{mol}$$

Step 2: Calculating amount of one product formed from each reactant

Moles of Fe formed from reactant Fe₂O₃

Let the number of moles be x.

$$\frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} = \frac{x}{0.125 \text{ mole Fe}_2\text{O}_3}$$

$$0.250 \text{ mol Fe} = x$$

Moles of iron (Fe) formed from the reactant Al

Let the number of moles be y . Then,

$$\text{Then, } \frac{2 \text{ mol Fe}}{2 \text{ mol Al}} \times \frac{y}{0.296 \text{ mol Al}}$$

$$2 \text{ mole Fe} \times \frac{0.296 \text{ mol Al}}{2 \text{ mol Al}} = y$$

$$0.296 \text{ mol Fe} = y$$

The amount of Al available can produce 0.296 mol (Fe), but the amount of Fe_2O_3 available can produce only 0.250 mole (Fe). This means the amount of Fe produced is limited by the amount of Fe_2O_3 present.

Therefore, Fe_2O_3 is the limiting reagent

Short cut

It can be seen from Example 4.9 that the amount of product (Fe) formed from each reactant (R) (Fe_2O_3 or Al) is proportional to the quantity:
$$\frac{\text{Mol (reagent R)}}{\text{Coefficient (reagent R.)}}$$

This can work for any number of reagents. Therefore, to determine the limiting reagent, just calculate this formula for each reactant, and the reactant that has the lowest value of this formula is the limiting reactant.

Exercise 4.23

1. Ammonia reacts with oxygen gas to produce nitrogen monoxide and water according to equation: $\text{NH}_3 + \text{O}_2 \rightarrow \text{NO} + \text{H}_2\text{O}$. In an experiment, 3.25g of NH_3 were allowed to react with 3.50g of O_2 :
 - a) Which reactant is a limiting reagent?
 - b) How many grams of NO are formed?
 - c) How many grams of the excess remains after the reaction?
2. If 4.95g of ethene (C_2H_4) are burned in 3.25g of oxygen (O_2):
 - a) What is the limiting reagent?
 - b) How many grams of CO_2 are formed?

(A: C = 12, O = 16, N = 14, H = 1)

Theoretical yield and percentage yield of chemical reactions

In chemistry yield is the amount of product obtained in a chemical reaction. Yield can be given as mass in grams or in moles.

Calculating theoretical and percentage yields

Theoretical yield

Theoretical yield is the amount of product based on the number of moles of all reactants present as given by the balanced chemical equation. It is the measure of how much product we expect to produce if the reaction represented by a given balanced chemical equation is perfectly efficient. To calculate the theoretical yield, we need the following information:

- The limiting reagent.
- The mole ratio of the limiting reactant with the desired product.
- Molar mass of the product.

Example 4.13

30g of Mg are mixed with 60g of Cl₂ to produce magnesium chloride. Calculate the theoretical yield of magnesium chloride. (A_r: Mg = 24, Cl = 35.5)

Working out

i. Balanced chemical equation



ii. Limiting reactant

Moles of each reactant available,

$$\text{Moles of Mg} = \frac{\text{mass of Mg}}{\text{molar mass of Mg}} \times 1\text{mole}$$
$$= \frac{30\text{ g Mg}}{24\text{ g Mg}} \times 1\text{mole} = 1.25\text{ mol}$$

$$\text{Moles of Cl}_2 = \frac{\text{mass of Cl}_2}{\text{molar mass of Cl}_2} \times 1\text{mole}$$

$$= \frac{60\text{ g Cl}_2}{71\text{ g Cl}_2} \times 1\text{mole} = 0.845\text{ mol.}$$

$$= \frac{1.25}{1} = 1.25.$$

$$= \frac{0.845}{1} = 0.845.$$

Cl₂ is the limiting reagent. This means when 0.845 moles of Cl are used up, the reaction will stop.

iii. Theoretical yield

From the balanced mole equation ($Mg + Cl_2 \rightarrow MgCl_2$)

The mole ratio of limiting reagent to desired product is 1 mol Cl_2 : 1 mol $MgCl_2$.

Therefore, the mass ratio of limiting reagent to desired product is 71g Cl_2 : 95g $MgCl_2$.

Amount of $MgCl_2$ produced when 60g of Cl are used



$$x g MgCl_2 = \frac{60g Cl_2 \times 95g MgCl_2}{71g Cl_2}$$
$$= 80.3g MgCl_2$$

Therefore, the theoretical yield when 60g of Mg reacts with 60g Cl is 80.3g $MgCl_2$.

Exercise 4.24

1. Methane, CH_4 burns in oxygen to produce carbon dioxide and water according to the following equation $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. In one experiment, 0.250 mol of methane was burned in 1.25 mol of oxygen in a sealed vessel.
 - a) Find the limiting reagent.
 - b) Calculate the theoretical yield of water (in moles).
2. Aluminium reacts with chlorine to form aluminium chloride, $AlCl_3$. If 20.0 g of Al and 30.0g of Cl are used calculate:
 - a) the limiting reactant.
 - b) the theoretical yield of $AlCl_3$.
3. Chloroform ($CHCl_3$) reacts with chlorine to produce carbon tetrachloride, CCl_4 and hydrogen chloride, HCl . In an experiment, 25g of chloroform and 25g of chlorine were mixed:
 - a) Which is the limiting reagent?
 - b) What is the maximum yield of CCl_4 in moles and in grams?

($A_r: Al = 27, C = 12, O = 16, H = 1, Cl = 35.5$)

Actual yield

Actual yield is the amount of product obtained from the actual chemical reaction. Actual yield is usually smaller than theoretical yield because of the following reasons:

1. Many reactions occur and reactants are not completely converted to the intended products.
2. Some products are lost during separation and purification of the desired product from the reaction mixture.

Calculating percentage yield

Percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100%. It indicates the percent of the theoretical yield that was obtained as the final product in an experiment. By definition,

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Example 4.14

What was the percentage yield for a reaction, if you predicted the formation of 21g C₆H₁₂ recovered only 3.8g.

Working out

$$\begin{aligned}\text{Percentage yield} &= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% \\ \text{Actual yield} &= 3.8\text{g} \\ \text{Theoretical yield} &= 21\text{g} \\ \text{Percentage yield} &= \frac{3.8\text{ g}}{21\text{ g}} \times 100\% \\ &= 18\%\end{aligned}$$

Exercise 4.25

1. If 4.95g of ethene (C₂H₄) are burned in 3.25g of oxygen (O₂) to produce carbon dioxide (CO₂):
 - a) What is the theoretical yield of CO₂?
 - b) If the actual yield of CO₂ is 2.09g, what is the percentage yield?
2. Consider the following reaction: C₆H₆ + Br₂ → C₆H₅Br + HBr:
 - a) What is the theoretical yield of C₆H₅Br if 42.2 g of C₆H₆ reacts with 73.0 g of Br₂?
 - b) If the actual yield of C₆H₅Br is 63.6 g, what is the percentage yield?

(A.: Br = 80, C = 12, O = 16, H = 1)

Topic review

1. What is the difference between molar volume and molar mass?
2. If you put equal masses of chlorine gas and sodium metal into a reaction vessel, which one will be the limiting reagent? (A_r : Cl = 35.5, Na = 23)
3. Equal masses of H₂ and O₂ are placed into a reaction container and ignited. If the reaction goes to completion, which gas is the excess reagent?
4. A certain organic compound has percentage composition of 62.1% carbon, 10.35% hydrogen and 27.6% oxygen. If the relative molecular mass of the compound is 116, calculate its molecular formula.
5. A certain organic compound has mass composition of 64.8 g carbon, 13.6g hydrogen and 21.6% oxygen. Its relative molecular mass is 74. Work out its:
 - a) empirical formula
 - b) molecular formula
6. Describe how you can prepare 500ml of 0.555M HCl from 11.3M HCl.
7. Describe how you can prepare 200ml of 0.2M NaCl solution.
8. a) What is the unit of:
 - i) mass concentration
 - ii) mass-volume percent**b)** A solution is made by dissolving 4 moles of NaCl in 2dm³ of solution; calculate its:
 - i) mass concentration
 - ii) mass-volume percent
9. Why is it important to balance chemical equations?
10. Brian made a solution by dissolving about three table-spoons of sugar in some water in a cup. Explain whether the solution he made is a standard solution or not.

Topic 5: Heats of reactions

Chemical changes and physical changes which substances undergo, involve breaking and formation of forces (or chemical bonds) between particles such as molecules, atoms and ions. The breaking and formation of the bonds involve energy in form of heat and work. As a result there are energy changes whenever a chemical or physical change has occurred. In this topic, we will study energy changes, as a result of heat transfers, that take place during physical and chemical changes.

By the end of this topic you must be able to:

1. define the terms exothermic and endothermic
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

5.1: Exothermic and endothermic reactions and processes

Chemical system and surroundings

A **system** is part of the world which we are interested in (or which we are studying). In chemistry a system is a chemical reaction such as $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$, or a chemical process such as dissolving of sodium chloride in water: $\text{NaCl}(\text{s}) + \text{water} \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$. Chemical systems consist of those molecules, atoms, ions or any combination of these particles. The rest of the world outside the system is called **surroundings**. For instance, if sodium hydroxide solution reacts with hydrochloric acid in a beaker, the chemical system is the reaction: $\text{NaOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$, and the water in which the reactants are dissolved, the wall of the beaker and the air around the beaker are the surroundings.

Energy of a system

Chemical systems possess energy due to kinetic and potential energies of their particles. The kinetic energy is due to motions of the particles. The motions are translational, rotational or vibrational.

The potential energy is due to electrostatic forces of attraction between the particles and within the particles. Potential energy is energy which is stored in the system.

The energy content of a system at constant pressure is known as **enthalpy**. It is denoted by **H**. It is the total potential energy in a system which is open to its surroundings.

Energy flow between a system and its surroundings

There is always transfer of energy between the system and its surroundings in form of **heat** whenever the system is undergoing a physical or a chemical change. Two possibilities exist concerning the direction of flow of the heat energy between the system and its surroundings:

- Heat can flow to a system from the surroundings (see Figure 5.1). This increases the enthalpy of the system.

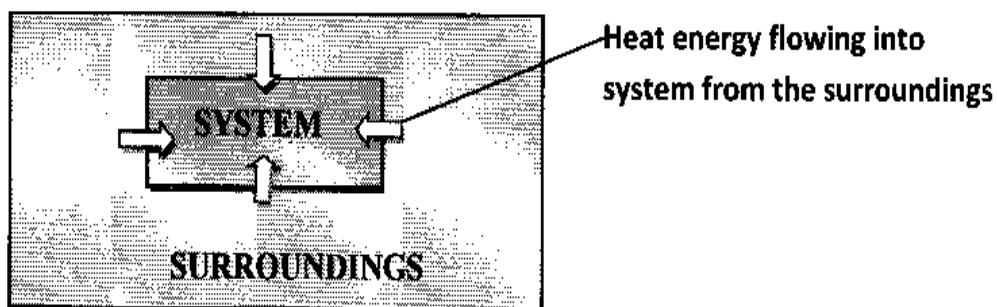


Figure 5.1: Heat energy flowing into a system

- Heat can flow from the system to the surroundings (see Figure 5.2). This decreases the enthalpy of the system.

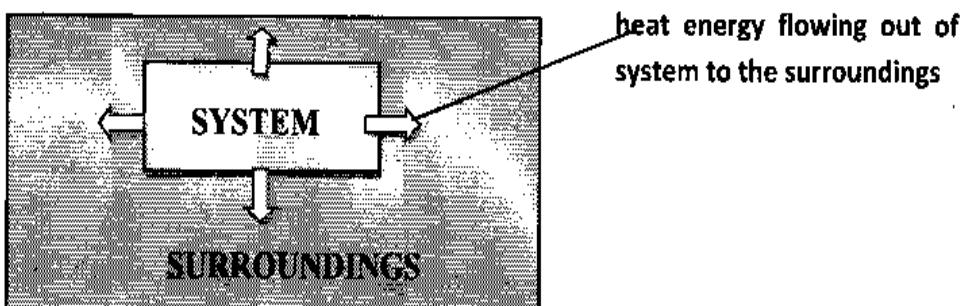


Figure 5.2: Heat energy flowing out of a system

Exercise 5.1

Suggest what happens to the surroundings when energy, in form of heat, flows:

- to the surroundings
- from the surroundings

Heats of reaction

In our discussions, we will focus on the **change in enthalpy** (or heat content) of systems when they have undergone physical or chemical changes.

The change in enthalpy of a system when reactants have reacted completely in a single reaction to form products is called **heat of reaction**. Heat of reaction is denoted by ΔH , (read as delta H). By definition, $\Delta H = \text{Enthalpy of reactants} - \text{Enthalpy of products}$.

Heat of reaction indicates amount of energy that has been transferred either into the system from the surroundings, or from the system to the surroundings per mole of reactions for a given reaction as represented by a given balanced equation. Therefore, the unit of heat of reaction is energy transferred per mole of reactions, for example, kJ/mol.

Exercise 5.2

Can the value of the heat of reaction be positive or negative? Explain.

Exothermic and endothermic changes

A change, whether physical or chemical, is described as exothermic or endothermic depending on the direction of flow of heat energy between the system and its surroundings.

1) Exothermic change

When heat energy flows out of a system to the surroundings, the change is called **exothermic**. During an exothermic change:

- i. energy content (or enthalpy) of the system decreases because energy, in form of heat, is transferred from it. This means the products have lower enthalpy than the reactants.
- ii. temperature of the surroundings increases because energy, in form of heat has been transferred to it.

Examples of exothermic changes in everyday life

Some of the changes that release heat to the surroundings include the following:

a. Burning

Burning is an exothermic reaction because the process releases heat.

b. Condensation of gases

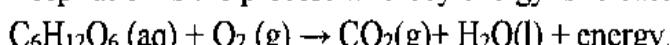
Heat is released when a gaseous substance is changing into liquid.

c. Dissolving of magnesium sulphate or calcium chloride

Dissolving of magnesium sulphate or calcium chloride is exothermic.

d. Respiration

Respiration is the process whereby energy is released from the sugar as follows:



2) Endothermic change

When heat energy flows into a system from the surroundings, the change is called **endothermic**. During endothermic change:

- i. the energy content of the system increases because the system has been transferred to it. This means that the products have higher energy content than the reactants.
- ii. the temperature of the surroundings decreases because they have lost heat.

Examples of endothermic processes in everyday life

Any process in life that requires heat for it to take place is an endothermic process. Some of the endothermic processes are as follows:

a. Cooking

Cooking of food is an endothermic process because it requires heat. For example to fry an egg, heat is absorbed from the frying pan.

b. Sweating

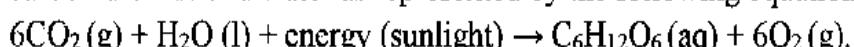
When sweating, water in the sweat takes heat from the body as it evaporates from the skin.

c. Boiling of liquids

Heat is required to make a liquid to boil.

d. Photosynthesis

Photosynthesis is the process by which green plants use energy (sunlight) to make sugar from carbon dioxide and water as represented by the following equation:



Exercise 5.3

How is each of the following processes important in our life?

- a) sweating;
- b) burning;
- c) dissolving of magnesium sulphate;
- d) respiration;
- e) photosynthesis?

Temperature changes during endothermic and exothermic processes

We have seen that the temperature of the surroundings, either decreases or increases when a system has undergone a change.

Temperature change is the difference between the temperature of the surroundings before and after a physical or a chemical change. It is denoted by ΔT (delta T).



By definition, $\Delta T = T_f - T_i$, where ΔT is change in temperature, T_f is the final temperature of the surroundings (temperature of the surroundings immediately after the physical or chemical change), and T_i is the initial temperature of the surroundings (temperature of the surroundings before the physical or chemical change).

Exercise 5.4

1. How would you use temperature change to determine if a given physical or chemical change which a system has undergone is exothermic or endothermic?
2. What would the following tell you about a given chemical or physical change:
 - a) positive value of ΔT ?
 - b) negative value of ΔT ?
3. Describe the relationship between ΔT and ΔH .
4. Explain the importance of temperature in chemistry.

Experiment 5.1: Investigating temperature changes during dissolving of sodium hydroxide and ammonium nitrate

Materials

- Water
- A thermometer
- Two test tubes
- A spatula
- Sodium hydroxide (NaOH) pellets
- Solid ammonium nitrate

Procedure

1. Pour 2cm³ of water into one test tube.
2. Measure the temperature of the water and record it as initial temperature in the table of results (see Table 5.1).
3. Add ½ spatula of NaOH into the test tube.
4. Shake the test tube to dissolve the sodium hydroxide.
5. Measure the temperature of the contents in the test tube as the sodium hydroxide is dissolving.
6. Record the extreme temperature reached as the final temperature in the table of results.
7. Repeat steps 1 to 5 with ammonium nitrate.

Table of results

Liquid in the test tube	T _i (°C)	Substance added to the test tube	T _f (°C)	ΔT (T _f – T _i) (°C)
Water		NH ₄ NO ₃		
Water		NaOH		

Discussion

- 1) Which process had:
 - a) increase in temperature?
 - b) decrease in temperature?
- 2) Explain, dissolving of which substance was:
 - a) exothermic;
 - b) endothermic?
- 3) Discuss any source of error in this experiment.
- 4) Suggest any applications of the process of dissolving of sodium hydroxide and ammonium nitrate in our life.

Experiment 5.2: Investigating temperature in neutralization reactions**Materials**

- A thermometer
- Two test-tubes
- A spatula
- A dropper
- Dilute hydrochloric acid (HCl)
- Dilute sodium hydroxide (NaOH)

Procedure

1. Pour 2cm³ of dilute HCl into one test tube.
2. Measure the temperature of the acid and record it.
3. Add five drops of dilute NaOH to the dilute HCl in the test tube.
4. Measure the temperature of the contents in the test tube as the reaction is taking place.
5. Record the extreme temperature reached.
6. Calculate the temperature change.



Discussion

- 1) What is the sign of ΔT ?
- 2) What does the sign of ΔT tell you about the overall direction of heat energy between the system and the surroundings?
- 3) Explain whether the neutralization reaction is endothermic or exothermic.
- 4) Discuss any source of error in this experiment.

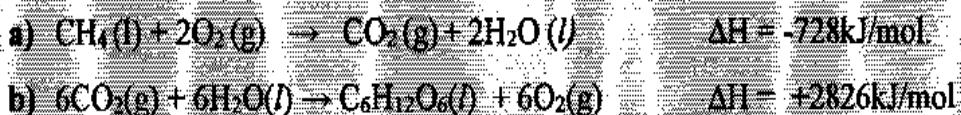
5.2: Energy level diagrams for exothermic and endothermic reactions

Thermo-chemical equations

Apart from using temperature changes to determine whether a change of a given system is exothermic or endothermic, we can also use thermo-chemical equations.

Exercise 5.5

The following equations are examples of thermo-chemical equations. Study them and answer the questions that follow.



1. What do you think is the meaning of the term ‘thermo-chemical equation’?
2. Discuss the difference between the two equations in terms of:
 - i. temperature changes during the change.
 - ii. enthalpy of reactants and products.
3. How much energy can be released to or removed from the surroundings if 5 moles of each of the reactions represented by equations a) and b) occur?

Energy level diagrams

We have learnt that enthalpy or energy content of a system changes when a chemical or physical change has occurred in the system. The products have either higher or lower heat enthalpy than the reactants after the change. The difference in enthalpies of the products and the reactants can be illustrated using diagrams called **energy level diagrams**.

Energy level diagram for an exothermic change

During an exothermic change a system loses energy. This means that after an exothermic change, products have lower enthalpy than the reactants. The value of ΔH has a negative sign.

Figure 5.3 is an energy level diagram showing the enthalpy levels of reactants and product after an exothermic change.

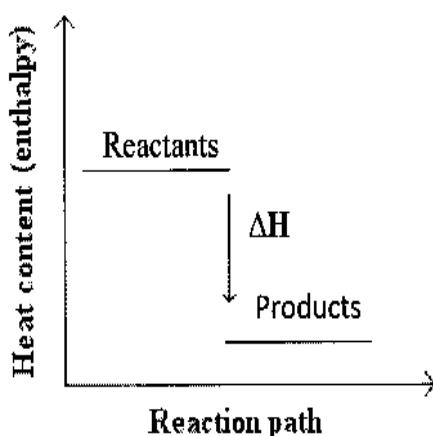


Figure 5.3: Energy level diagram for an exothermic change

Energy level diagrams for an endothermic change

After an endothermic change, products have higher enthalpy than reactants because the system loses energy. The value of ΔH has a positive sign. Figure 5.4 is an energy level diagram showing the enthalpy levels of reactants and products after an endothermic change.

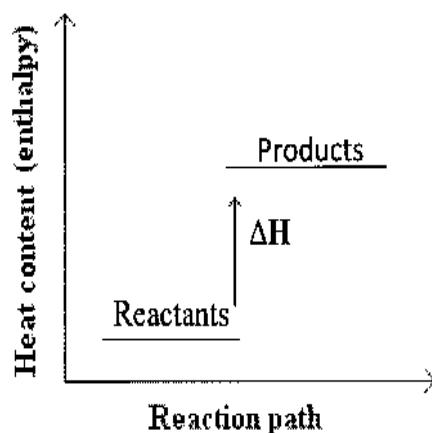


Figure 5.4: Energy level diagram for an endothermic change

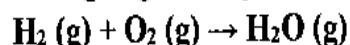
Exercise 5.6

- State whether the reactions represented by each of the following thermo-chemical equations are exothermic or endothermic. Explain your answer in each case.
 - $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \quad \Delta H = -487\text{ kJ}$
 - $\text{NaHCO}_3 + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2 \quad \Delta H = +197\text{ kJ}$
- Draw an energy level diagram for the reactions represented by each of the thermo-chemical equations in question 1 above.

5.3: Bond energy

Energy changes involved in bond breaking and bond making

All chemical changes and some physical changes involve breaking and formation of chemical bonds (or forces between particles). This means two processes are involved when a system is undergoing a change. Consider the reaction between hydrogen and oxygen to produce water:



The first process in this reaction, involves breaking existing bonds in hydrogen and oxygen molecules to produce gaseous (isolated) atoms of hydrogen and oxygen. Energy is transferred into the system during this process (see Figure 5.4).

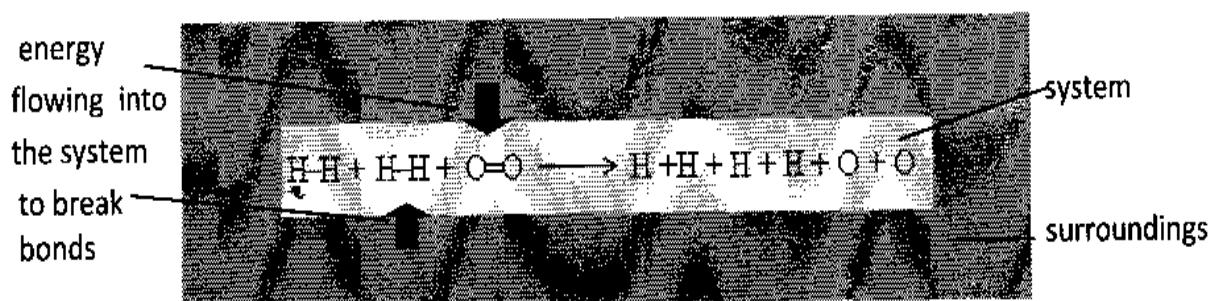


Figure 5.4: Heat energy flowing into a system when bonds are breaking

The second process involves formation of new bonds when water molecules are formed. Energy flows out of the system during this process (see Figure 5.5)

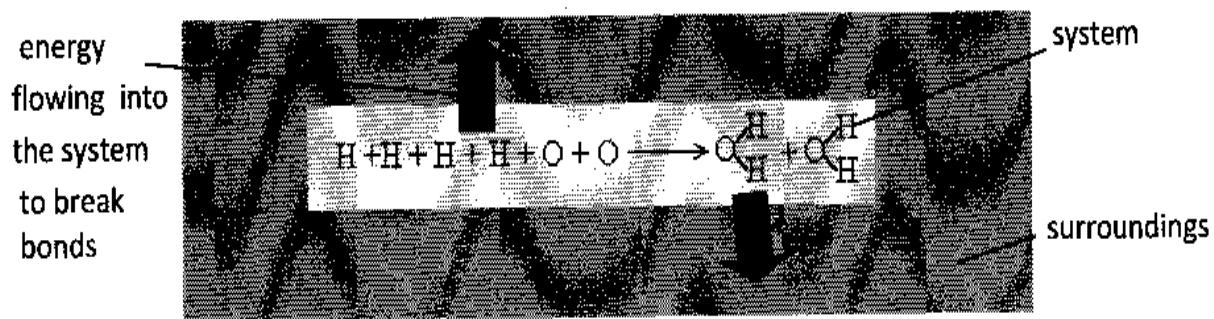


Figure 5.5: Heat energy flowing out of a system when new bonds are forming

The two processes that are involved during the reaction of hydrogen and oxygen to form water can be represented using an energy level diagram as shown in Figure 5.6.

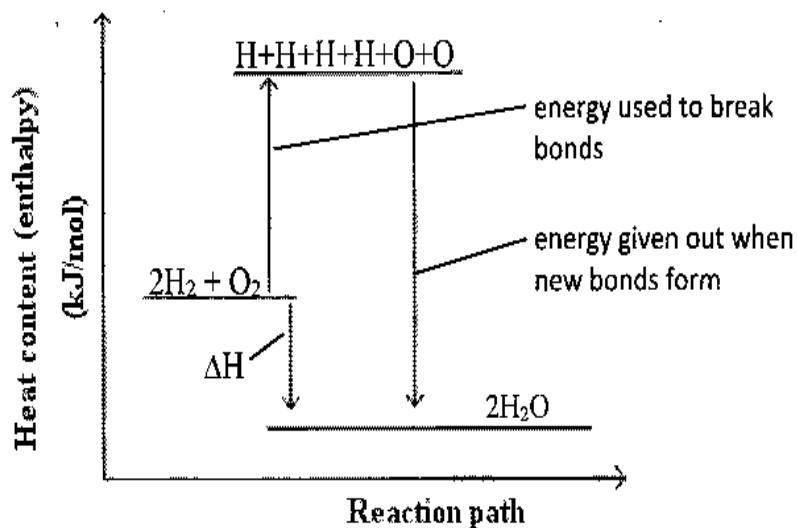


Figure 5.6: Energy level diagram for reaction of hydrogen and oxygen

Bond breaking and bond making processes take place at the same time. This means that during chemical reactions as heat energy is flowing into a system to break bonds, heat energy is also flowing out of the system when new bonds in the products are forming. The difference between the heat energy that flows into the system to break bonds and the energy that flows out when new bonds are forming gives the heat of reaction (ΔH). Therefore, heat of reaction, can also be defined as:

$$\Delta H = \text{energy used when bonds break in the reactants} - \text{energy released when new bonds form in the products}$$

Exercise 5.7

1. What would the following signs of the value of ΔH tell you about the overall direction of flow of heat energy between a system and its surroundings?
 - a) negative sign;
 - b) positive sign?
2. What would the following signs of the value of ΔH tell you about the enthalpies of reactants and products in a system?
 - c) negative sign;
 - d) positive sign?
3. How can you tell if a given physical or chemical change is exothermic or endothermic using the signs of the values of ΔH ?

Meaning of bond energy

We have seen that breaking a chemical bond (or a force of attraction between particles) requires energy. Bond energy is the energy required to break one mole chemical bonds. It is measured in kJ/mol.

The amount of energy that is required to break a bond equals the amount of energy released when the same bond is formed. Stronger bonds require more energy to break than weaker ones. It therefore follows that stronger bonds also release more energy when they form than weaker ones.

We can, therefore, conclude at this point that if a system has undergone an exothermic change, then the chemical bonds in the products are stronger than the bonds in the reactants. The opposite is also true for an endothermic change.

Table 5.1 shows bond energies for some chemical bonds. In terms of bond energies, heat of reaction is the difference between sum of bond energies of all chemical bonds broken and all chemical bonds formed when a system has undergone a change.

$$\Delta H = \frac{\text{sum of bond energies of all bonds broken}}{\text{sum of bond energies of all bonds formed}}$$

Example 5.1

Calculate the ΔH for the following reaction using the bond energies given in **Table 5.1**:



Working out

Write the equation showing the full structures of all the molecules:



$$\Delta H = (\text{sum of bond energies of all bonds broken}) - (\text{sum of bond energies of all bonds formed})$$

a) Bonds which are broken

- 2 moles of H-H bonds
- 1 mole of O=O bonds

Sum of energies of all bonds broken:

$$(2\text{mol of H-H bonds} \times \frac{436\text{ kJ}}{\text{mol of H-H bonds}}) + (1\text{mol of O=O bonds} \times \frac{498\text{ kJ}}{\text{mol of O=O bonds}}) = 1371\text{ kJ}$$

b) Bonds which are formed

- 4 moles of O-H (2 moles of O-H per 1 mole of H₂O and there are 2 moles of H₂O)

$$\text{Sum of energies of all bonds formed} = 4\text{mol of O-H bonds} \times \frac{463\text{ kJ}}{\text{mol of O-H bonds}} = 1852\text{ kJ}$$

$$\text{Therefore, } \Delta H = 1371\text{ kJ} - 1852\text{ kJ} = -481\text{ kJ}$$

The reaction is exothermic since the heat of reaction (ΔH) has a negative sign.

Table 5.1: Average bond energies (kJ/mol)

Single bonds							
C-H	413	N-H	391	O-H	463	F-E	155
C-C	348	N-N	163	O-O	146	Cl-F	253
C-N	293	N-O	201	O-F	190	Cl-Cl	242
C-O	358	N-F	272	O-Cl	203	Br-F	237
C-F	485	N-Cl	200	O-I	234	Br-Cl	218
C-Cl	328	N-Br	243	S-H	339	Br-Br	193
C-Br	276	H-H	436	S-F	327	I-Cl	208
C-I	240	H-F	567	S-Cl	253	I-Br	175
		H-Cl	431	S-Br	218	I-I	151
		H-Br	366	S-S	166		
		H-I	299				
Multiple bonds							
C=C	614	N=N	418	O=O	498	S=O	523
C≡C	839	N≡N	941			S=S	418
C=N	615*	N=O	631				
C≡N	891						
C=O	799						
C≡O	1072						

Exercise 5.8

I. Figure 5.7 shows an energy level diagram for the reaction between oxygen and nitrogen.

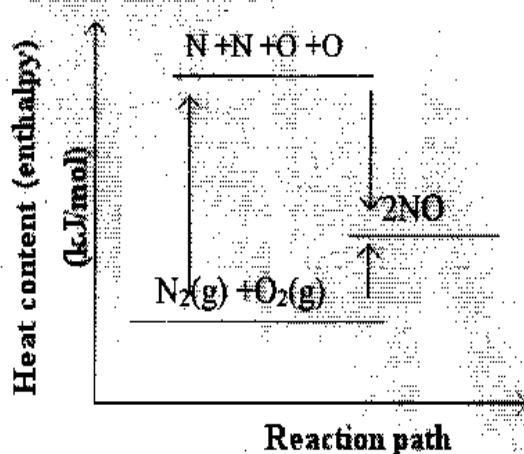


Figure 5.7: Energy level diagram for the reaction of N_2 and O_2 .

- Explain whether the reaction is exothermic or endothermic.
- Which bonds are stronger; those in the reactants or those in the products? Explain your answer.

- c) Use bond energies in Table S.1 to calculate the heat of reaction for the reaction between nitrogen and oxygen.
2. Calculate heat of reaction in each of the following reactions. In each case, explain whether the reaction is endothermic or exothermic.
- $2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}$
 - $\text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$
 - $\text{CH}_4 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \text{HCl}$
3. For each of the following equations i. to iii.
- use bond energies to determine if the reaction is endothermic or exothermic.
 - draw an energy level diagrams for each reaction.
- $2\text{NH}_3 \rightarrow \text{N}_2 + 3\text{H}_2$
 - $\text{H}_2 + \text{I}_2 \rightarrow 2\text{HI}$
 - $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$
4. Chlorine gas reacts with hydrogen gas to produce hydrogen chloride according to the following equation: $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl(g)}$
- Calculate enthalpy change (ΔH) at each stage and the overall enthalpy change of the reaction.
 - Is the reaction between chlorine and hydrogen, exothermic or endothermic? Explain your answer.
 - Draw an energy level diagram representing the two processes and the overall change that take place during the reaction.

Topic review

1. The equation for the reaction between ethanol and oxygen is as follows
 $C_2H_5OH(l) + O_2(g) \rightarrow CO_2(l) + H_2O(l) \quad \Delta H = -1371\text{ kJ/mol}$
 - a) Balance the equation.
 - b) What does ' $\Delta H = -1371\text{ kJ/mol}$ ' mean?
 - c) Is the reaction exothermic or endothermic? Give a reason.
 - d) Draw an energy level diagram for the reaction.
2. Mention one endothermic reaction and one exothermic reaction on which life depends. Explain each answer.
3. Explain the importance of the reaction and/or the process that take place in our bodies when doing sporting activities such as running.
4. With the aid of an energy level diagram describe the two processes that occur when sodium chloride is dissolving in water, given that its thermo-chemical equation is as follows:
 $NaCl(s) + \text{water} \rightarrow Na^+(aq) + Cl^-(aq) \quad \Delta H = +4\text{ kJ/mol}$
(Use bond energies in Table 5.1)

Core Element 5: Organic chemistry

Topic 6: Alkanols

Topic 7: Alkanals and alkanones

Topic 8: Alkanoic acids

Topic 9: Alkanoates

Topic 10: Identification of unknown organic compounds

Topic 6: Alkanols

Alkanols are organic compounds whose molecules are made up of carbon, hydrogen and oxygen atoms. They are common substances that are used in a number of ways in our every life. They are also called alcohols.

By the end of this topic you must be able to:

1. identify the functional groups of alkanols.
2. draw and name the structures of the first ten unbranched alkanols.
3. write the molecular formula of alkanols given the number of carbon atoms.
4. classify the alkanols as primary, secondary and tertiary.
5. describe methods of preparing alkanols.
6. explain the physical properties of alkanols.
7. describe the chemical reactions of alkanols.
8. explain uses of ethanol.

6.1: Functional group of alkanols

A functional group is a single atom or a group of atoms joined together that decides physical and chemical properties of the members of a family of organic compounds. The functional group of alkanols is the hydroxyl group (-OH).

6.2: Structures and nomenclature

Structural formulae and the skeletal formulae of alkanols

The structural and formulae of alkanols are drawn by replacing one hydrogen atom in the corresponding alkane with -OH. For example, the structural formula of an alkane with two carbon atoms is $\begin{array}{c} \text{H} & \text{H} \\ | & | \\ \text{H}-\text{C} & -\text{C}-\text{H} \\ | & | \\ \text{H} & \text{H} \end{array}$. The structure of the corresponding alkanol is $\begin{array}{c} \text{H} & \text{H} \\ | & | \\ \text{H}-\text{C} & -\text{C}-\text{OH} \\ | & | \\ \text{H} & \text{H} \end{array}$.

The same applies to the skeletal formulae; the skeletal formula of an alkane with two carbon atoms is C-C, and that of the corresponding alkanol is C-C-OH.

Exercise 6.1

Draw the structures and the skeletons of the alkanols with the following number of carbon atoms: 1, 3, 4, 5, 6, 7, 8, 9, 10.

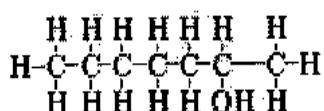
Naming Alkanols

Alkanols are named by replacing the ending -e in a corresponding alkane by -ol. For example, the corresponding alkane of an alkanol with one carbon atom is methane. Therefore, the alkanol with one carbon atom is named methanol. When naming alkanols, the carbon atom to which the functional group is attached should be indicated by a number. The following steps are followed to find a full name of an alkanol:

- Number the carbon atoms in the longest chain that contains the -OH group.
- Name the corresponding alkane.
- Modify the name of the alkane by removing the end -e and replacing it with -ol.
- If there are more than two carbon atoms in the alkanol, number the carbon atoms in the longest chain that contains the -OH group in such a way that the carbon atom to which the -OH group is attached has that lowest number possible. The number of the carbon atom to which the -OH group is attached is written before the ending -ol.

Example 6.1

What is the name of an alkanol with the following structure:

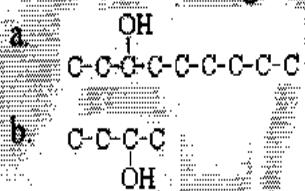


Working out

- This alkanol has 7 carbon atoms.
- The corresponding alkane is heptane.
- Replacing -e with -ol, we get heptanol.
- Counting the carbon atoms from left to right, the carbon atom with the -OH group is number 6, while counting from the right the same carbon atom is number 2:
$$\begin{array}{ccccccccc} & 7 & 6 & 5 & 4 & 3 & 2 & 1 & \leftarrow \\ & \text{C} & -\text{C} & -\text{C} & -\text{C} & -\text{C} & -\text{C} & -\text{C} & \\ \rightarrow & 1 & 2 & 3 & 4 & 5 & 6 & 7 & \text{OH} \end{array}$$
- Therefore, the correct number of the carbon is 2 and not 6.
- Hence, its name is heptan-2-ol.

Exercise 6.2

- Name the following alkanols:



- Draw the structural formula of each of the following alkanols:

- Hexan-2-ol
- Butan-1-ol

- Explain why it is not possible to have an alkanol with the name pentan-4-ol.

Molecular formula and condensed formula

A condensed formulae of an alkanol shows separate units of atoms in the alkanol and the position of the -OH group. They can be considered as shortened structural formulae. For

example, the condensed formula of ethanol is $\text{CH}_3\text{CH}_2\text{OH}$. When the number of ' CH_2 ' units is more than one, they can be grouped together. For example, the condensed formula of butanol is $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$ and it can be written as $\text{CH}_3(\text{CH}_2)_3\text{OH}$.

Molecular formula of alkanols gives the numbers and types of atoms of each kind in the alkanol except those in the functional group. For instance, the condensed formula of butanol shows that it has four carbon atoms, nine hydrogen atoms and the -OH group. Therefore, its formula is $\text{C}_4\text{H}_9\text{OH}$.

Exercise 6.3

Write the condensed formulae and molecular formulae of alkanols with the following numbers of carbon atoms: 1, 2, 3, 5, 6, 7, 8, 9, 10.

Molecular and general formulae of alkanols

Alkanols are considered as the derivatives of alkanes, where one hydrogen atom in the corresponding alkane is replaced by an -OH group.

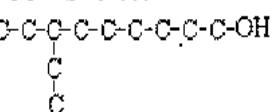
Exercise 6.4

1. In Form 2, you learnt the general formula of alkanes and alkenes.
 - a) what is the general formula of alkanes.
 - b) if one hydrogen atom in the general formula of the alkanes is replaced with an -OH group to form the general formula of alkanols, write the general formula of alkanols.
2. Write the molecular of the alkanols with
 - a) 6 carbon atoms.
 - b) 9 carbon atoms.

6.3: Classification of alkanols

Branched alkanols

Branched chain alkanols contain a branch in the carbon skeleton where three or four carbon atoms are attached to one carbon atom as follows:



The branches are called alkyl groups. Table 6.3 shows some alkyl groups.

Table 6.1: Alkyl groups

Alkyl group	Name of alkyl group
-CH ₃	Methyl
-CH ₂ CH ₃	Ethyl
-CH ₂ CH ₂ CH ₃	Propyl
-CH ₂ CH ₂ CH ₂ CH ₃	Butyl

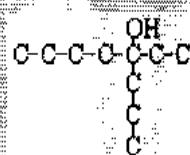
Naming branched chain alkanols

The following steps that are followed when naming branched chain alkanols are as follows:

1. Number the carbon atoms from in the longest carbon chain containing the -OH group from the end closest to the functional group. This will give the carbon atom with the functional the lowest number.
2. Identify the number of the carbon atom to which the -OH group is attached.
3. Identify the alkyl group or groups attached to the main branch.
4. Identify the number of the carbon atom to which the alkyl group is attached. Write the name of the alkyl group in front of the name of the alkanol, with the number of the carbon atom to which it is attached written in front of the name of the alkyl group.

Example 6.2

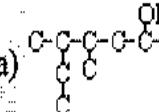
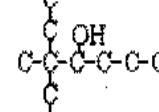
Name the following alkanols



Working out

- a) The longest carbon chain containing the -OH group has 7 carbon atoms.
 Numbering the carbon atoms in the longest carbon chain from right to left; the -OH group is at carbon number 3, while the branch (alkyl group) is attached to carbon number 6.
 The name of the alkyl group is methyl.
 Therefore, the name of the alkanol is 6-methyl heptan-3-ol.
- b) The longest carbon chain containing the -OH group has 8 carbon atoms.
 The -OH group is at carbon number 4.
 The alkyl group is attached at carbon number 4.
 The name of the alkyl group is ethyl.
 Therefore, the name of the alkanol is 4-ethyl octan-4-ol.

Exercise 6.5

1. Name the following alkanol. (a)  (b) 
2. Draw the structure of:
 - a) 3-methylbutan-2-ol.
 - b) 2-ethyl, 3-methyl pentan-1-ol.

Types of alkanols

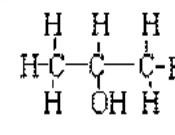
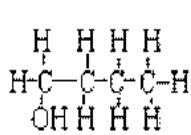
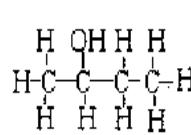
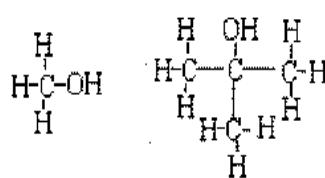
Different alkanols can have the carbon atom carrying the OH group bonded to different numbers of other carbon atoms. The numbers of carbon atoms bonded to the carbon atom with the -OH group are used to classify the alkanol as shown in Table 6.2.

Table 6.2: Classification of alkanols

Number of carbon atoms attached to the carbon atom carrying the -OH group	Alkanol type
0 or 1	Primary
2	Secondary
3	Tertiary

Exercise 6.6

Classify the following alkanols into primary, secondary and tertiary alkanols:



Unit 6.4: Preparation of alkanols

Ethanol

Ethanol is the most common member of all the alkanols. It is commonly called alcohol. Ethanol is prepared either through fermentation of sugar or hydration of ethene.

Exercise 6.7

Figure 6.1 shows the apparatus used in indigenous way of preparing ethanol. Describe in detail how ethanol is prepared in the indigenous way.

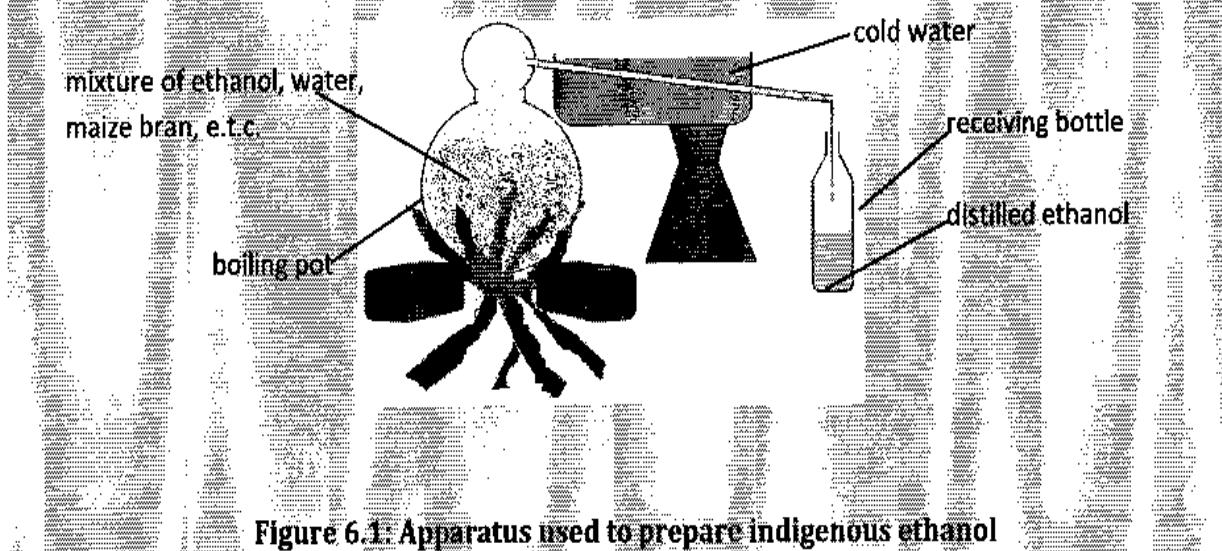


Figure 6.1: Apparatus used to prepare indigenous ethanol

*

Preparation of ethanol through fermentation of sugar by yeast

Ethanol is prepared through fermentation of sugar. The process can be speeded up by adding yeast to a mixture of sugar and water.

Experiment 6.1: Investigating preparation of ethanol through fermentation of sugar by yeast

Materials

- A conical flask (or any bottle).
- A teat pipette
- A balloon
- A conical flask (or any bottle).
- Sugar solution ($C_6H_{12}O_6$).
- Yeast
- Lime water
- Sugar solution ($C_6H_{12}O_6$).

Procedure

1. Place the sugar solution in the conical flask.
2. Add yeast to the sugar solution.
3. Cover the mouth of the conical flask with a balloon as shown in Figure 6.2.
4. Arrange the apparatus as shown in Figure 6.1.1
5. Leave the apparatus for about 3 to 4 days.
6. After three or four days, observe the balloon.
7. Collect the gas inside the balloon and bubble it through the lime water.

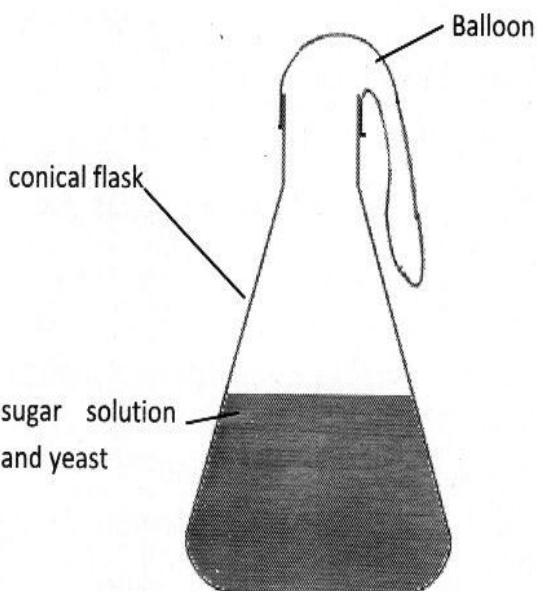


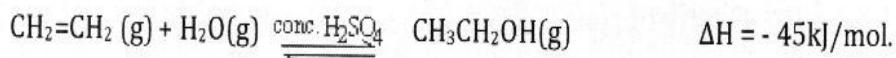
Figure 6.2: Apparatus for investigating formation of ethanol by fermentation of sugar by yeast

Conclusion

1. What is the function of the yeast in this experiment?
2. What happened to the balloon after the 3 or 4 days? Suggest a reason for your observation.
3. What happened to the lime water, when the gas was bubbled in it.
4. In Chapter 1 we discussed tests for gases. Using that information, explain your observation on the lime water in this experiment.
5. What gas was produced in this experiment? Write its chemical formula.
6. What does fermentation mean?
7. What are the products of fermentation of sugar?
8. Write a balanced chemical equation for fermentation of sugar ($C_6H_{12}O_6$).
9. How can pure ethanol be obtained from the mixture of ethanol and water that was used to dissolve the sugar?
10. Write a laboratory report.

Formation of ethanol by hydration of ethene

Ethanol can also be prepared by hydrating ethene. To do this, steam (water) is added to an alkene in the presence of a strong acid such as concentrated sulphuric acid as a catalyst. The reaction is reversible, where once ethanol is produced, it decomposes back into ethene and water as represented by the equation below:



Formation of ethanol by dehydration of ethene is an exothermic reaction.

Exercise 6.8

1. Hydration of ethene to produce ethanol is a reversible reaction, where there is a forward reaction and a reverse reaction. What is the product(s) in the:
 - a) forward reaction?
 - b) reverse reaction?
2. Hydration of ethene is addition of water to ethene to produce ethanol; what name would be given to the reverse reaction where water is removed from alkanols molecules?
3. If the forward reaction is exothermic, explain whether the reverse reaction will be exothermic or endothermic.
4. What will be the effect of each of the following actions on the formation of ethanol by hydration of ethene?
 - a) increasing the temperature of the reaction environment?
 - b) decreasing the temperature of the reaction environment?
5. If you want to increase the amount of ethanol produced from a given hydration reaction of ethene, do you have to conduct the reaction in hot or cold environment? Explain your answer.

6.5: Physical and chemical properties of alkanols

Physical properties of alkanols

Polarity of alkanols

Alkanols are polar compounds and in addition to that, they contain oxygen and hydrogen atoms. Because of this, the molecules of alkanols join together by strong intermolecular forces.

Exercise 6.9

In Chapter 2 we learnt about polar molecules and different types of intermolecular forces. Use that information to discuss the following questions:

- 1) Explain why alkanols are polar by referring to the structures of their molecules.
- 2) What types of intermolecular forces are found in alkanols?
- 3) Which intermolecular force becomes stronger with increase in molecular size?
- 4) Explain the effect of the polarity of the alkanols and the types of the intermolecular forces that exist between their molecules on the following properties of the alkanols:
 - a) melting points and boiling points?
 - b) solubility in water?
 - c) volatility?
 - d) viscosity?
 - e) density?

Experiment 6.2: Investigating the effect of the size of the molecules of alkanols on their solubility in water.

Materials

- water, ethanol, butanol, octanol, 3 test tubes

Procedure

1. Pour 3 cm³ of water into a test tube.
2. Add 3 drops of ethanol to the test tube
3. Shake the test tube and wait for 2 minutes
4. Record your observation
5. Repeat steps (1) to (4) with butanol and octanol respectively.

Table of results

Liquid added to water	Observation
Ethanol	
Butanol	
Octanol	

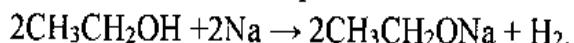
Conclusion

1. How many layers did you observe in each test tube?
2. What does the number of layers in the test tube tell you about the solubility of the alkanol in water?
3. What happens to the solubility of the alkanols as the size of the molecules increase?
4. What can you conclude from the results of the experiment?
5. Write a laboratory report.

Chemical properties of alkanols

1. Alkanols react with alkali metals

Alkanols react with alkali metals to produce organic compounds (metal alkoxides). For example, ethanol reacts with sodium to produce sodium ethoxide as represented by the following equation:



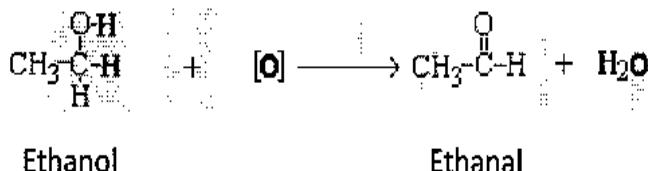
2. Alkanols undergo oxidation reactions to produce alkanals, alkanones or alkanoic acids

A substance is said to have been oxidised when its molecules accept oxygen atoms or when it loses a hydrogen atom. Alkanols can be oxidised when they react with a solution of sodium dichromate (IV), (Na₂Cr₂O₇) or potassium dichromate (IV), (K₂Cr₂O₇) acidified with dilute sulphuric acid or air. Full equation for this reaction is complicated. To simplify it, the oxidising agent is represented by the oxygen atom only, which is written as [O]. [O] stands for an oxygen atom from an oxidising agent. During the reaction the oxygen atom from the oxidising agent removes a hydrogen atom from the -OH group, and another hydrogen atom from the carbon

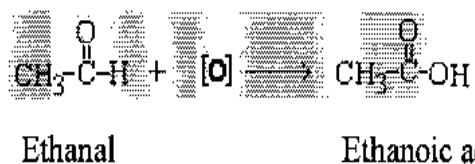
atom to which the $-OH$ group is attached to form a water molecule. Different types of alkanols produce different organic compounds of different families.

A. Oxidation of primary alkanols

Primary alkanols produce alkanals (see Topic 7) when they undergo oxidation reaction as follows:



Ethanal belongs to the family of alkanals. If enough oxygen is available, ethanal can be oxidised further to ethanoic acid, which is a member of alkanoic acids (see chapter 8).



B. Oxidation of secondary alkanols

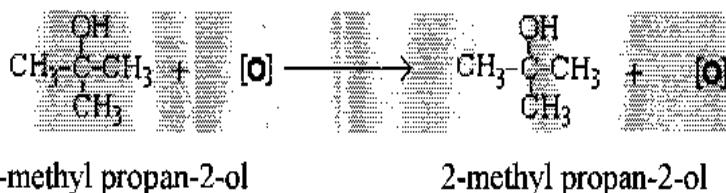
Secondary alkanols produce alkanones (Topic 7) when they undergo oxidation reaction.



Propanone is a member of alkanones.

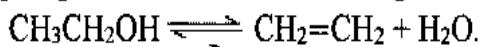
C. Tertiary alkanols

Tertiary alkanols do not undergo oxidation reaction as shown by the following equation:



3. Alkanols undergo dehydration reactions to produce ethene

Ethanol can be dehydrated to ethene when it is heated at about 85°C with sulphuric acid or phosphoric acid as the following reaction equation indicates:



4 Alkanols undergo combustion reaction

Alkanols can burn in air to produce carbon dioxide and water. For example, ethanol can burn as follows: $\text{C}_2\text{H}_5\text{OH(l)} + 3\text{O}_2\text{(l)} \rightarrow 2\text{CO}_2\text{(g)} + 3\text{H}_2\text{O(g)}$.



Exercise 6.10

1. Suggest a reason why 2-methyl propan-2-ol cannot be oxidised.
2. Write the equation for the oxidation (where possible) of each of the following alkanols:
 - a) methanol
 - b) butan-2-ol
 - c) hexan-3-ol
 - d) 3-methyl pentan-2-ol
 - e) 3-ethyl pentan-1-ol
3. Suggest the reason why dehydration of ethanol to ethene takes place at a higher temperature such as 85°C ?

6.6: Uses of ethanol

Ethanol has many applications in our every day life. Some of the uses of ethanol are as follows:

1) As a solvent

- Industrial methylated spirit is used in industries as a solvent.
- Ethanol is used as a solvent for perfumes.

2) As alcoholic beverages

Alcoholic beverages such as Carlesberg and Malawi gin contain 3% to 40% ethanol by volume. These are consumed by human beings for various reasons.

3) As a raw material in industries

Ethanol is used as raw material in the production of vinegar and yeast.

4) As medicine

Ethanol is used in the processing of antibiotics, vaccines, tablets, pills and vitamins.

5) As Antiseptic

Many cleaning products contain high volumes of industrial alcohol.

6) As fuel

Ethanol can be used as fuel as follows:

a) In car engines

When mixed with petrol, bioethanol is used in car engines.

b) As a source of heat

Bioethanol is used in spirit burners to heat substances such as water.

Dangers of excessive consumption of alcoholic drinks and medicines on human health

Excessive consumption of alcoholic drinks and medicines is dangerous; it can affect your health and ability to function and think.

A. Dangers of excessive consumption of alcoholic drinks

Excessive consumption of alcoholic drinks means drinking too much on a single occasion or over a long time time. This can have serious effects on one's health as follows:

1. Spread of HIV/AIDS

Excessive consumption of alcoholic drinks promotes the spread of HIV/AIDS, and increases illnesses and deaths in people with the virus as follows:

- a) Alcohol consumption takes place in bars and bottle stores, where people meet potential sex partners and engage in high-risk sexual behaviours such as multiple sexual partners, unprotected sexual intercourse and sex with high-risk partners such as prostitutes.
- b) Alcohol consumption increases the rate at which the virus attacks the lungs. It also increases susceptibility to some infections that occur as complications of AIDS such as pneumonia, tuberculosis (tb) and hepatitis C.
- c) Alcohol increases the rate at which the virus grows. This leads to higher amounts of the virus in the body.
- d) Alcohol interferes with ART (Antiretroviral therapy) medications for blocking the effectiveness of HIV/AIDS.
- e) Excessive consumption of alcohol leads to heart disease in HIV-infected people.

2. Heart damage

Drinking a lot on a single occasion or too much over a long time can damage the heart, thereby causing many problems such as the following:

- a) Stretching and drooping of the heart
- b) Irregular heart beat
- c) Stroke
- d) High blood pressure

3. Liver diseases

Excessive consumption of alcohol can affect the liver, and can lead to a variety of liver diseases such as:

- a) Fatty liver
 - This is a build-up of extra fat in liver cells.
 - Symptoms are: fatigue, weakness, and weight loss.
- b) Alcoholic hepatitis [Figure 6.3(b)]

Alcoholic hepatitis is the swelling and damaging of the liver. Symptoms of alcohol hepatitis include the following: loss of appetite, vomiting, abdominal pain, fever, and jaundice. If it is severe, it leads to serious complications such as liver failure and death.

- c) Alcoholic cirrhosis [Figure 6.3(c)]

Alcoholic cirrhosis is the scarring of the liver; hard scar replaces the soft healthy tissue. It is the most serious type of alcohol-related liver diseases. The symptoms of alcoholic cirrhosis are similar to those of alcoholic hepatitis.

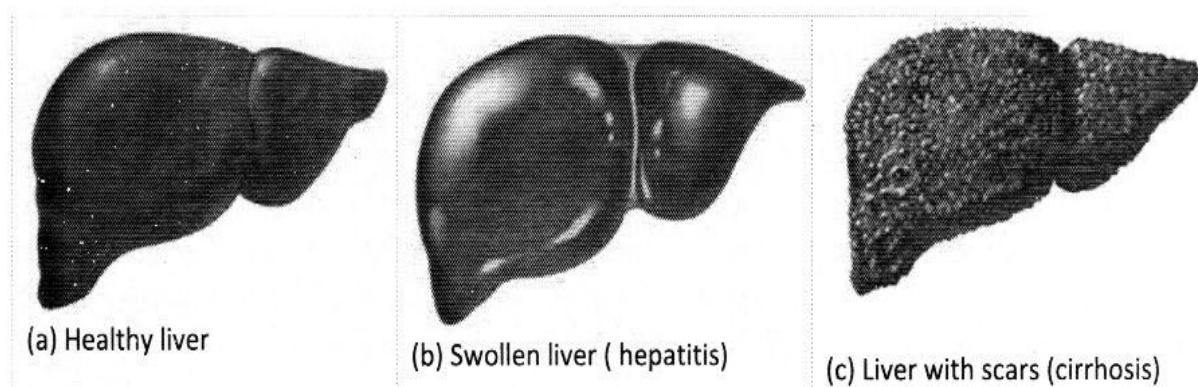


Figure 6.3: Effects of excessive consumption of alcoholic drinks on the liver

4. Inflammation and swelling of the blood vessels in the pancreas

Alcohol causes the pancreas to produce toxic substances that can lead to pancreatitis; which is the inflammation and swelling of the blood vessels in the pancreas that prevent proper digestion.

5. Cancers

Drinking too much alcohol increases the risk of developing certain cancers such as the cancer of the:

- a) mouth.
 - b) oesophagus.
 - c) throat.
 - d) liver.
 - e) breast.

6. Weakening of immune system

Drinking too much alcohol can weaken one's immune system, thereby making the body susceptible to various kinds of diseases. For example, people who consume alcoholic drinks excessively are more liable to contract diseases like pneumonia and tuberculosis (tb) than people who do not. In addition, drinking too much alcohol on a single occasion reduces the body's ability to fight infections even up to 24 hours after getting drunk.

7. Infertility

Excessive consumption of alcohol in late teens and early twenties can lead to serious infertility problems later in life in both men and women (<http://www.drinkaware.co.uk>). For example:

- a) In men, it can cause impotence and infertility. This is because excessive consumption of alcohol lowers the levels of testosterone and sperm quality and quantity.
- b) In women excessive consumption of alcoholic drinks may have the following effects:
 - i. stopping having menstruation periods or having an early menopause.
 - ii. miscarrying if they fall pregnant.

A. Dangers of excessive consumption of medicines on human health

There are two types of medicines: over the counter (O.T.C) and prescription medicines. **Over the counter medicines** are the medicines that can be bought in shops or pharmacies, while **prescription medicines** are the ones that are prescribed by doctors for patients at the hospital. Medicines are safe and effective when taken as directed, but can have negative effects on your health if taken excessively. Medicines are said to have been consumed excessively if one takes too many medicines than recommended or when one takes medicine he or she is not ill. Some of the dangers of excessive consumption of medicines include the following:

- 1) Impaired judgement, seizures, blurred vision, shallow breathing and a fast heart rate, high blood pressure, heart damage and permanent brain damage. These effects result when one takes large doses of cough and cold medicines which contain dextromethorphan.
- 2) Liver failure, stomach bleeding, kidney failure and cardiac risks.
These effects result when one takes large doses of pain relievers (pain killers) such as ibuprofen and acetaminophen.
- 3) Serious dehydrations, panic attacks and heart irregularities that sometimes lead to sudden deaths. These result from large doses of caffeine medicines such as pain relievers with caffeine and energy drinks.
- 4) Heart palpitations, irregular heartbeats and heart attacks which result from excessive consumption of cold medicines that contain pseudoephedrine.
- 5) Resistance to antimicrobial medicines.
Taking antimicrobial medicines for non-bacterial infections leads to antimicrobial resistance. For example, taking penicillin to treat malaria causes body's resistance to penicillin.

Exercise 6.11

1. Suppose that you have realised that some boys and girls in your class are secretly taking alcoholic drinks. What advice can you give them with regard to:
 - a) the prevalence of the HIV in our community?
 - b) their future reproductive health?
 - c) their general health?
2. Your friend is fond of taking aspirin every time before he or she starts her study
 - a) What health risk is this friend of yours taking?
 - b) What can be your advice to him or her?

Topic review

- 1.** Describe four chemical properties of alkanols.
- 2.** Name the product that will be formed when each of the following alkanols is mixed with acidified potassium dichromate.
 - a) butan-2-ol;
 - b) butan-1-ol;
 - c) 3-methyl pentan-1-ol;
 - d) 3-methyl pentan-3-ol.
- 3.** Explain the effect of increase in molecular sizes on the following properties of alkanols:
 - a) boiling points;
 - b) solubility in water;
 - c) density;
 - d) volatility;
- 4.** What are the effects of excessive consumption of alcohol on the:
 - a) liver?
 - b) heart?
- 5.** Explain the effects of excessive consumption of alcohol in late teens and early twenties on the fertility of men and women.

Topic 7: Alkanals and alkanones

Under mild oxidation conditions, it is possible to convert alcohols to other families of organic compounds such as alkanals and alkanones. Alkanals are also called aldehydes while alkanones are also called ketones. In this topic we are going to look at the alkanals and the alkanones.

By the end of this topic you must 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 from alkanones.
5. describe the uses of alkanals and alkanones.

7.1: Functional group of alkanals and alkanones

Alkanals and alkanones are organic compounds whose functional group is the carbonyl group,

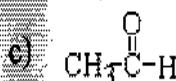
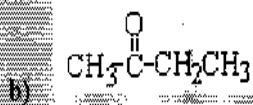
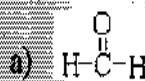


Structural difference between Alkanals and alkanones

In alkanals at least one hydrogen atom is bonded to the carbon in the carbonyl group. In alkanones, the carbon atom in the carbonyl group is bonded to two other carbon atoms. Alkanones do not have a hydrogen atom attached to the carbonyl group.

Exercise 7.1

Consider the following structures of alkanals and alkanones:



1. Describe the structural differences between the structures.

2. Which one(s) of these structures are structures of:

- i. alkanals?
- ii. alkanones?

7.2: Structures of alkanals and alkanones

The general structure of the alkanals is $\text{R}-\overset{\text{O}}{\underset{\text{H}}{\text{C}}}-\text{H}$, and the general structure of the alkanones is

$\text{R}_1-\overset{\text{O}}{\underset{\text{H}}{\text{C}}}-\text{R}_2$, where R, R¹ and R² stand for the alkyl groups such as -CH₃, -CH₂CH₃ and

-CH₂CH₂CH₃. For example, the structure of an alkanal with three carbon atoms is $\text{CH}_3-\text{CH}_2-\overset{\text{O}}{\underset{\text{H}}{\text{C}}}-\text{H}$.

Exercise 7.2

1. Draw the structures of alkanals and alkanones with the following numbers of carbon atoms:

- a) 1
- b) 2
- c) 3
- d) 4
- e) 5

2. How many carbon atoms do the following have:

- a) First alkanal?
- b) First alkanone?

Molecular formulae of alkanals and alkanones

The general formula of alkanals and alkanones is C_nH_{2n}O. Therefore, given the number of carbon atoms in the alkanal or the alkanone, one can use the general formula to come up with the molecular formula of the alkanal or the alkanone.

Exercise 7.3

Write the molecular formulae of:

a) alkanals with:

- i. n=1
- ii. n=2
- iii. n=3
- iv. n=4
- v. n=5

b) alkanones with:

- i. n=4
- ii. n=5

Naming straight-chain alkanones

Names of alkanones end with the suffix '-one'. To come up with the name of an alkanone follow these steps:

1. Number the carbon atoms in the longest chain that contain the C=O group.

2. Name the corresponding alkane.
3. Modify the name of the alkane by removing the end ‘-e’ and replacing it with ‘-one’.
4. If there are 4 or more carbon atoms in the longest carbon-chain of the alkanone, a number which indicates the carbon atom to which the oxygen atom, O, is bonded is placed before the suffix ‘-one’. The number is found by numbering the carbon atoms in the longest carbon chain of the alkanone in such a way that the carbon atom to which the oxygen atom is attached has that lowest number possible.

Example

Name the following alkanone: $\text{CH}_3\text{CH}_2\overset{\text{O}}{\underset{\parallel}{\text{C}}}\text{CH}_3$

Working out

1. The number of carbon atoms in the longest carbon chain of the alkanone is 4.
2. The corresponding alkane is butane.
3. Replacing -e with -one, we have butanone.
4. The oxygen atom, ‘O’ is attached to carbon number 2.
5. Therefore, the name is butan-2-one.

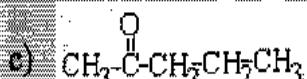
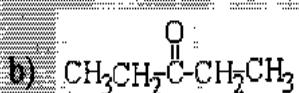
Naming alkanals

The names of alkanals end with the suffix ‘-al’. The names can be found by following these steps:

1. Number the carbon atoms in the longest chain that contain the C=O group.
2. Name the corresponding alkane.
3. Modify the name of the alkane by replacing the -e in the name of the alkane with ‘-al’.
4. In alkanals, the oxygen atom, O, is always bonded to carbon number 1. So there is no need to indicate the number of the carbon to which the O is attached.

Exercise 7.4

1. Name the following organic compounds:



2. Draw the structure of:

- propanone
- propanal

7.3: Sources of alkanals and alkanones

Alkanals and alkanones are produced by the oxidation reactions of alkanols.

Experiment 7.1: Preparing alkanals and alkanones

Materials

- Two 25ml measuring cylinders
- Three 25 ml beakers
- Potassium dichromate (IV) solution ($K_2Cr_2O_7$)
- Sulphuric acid
- Ethanol
- Propan-1-ol
- Propan-2-ol

Procedure

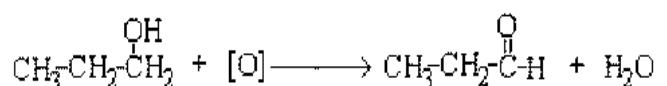
1. Pour about 2cm^3 of potassium dichromate (IV) solution into a measuring cylinder.
2. Add about 1cm^3 of sulphuric acid and stir with a glass rod.
3. Add 10 drops of the acidified potassium dichromate (IV) solution into each of three beakers
4. Add three drops of ethanol to one of the three beakers.
5. Repeat steps 4 and 5 with propan-1-ol and propan-2-ol in the other beakers.
6. Observe for 15 minutes.
7. Record your observations.

Discussion

- 1) What was the colour of the acidified potassium dichromate (IV) solution?
- 2) What was the colour of the acidified dichromate (IV) in each beaker after 15 minutes?
- 3) What conclusion can you draw from the results of this experiment?
- 4) Write a laboratory report.

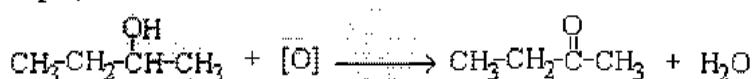
Alkanals

Alkanals are produced from the reaction of oxidising agent with excess of primary alkanols. For example, propanal is produced by oxidising propanol as shown by the following equation:



Alkanones

Alkanones are produced from the reaction of secondary alkanols with an oxidising agent. For example, butan-2-ol can be oxidised to butanone as follows:



Exercise 7.5

1. Name the product of the oxidation of the following alkanols:
 - a. propan-2-ol.
 - b. butan-1-ol.
2. Which alkanol is oxidized to give the following products?
 - a. ethanal?
 - b. pentan-3-one?
3. Suggest a reason why tertiary alkanols cannot be oxidized to alkanones or alkanals.

Properties of alkanones and alkanals

Polarity of -C=O bond

The carbon atom and the oxygen atom in the carbonyl group, (-C=O) have different electronegativities. Therefore, the -C=O bond is polar. However, molecules of alkanals and alkanones cannot form hydrogen bond amongst themselves, but can form hydrogen bonds with other substances that have hydrogen bonds such as alkanols and water.

The physical properties of alkanones and alkanals

The polarity of the C=O bond affects the physical properties of the alkanals and the alkanones as follows:

1. Alkanals and alkanones have higher melting points and high boiling points than the corresponding alkanes. This is because the C=O bond is polar.
2. Alkanals and alkanones have lower melting points and boiling points than the corresponding alkanols. This is because there are no hydrogen bonds between the alkanal or the alkanone molecules.

Trends in melting points and boiling points and solubility

1. Both melting points and boiling points of alkanals and alkanones increase as the molecules get bigger. This is because melting points depend on the strengths of the intermolecular forces. As the molecules get bigger these forces become stronger.

2. Solubility of in water alkanals and alkanones decrease with increase in molecular sizes.

Small alkanals and alkanones are able to dissolve in water because they are able to form hydrogen bonds with water molecules. But the solubility decreases as the sizes of the molecules get bigger. The reason for this is that, as the carbon chain becomes longer in the alkanals and alkanones, the induced dipole forces in the hydrocarbon tail become stronger than the hydrogen bonds. As a result, the hydrogen bonds fail to disrupt them. This makes the alkanal and the alkanone molecules remain joined to each other.

Chemical properties of alkanals and alkanones

1. Alkanals undergo oxidation reactions

Alkanals are oxidized to carboxylic acid when they are mixed with oxidizing agents such as acidified potassium dichromate (IV) ($K_2Cr_2O_7$) or with potassium permanganate ($KMnO_4$) as represented by the following equation:



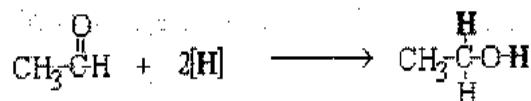
When the reaction has taken place:

- the acidified potassium dichromate (IV) solution changes its colour from orange to green.
- the solution of potassium permanganate changes its colour from purple to colourless.

Alkanones are not readily oxidised because they do not contain a hydrogen atom bonded to the carbonyl group

2. Alkanals and alkanones undergo reduction reactions

Alkanals and alkanones are reduced to alkanols when they are reacted with reducing agents such as lithium aluminium hydride ($LiAlH_4$) or sodium boron hydride ($NaBH_4$). The reactions are complicated and we use only the simplified equation where the reducing agent is represented by $[H]$. During the reaction, two hydrogen atoms are added to the alkanal or the alkanone; one hydrogen atom is added to the oxygen atom and the other hydrogen atom is added to the carbon atom in the functional group as the following equation represents:



Exercise 7.6

- Write a chemical equation that represents the oxidation of the following compounds where the reaction will take place. Where the reaction is not possible, explain why it is not possible.
 - propanal.
 - propanone.
 - pentan-2-one.
 - butanal.
- For the reduction of each of the following organic compounds:
 - Name the alkanol that is formed.
 - Name the class of the alkanol that is formed.
 - propanal.
 - butanone.
 - butanal.
 - pentan-1-one.
 - pentan-3-one.

7.4: Tests to distinguish alkanals from alkanones

Experiment 7.2: Distinguishing alkanals and alkanones using acidified potassium (IV) dichromate.

Materials

- An alkanal in a test tube
- An alkanone in a test tube
- Acidified potassium dichromate(IV)
- Warm water

Procedure

1. Add a few drops of acidified potassium dichromate (IV) solution to an alkanal in a test tube.
2. Repeat step 1 with an alkanone.
3. Observe for three minutes.
4. If nothing happens in the cold, warm the mixture in a beaker of hot water.
5. Record your observations

Discussions.

1. What was the colour change of the acidified potassium dichromate(IV) solution in the:
 - a) alkanal?
 - b) alkanone?
2. What conclusion can you make from this experiment on how alkanals can be distinguished from the alkanones?

Distinguishing alkanals and alkanones using tests

Alkanals and alkanones can be distinguished using different tests as follows:

a) The Brady's test

Experiment 7.3: Distinguishing alkanals and alkanones from other liquids using the Brady's test

Materials

- | | |
|-------------------|---------------|
| • Test-tubes | • An alkanone |
| • Brady's reagent | • Ethanol |
| • An alkanal | • Hexane |

Procedure

1. Add a few drops of Brady's reagent into a test tube.
2. Add a few drops of the alkanal.
3. Record your observations.
4. Repeat steps 1 to 3 with the alkanone, ethanol and hexane.

Conclusion

- 1) What did you observe in each test tube?
- 2) Which substances reacted with the Brady's reagent?
- 3) How can you use the results of this experiment to test for alkanals and alkanones?

b) The Tollen's test (silver mirror test)

Experiment 7.4: Distinguishing alkanals from alkanones using the Tollen's test

Materials

- Three test-tubes
- Freshly prepared Tollen's reagent
- An alkanal
- An alkanone

Procedure

1. Prepare Tollen's reagent as follows:
 - a. Dissolve silver (I) nitrate in distilled water.
 - b. Add a drop of sodium hydroxide solution to give precipitate of silver (I) oxide.
 - c. Add just enough dilute ammonia solution to dissolve the precipitate
2. Add a few drops of the freshly prepared Tollen's reagent into one of the three test tubes
3. Add a few drops of the alkanal to the Tollen's reagent in the test tube.
4. Warm the mixture gently in a hot water bath for a few minutes.
5. Record the results.
6. Repeat steps 1 to 4 with the alkanone.

Discussion

- 1) What was the colour of the Tollen's solution before adding any substance?
- 2) What colour change did you observe when the substances were added to the Tollen's reagent into each test tube?
- 3) What conclusion can you draw from the results of this experiment?

c) The Fehling's test

Experiment 7.5: Distinguishing alkanals from alkanones using the Fehling's test

Materials

- Three test tubes
- Fehling's reagent (Benedict's reagent can also be used)
- An alkanal
- An alkanone



Procedure

1. Pour some of the Fehling's reagent (blue in colour) into one test tube.
2. Add a few drops of the alkanal to the Fehling's reagent in the test tube.
3. Warm the mixture gently in a hot water bath for a few minutes.
4. Record your observations.
5. Repeat steps 1 to 4 with the alkanone.

Discussion

- 1) What colour change did you observe when each of the following was added to the Fehling's solution:
 - a) Alkanal?
 - b) Alkanone?
- 2) How can you use this experiment to distinguish between alkanals and alkanones?

7.5: Uses of alkanals and alkanones

Alkanals and alkanones are used in a number of ways in our everyday life. Some of the uses are as follows:

1. Synthesising other organic compounds

For example ethanal is used to make ethanoic acid.

2. Making plastics

Methanal is used to make Bakelite while propanone is used to make Perspex.

3. As solvents

Propanone is used as a solvent in both laboratory and industry. For example, propanone is used as a solvent for cellulose nitrate and cellulose ethanoate.

4. As preservatives

Formalin solution which is used to preserve animal specimens contains 40% methanal.

5. As food flavours

Benzaldehyde, which has a characteristic odour, is used to flavour foods.

Topic review

1. What is the general formula for alkanals and alkanones?
2. Draw the structures of the following alkanals and alkanones:
 - a) butanal
 - b) butan-2-one.
3. Explain why alkanals and alkanones have:
 - a) higher boiling points than corresponding alkanes and alkenes.
 - b) lower boiling points than corresponding alkanols.
4. What are the effects of increase in molecular sizes of alkanals and alkanones on their boiling points?
5. What is the main structural difference between alkanal and alkanones?
6. Explain which one of the two, alkanals and alkanones, can react with acidified potassium dichromate(IV).
7. Write the equation for the reaction between sodium boron hydride, NaBH_4 , and:
 - a) butanal
 - b) propan-2-one.
8. Name each product formed in each reaction in (7) above.

Topic 8: Alkanoic acids

We usually use vinegar to flavour potato chips and for household cleaning. The main component of this important substance is ethanoic acid, which is just one of many compounds called alkanoic acids. Alkanoic acids are also called carboxylic acids. In this topic we will learn more about alkanoic acids.

By the end of this topic you must be able to:

1. identify the functional group of alkanoic acids.
2. name and draw the structures of the first ten unbranched alkanoic acids.
3. write the molecular formula of alkanoic acids given the number of carbon atoms.
4. describe the sources of alkanoic acids.
5. explain the physical properties of alkanoic acids.
6. explain the chemical properties of alkanoic acids.
7. describe the uses of alkanoic acids.

8.1: Functional group of alkanoic acids

Alkanoic acids are organic compounds whose functional group is the carboxyl group, -COOH. The structure of the functional group is as follows:



8.2: Structure and nomenclature of alkanoic acids

Nomenclature of alkanoic acids

Alkanoic acids are named by replacing the ending ‘-e’ in a corresponding alkane with ‘-oic acid’. For example, the name ethanoic acid is made by replacing the ending ‘-e’ in an ethane. To name an alkanoic acid, follow these steps:

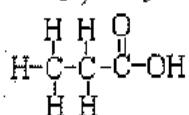
- a) Count all the carbon atoms in the molecule.
- b) Name the corresponding alkane.
- c) Modify the name of the alkane by removing the end -e and replacing it with -oic acid.

Exercise 8.1

1. Name the alkanoic acid with the following number of carbon atoms.
a) 6 b) 8 c) 9

Structures of alkanoic acids

The general structural formula of alkanoic acids is $\text{R}-\overset{\text{O}}{\underset{\text{H}}{\text{C}}}(\text{OH})$, where R is a hydrogen atom or an alkyl group such as CH_3- , CH_3CH_2- or $\text{CH}_3\text{CH}_2\text{CH}_2-$. For example, the structure of propanoic acid is as follows:



Exercise 8.2

Draw the structures of alkanoic acid with the following number of carbon atoms.

- a) 5
- b) 7
- c) 10

8.3: Molecular and general formulae

Molecular formula of alkanoic acids

The condensed formula of propanoic acid, whose structure is given above, is $\text{CH}_3\text{CH}_2\text{COOH}$, and its molecular formula is $\text{C}_2\text{H}_5\text{COOH}$.

Exercise 8.3

1. In the molecular formula $\text{C}_2\text{H}_5\text{COOH}$, express the '5' in terms of the number of carbon atoms.
2. What will be the molecular formula of an alkanoic acid with the following numbers of carbon atoms?
 - a) 5
 - b) 6
 - c) 7
3. What is molecular formula of an alkanoic acid with any number 'n' of carbon atoms?
4. What is the general formula of alkanoic acid? Explain your answer.
5. Use the general formula of alkanoic acids to work out the molecular formulae of alkanoic acids with the following numbers of carbon atoms.
 - a) 1
 - b) 8
 - c) 9
 - d) 10
6. What will be the condensed formula for the alkanoic acid with the number of carbon atoms given (2.) above?

8.4: Sources of alkanoic acids

Natural sources

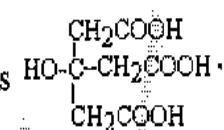
Citrus fruits, sour milk and vinegar taste sour because they contain acids. Some alkanoic acids are found in both plants and animals as follows:

1. Ant stings

Methanoic acid is found in ant stings.

2. Citrus fruits

Citrus fruits such as oranges or lemons contain citric acid, whose structure is



3. Vinegar

Vinegar contains ethanoic acid.

4. Goat milk

Goat milk contains hexanoic acid, $\text{CH}_3(\text{CH}_2)_4\text{COOH}$.

5. Sour milk

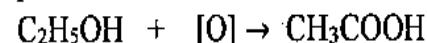
Sour milk contains lactic acid, $\text{CH}_3\text{CHOHCOOH}$.

Preparation of alkanoic acids from oxidation of alkanols and alkanals

Alkanoic acids can be produced from the oxidation reactions of alkanols and alkanals as follows:

1. Oxidation of alkanols

Alkanoic acids can be produced when excess of alkanols are reacted with acidified solutions of potassium dichromate (IV) as the following equation represents:



2. Oxidation of alkanals

Alkanoic acids can also be produced from the oxidation of alkanals as follows:



Experiment 8.1: Investigating formation of ethanoic acid using the reaction of ethanol and oxygen [O].

Materials

- Dilute ethanoic acid (vinegar)

- Ethanol
- Acidified solution of potassium dichromate (IV)
- Test tube
- Hot water in a water bath

Procedure

- 1) Smell the dilute ethanoic acid by wafting your hand slowly across the mouth of the container towards your nose.
- 2) Place 2cm³ of acidified potassium dichromate(IV) into a test tube.
- 3) Note the colour of the acidified potassium dichromate.
- 4) Add four drops of the ethanol into the test tube and shake.
- 5) Place the test tube in the hot water.
- 6) Wait for about 15 minutes.
- 7) Note the colour of the contents in the test tube after 15minutes.
- 8) Smell the gas from the test tube by wafting it with your hand towards your nose.

Conclusion

1. What colour change, if any, did you observe in the test tube after 15 minutes?
2. How did the smell of the dilute ethanoic acid compare with the smell from the test tube?
3. What substance was formed in the test tube?

8.5: Physical and chemical properties of alkanoic acids

A. Physical properties of alkanoic acids

The alkanoic acids are polar. The carboxyl group, -COOH, in alkanoic acids is a combination of the carbonyl group, -C=O and the hydroxyl group, -OH. The carbonyl bond is a hydrogen bond acceptor while the hydroxyl group is a hydrogen bond donor. Therefore, alkanoic acids participate in hydrogen bonding, and usually form dimmers (Figure 8.2). This affects their physical properties as follows:

1. Alkanoic acids dissolve in water.
2. Alkanoic acids have higher boiling points than water and the corresponding alkanols. For example, the boiling point of methanoic acid is 101°C and it is higher than that of either methanol, 64°C which has the same number of carbon atoms, or ethanol 78°C which has the same relative molecular mass of 46. The reasons why alkanoic acids have higher boiling points than the corresponding alkanols are as follows:
 - a) Alkanoic acids form stronger hydrogen bonds than the alkanols. This is because the -OH group of the alkanoic acids is more polar than the -OH group of the alkanols because of the presence of the electron-withdrawing group, -C=O as shown Figure 8.1:



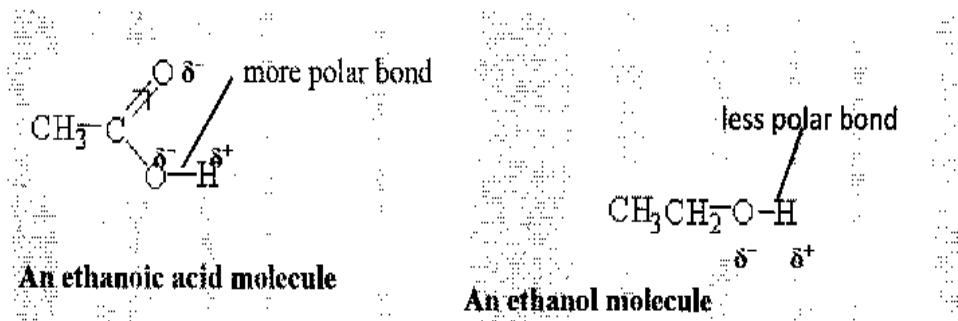


Figure 8.1: Polarity of hydrogen bond in ethanoic acid and in ethanol

- b) Alkanoic acids exist as **dimers**. Dimers are two molecules which are bonded together by two hydrogen bonds as shown in Figure 8.2.

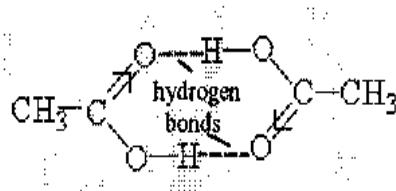


Figure 8.2: Two molecules of ethanoic acid forming a dimer

The formation of the dimers leads to increased melting and boiling points when compared to similar molecules that form only one hydrogen bond. This is because the dimers lead to the formation of stronger intermolecular forces between the molecules and also between the dimers themselves.

Exercise 8.4

Explain by referring to the type of intermolecular forces why alkanoic acids:

- are soluble in water
- have higher boiling point than water.

Trends in physical properties of alkanoic acids

Experiment 8.2: Investigating the trend in the boiling points of alkanoic acids

Materials

- source of heat
- ethanoic acid
- propanoic acid
- hexanoic acid
- thermometer
- 100ml beaker

Safety precaution

Care must be taken when handling hot substances.

Procedure

1. Arrange apparatus as shown in Figure 8.3.
2. Heat the ethanoic acid until it boils.
3. Record the temperature at which it boils..
4. Repeat steps 1 to 3 with propanoic acid and hexanoic acid

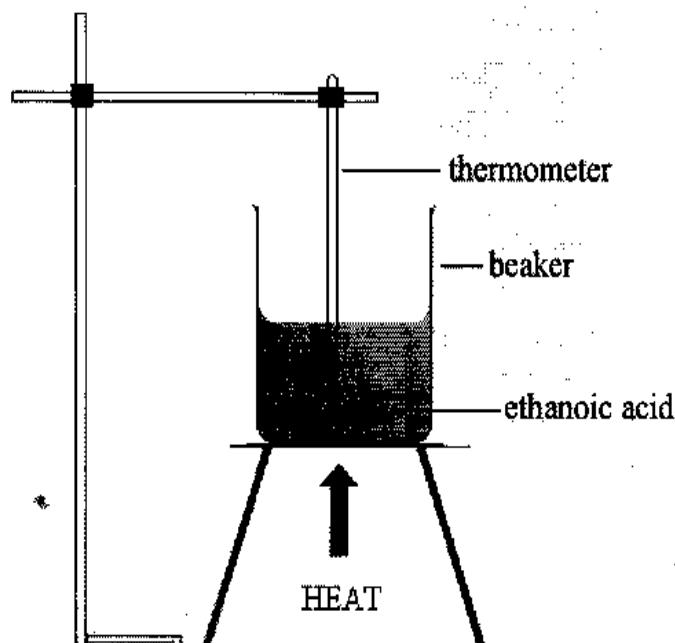


Figure 8.3: Measuring boiling of ethanoic acid

Discussion

1. Arrange the alkanoic acids in the order of increasing boiling points.
2. What is the trend in the boiling points of the alkanoic acids?
3. Explain the trend in the boiling points of alkanoic acids by referring to the type of intermolecular forces.

Experiment 8.3: Investigating the trend in the densities of alkanoic acids

Materials

1. Ethanoic acid
2. Butanoic acid
3. Hexanoic acid
4. Decanoic acid
5. Measuring cylinder
6. Beam balance

Procedure

- 1) Copy Table 8.1 in your note books.
- 2) Measure mass of empty measuring cylinder and record it in the table.
- 3) Pour 10cm³ of ethanoic acid into the measuring cylinder.
- 4) Measure the mass of the ethanoic acid plus the measuring cylinder.
- 5) Find the mass of the ethanoic acid (mass of measuring cylinder plus ethanoic acid minus mass of empty measuring cylinder).
- 6) Record the mass of the ethanoic acid in the table of results.
- 7) Calculate the density of the ethanoic acid and record it in the table of results.
- 8) Repeat steps 2 to 7 with Butanoic acid, hexanoic acid and decanoic acid.

Table 8.1: Trends in densities

Mass of empty measuring cylinder	Alkanoic acid	Mass of measuring cylinder + alkanoic acid	Mass of alkanoic acid (g)	Volume of alkanoic acid (cm ³)	Density (mass /volume) (g/cm ³)
	Ethanoic acid			10	
	Butanoic acid			10	
	Hexanoic acid			10	
	Decanoic acid			10	

Discussion

1. Arrange the alkanoic acids in the order of increasing densities?
2. What is the trend in the densities of the alkanoic acids?
3. Write a laboratory report.

Experiment 8.4: Investigating the trend in the solubilities of alkanoic acids

Materials

- Different alkanoic acids such as methanoic acid, ethanoic acid, pentanoic acid, octanoic acid and decanoic acid
- Water
- Test tubes in a rack
- Droppers

Procedure

- 1) Label the test tubes with the names of the alkanoic acids.
- 2) Pour 2cm³ of methanoic acid into a test tube labelled methanoic acid (2cm³ is as about 2cm high in the test tube).
- 3) Add 4 drops of water into the test tube and shake.
- 4) Put it into the rack.

- 5) Repeat steps 2 to 4 with the other alkanoic acids.
- 6) Observe the contents in each test tube and record your observations.

Discussion

- a) How many layers of liquids did you observe in each test tube?
- b) What does the number of layers tell you about the solubility of the alkanoic acid in water?
- c) Which alkanoic acids are soluble in water?
- d) What is the trend in the solubilities of the alkanoic acids in water?

Viscosities of alkanoic acids

Viscosity is defined as a liquid's resistance to flow. Liquids with stronger intermolecular forces have higher viscosities than those with weaker intermolecular forces.

Experiment 8.5: Investigating the trend in the viscosities of alkanoic acids

Materials

- 25 ml pipette
- Alkanoic acids (methanoic acid, ethanoic acid, pentanoic acid, octanoic acid, decanoic acid)
- Stop watch
- 25ml measuring cylinder
- A marker

Procedure

1. Use the marker to mark the 20ml line on the cylinder.
2. Fill the pipette with methanoic acid up to the 25 ml mark.
3. Hold the pipette above the cylinder.
4. Let the methanoic acid flow from the pipette.
5. Begin the stopwatch at the moment the first drop of the acid hits the bottom of the beaker.
6. Stop the clock when the methanoic acid reaches the 20ml mark on the measuring cylinder.
For the sake of validity continue pouring from the pipette until all the acid is in the cylinder.
7. Record the flow time.
8. Clean the pipette and the measuring cylinder thoroughly. Reset the stopwatch and repeat steps 2 to 7 for a total of three tests with methanoic acid.
9. Repeat steps 2 to 8 for each alkanoic acid.
10. Calculate the average test time for each alkanoic acid.
11. Analyse the viscosity data.

Discussion

- 1) Arrange the alkanoic acids in the order of increasing average flow time.

- 2) The liquid with the highest average time is the most viscous and that with the lowest average time is the least viscous. Which alkanoic acid is:
 - a. most viscous?
 - b. least viscous?
- 3) What is the trend in the viscosities of alkanoic acids?
- 4) By referring to the types of intermolecular forces in alkanoic acids, explain the trend in their viscosities.

Electrical conductivity of alkanoic acids

Electrical conductivity is the measure of electric current a material can allow to pass through it. Some materials allow electric currents to flow through them, while others do not.

Experiment 8.6: Investigating electrical conductivity of aqueous solution of ethanoic acid

Materials

1. Dilute ethanoic acid (vinegar can do)
2. Two carbon electrodes
3. Four connecting wires
4. An ammeter
5. Two cells
6. A cell holder
7. A switch

Procedure

- 1) Connect a circuit as shown in Figure 8.4.
- 2) Record the ammeter reading.
- 3) Close the switch.
- 4) Record the ammeter reading again.

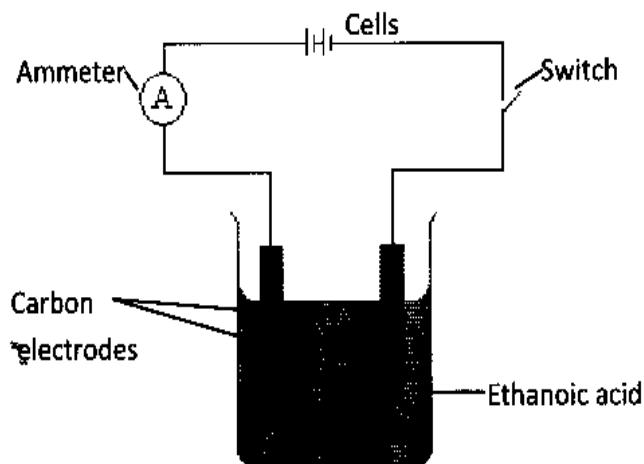


Figure 8.4: Electrical conductivity apparatus



Discussion

1. What was the reading of the ammeter before closing the switch?
2. What was the reading of the ammeter after closing the switch?
3. Does ethanoic acid conduct electricity? Explain your answer.

Effect of alkanoic acids on acid-base indicators

As we learnt in Form 2, that an acid-base indicator is a substance that changes colour on going from acid to base or vice versa. Examples of acid-base indicators are red and blue litmus papers, universal indicator and phenolphthalein indicator. Acids and bases affect the indicators differently.

Experiment 8.7: Investigating the effect of alkanoic acids on blue and red litmus papers

Materials

- Ethanoic acid
- Blue litmus paper
- Red litmus paper

Procedure

- 1) Dip a piece of red litmus paper into the ethanoic acid.
- 2) Record the colour of the litmus paper.
- 3) Repeat steps 1 and 2 with blue litmus paper.

Conclusion

1. What was the colour of the litmus paper :
 - a) before dipping it into the ethanoic acid?
 - b) after dipping it into the ethanoic acid?
2. What is the effect of alkanoic acid on litmus papers?

Experiment 8.8: Investigating the pH of ethanoic acid

Materials

- Universal indicator
- Universal indicator colour chart
- Ethanoic acid test tube

Procedure

- 1) Put 2cm³ of ethanoic acid into the test tube.
- 2) Add three drops of the universal indicator to the ethanoic acid in the test tube.
- 3) Use the universal colour chart to find the pH value of the ethanoic acid.



Discussion

1. What was the pH value of the ethanoic acid?
2. Is ethanoic acid neutral, acidic or basic? Explain your answer.

Experiment 8.9: Investigating the effect of ethanoic acid on phenolphthalein indicator

Materials

- Ethanoic acid
- A test tube
- Phenolphthalein indicator

Procedure

- 1) Place 10 drops of ethanoic acid into the test tube.
- 2) Add 3 drops of phenolphthalein indicator into the test tube.
- 3) Record your observations

Discussion

1. What changes (if any) did you observe after adding ethanoic acid into the test tube?
2. What is the effect of ethanoic acid on phenolphthalein indicator?

B. Chemical properties of alkanoic acids

Alkanoic acids undergo a number of chemical reactions. They can react with alkali metals, hydroxides, carbonates to produce metal alkanoates (salts) such as sodium methanoate (HCOONa) and other products. They also react with alkanols to produce alkyl alkanoates. For example, propanoic acid can react with methanol to produce methyl propanoate ($\text{CH}_3\text{CH}_2\text{COOCH}_3$) and water.

Experiment 8.10: Investigating the reaction of ethanoic acid with alkali metals

Materials

- Sodium or potassium metal
- Ethanoic acid
- Test tube
- Matches
- A wood splint

Procedure

- 1) Place 2cm^3 of ethanoic acid in a test tube.
- 2) Add a small piece of sodium metal.



- 3) Light the splint with the matches.
- 4) Bring the glowing splint at the mouth of the test tube.
- 5) Record your observation.

Discussion

1. What happened to the lithium metal in the test tube?
2. What did you observe about the glowing splint when it was placed at the mouth of the test tube?
3. What gas was produced during the reaction?
4. Write a balanced chemical reaction for the reaction between lithium metal and ethanoic acid.

Experiment 8.11: Investigating the reaction of ethanoic acid with hydroxides

Materials

- Dilute sodium hydroxide solution
- Dilute ethanoic acid
- A test tube
- Phenolphthalein indicator

Procedure

1. Place 2 cm³ of sodium hydroxide solution into a test tube.
2. Add 3 drops of phenolphthalein indicator into the test tube.
3. Record your observations.
4. Add 10 drops of ethanoic acid into the test tube.
5. Record your observations.

Discussion

1. What was the colour of the phenolphthalein indicator in sodium hydroxide solution?
2. What changes (if any) did you observe when ethanoic acid was added to the hydroxide solution? Explain your observations.
3. What are the products of the reaction between a base and an acid?
4. Write down a balanced chemical equation for the reaction between ethanoic acid and the sodium hydroxide.

Experiment 8.12: Investigating the reaction of ethanoic acid with carbonates

Materials

- | | |
|-----------------------------|-----------------------------|
| • Calcium carbonate | • Rubber tubing |
| • Ethanoic acid | • Glass tubing |
| • A test tube with side arm | • Lime water |
| • A test tube | • Two clamp stands |
| • Stopper | • Phenolphthalein indicator |



Procedure

- 1) Arrange the apparatus as shown in figure 8.5.
- 2) Place 2cm³ of ethanoic acid into the test tube with side arm.
- 3) Add 3 drops of phenolphthalein indicator.
- 4) Record your observations.
- 5) Carefully add a $\frac{1}{4}$ spatulaful of calcium carbonate.
- 6) Close the test tube with the stopper.
- 7) Check the colour of the lime water.
- 8) Record the results.

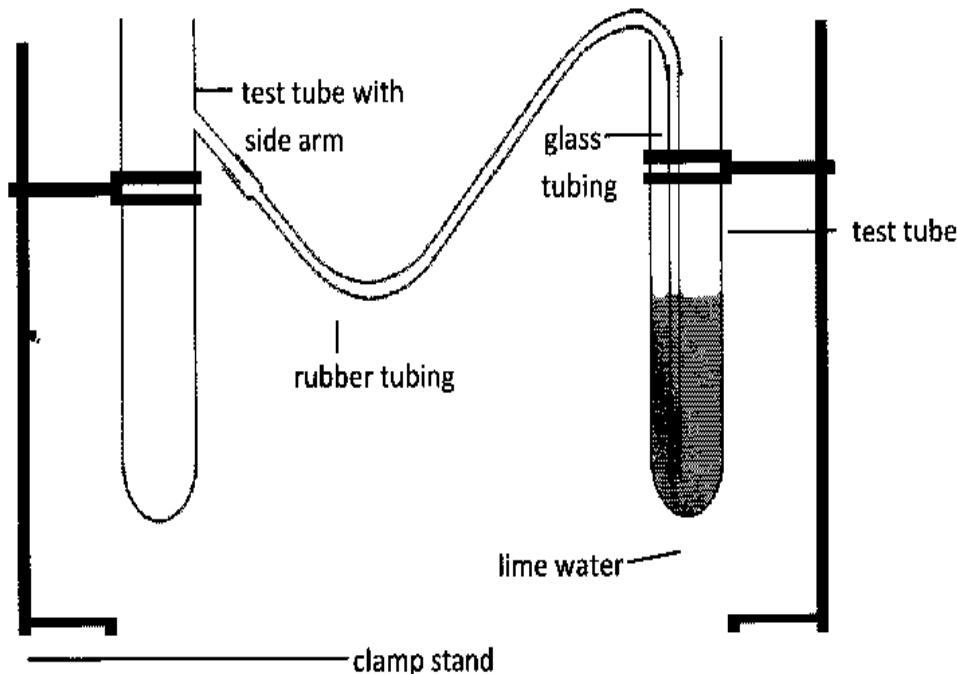


Figure 8.5: Apparatus for investigating the reaction of ethanoic acid with calcium carbonates

Discussion

1. Explain the changes that took place in the test tube (if any) after adding calcium carbonate.
2. What was the colour of the lime water:
 - a) before the reaction?
 - b) after the reaction?
3. What gas was produced during the reaction? Explain your answer.
4. Write a balanced chemical equation for the reaction between ethanoic acid and calcium carbonate.
5. What was the use of the lime water this experiment?
6. What is your conclusion on the reaction between alkanoic acids and carbonates?

Experiment 8.13: Investigating the reaction of ethanoic acid with alkanols

Materials

- Ethanol
- Ethanoic acid
- Cold water
- Test tube
- Concentrated sulphuric acid
- A 100ml beaker
- Source of heat
- Warm water in a water bath

Procedure

- 1) Place 2cm³ of ethanoic acid in the test tube.
- 2) Add an equal amount of ethanol followed by two drops of concentrated sulphuric acid.
- 3) Warm the mixture gently (in a warm water bath).
- 4) Pour the mixture into a beaker with about 50cm³ of cold water.
- 5) Smell the product by wafting your hand across the mouth slowly towards your nose.

Discussion

1. What did the product smell like?
2. What was the use of the concentrated sulphuric acid?
3. Suggest a chemical equation for the reaction between ethanoic acid and ethanol.

8.6: Uses of alkanoic acids

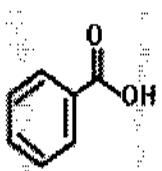
Alkanoic acids have a number of uses in our everyday life. Some of the uses include as solvents and food preservatives.

a) As solvents

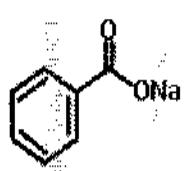
Alkanoic acids such as ethanoic acid dissolve both polar compounds such as sugar and non-polar compounds such as oils and elements such as sulphur and iodine.

b) As food preservatives

Alkanoic acids are used as preservatives in foods and drugs. For example, benzoic acid (Figure 8.6(a)), which is an alkanoic acid with the carboxyl group bonded to benzene ring, can be converted to its salts such as sodium benzoate (Figure 8.6(b)) for use as preservative in foods and drugs.



(a) Benzoic acid



(b) Sodium benzoate

Figure 8.4: Benzoic acid and sodium benzoate

Exercise 8.5

1. Write a balanced equation for a reaction of benzoic acid to produce sodium benzoate.
2. Collect labels of containers of different food and beverages. Check the ingredients on each label. Which foods and beverages use sodium benzoate as preservative?

Topic review

1. What is the functional group of alkanoic acids?
2. What is the general formula of alkanoic acids?
3. Name the alkanoic acid that is found in each of the following:
 - a) goat milk.
 - b) sour milk.
 - c) ant stings.
4. Which type of alkanols can be oxidised to give alkanoic acids?
5. Name the alkanoic acid that will be produced from the reaction of potassium dichromate (IV) with:
 - a) propanal.
 - b) butan-1-ol.
6. Explain why alkanoic acids are soluble in water.

Topic 9: Alkanoates (Esters)

We are familiar with the pleasant smells of ripe fruits such as bananas and pineapples. We are also, probably aware of the pleasant flavours of different non-alcoholic beverages such as fanta orange, and pineapple. The substances that are responsible for these smells, flavours and the fragrances of various types of perfumes are organic compound that belong to a family known as alkanoates (also known as esters). Alkanoates are produced naturally in ripening fruits, but can also be produced artificially in the laboratory.

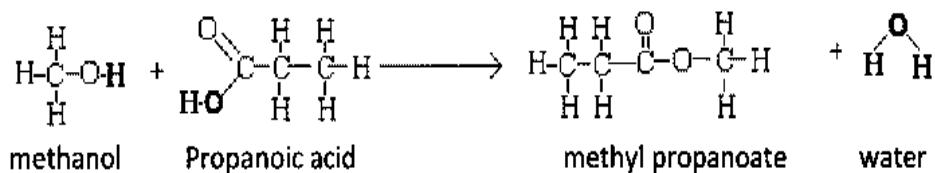
In this topic, we will learn more about these sweet-smelling and sweet-tasting organic compounds called alkanoates or esters.

By the end of this topic you must be able to

1. identify the functional group of alkanoates (esters).
2. name and draw the structures of alkanoates.
3. describe the sources of alkanoates.
4. describe properties of alkanoates.
5. state uses of alkanoates.
6. describe the process of soap making (saponification).

9.1: Functional group

We have seen, in the Topic 8, that alkanoic acids react with alkanols to form alkanoates (or esters) in a reaction called **esterification**.



The functional group of the alkanoates is ' $\text{C}=\text{O}-$ '. In the functional group, the bridging ' -O- ' comes from the alkanol parent, whereas the ' C=O ' comes from the alkanoic acid parent.

9.2: Structure and nomenclature

Naming alkanoates

Alkanoates are named by combining the names of the parent alkanol and the parent alkanoic acid from which they are made. The parent alkanol gives the first part of the name, while the parent alkanoic acid gives the second part of the name.

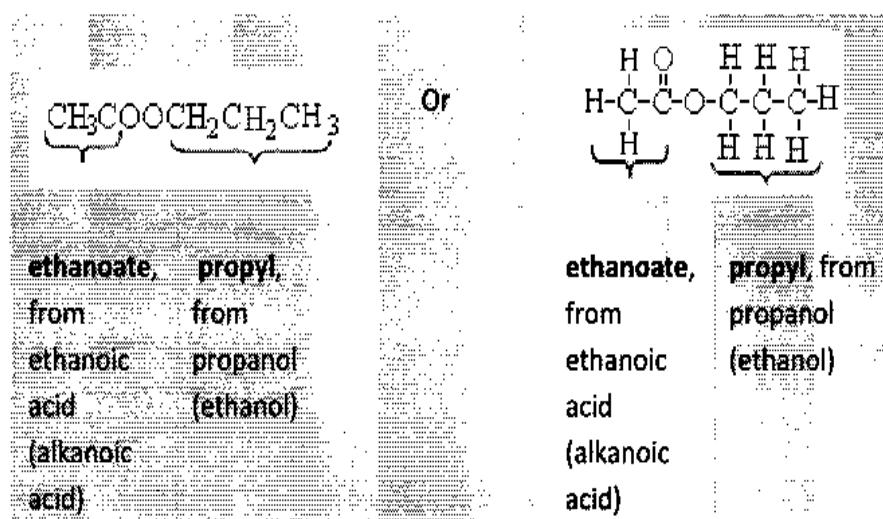
Naming alkanoates given parent alkanols and alkanoic acids

To name an alkanoate, given the parents alkanol and alkanoic acid:

- Replace the ending ‘-anol’, in the name of the alkanol, with ‘-yl’.
- Replace the ‘-oic acid’ in the parent alkanoic acid with ‘-oate’.
- Combine the two words to form the name of the alkanoate.

Naming alkanoates given condensed or full structural formulae

In the condensed and the structural formulae of alkanoates, the first part of the formula is derived from the parent alkanoic acid while the second part is derived from the parent alkanol. For example, consider the following formulae of propyl ethanoate:



To name an alkanoic acid, given its condensed or the full structural formulae, identify the parent alkanoic acid and the parent alkanol and then follow the steps above to come up with the name.

Exercise 9.1

- Name the alkanoate that is formed from the reaction of
 - propanol and butanoic acid
 - methanoic acid and pentanol
- Name the following alkanoates:



Drawing the structures of the alkanoates formed from given alkanoic acids and alkanols

When drawing the structures of the alkanoates formed from given alkanoic acid and alkanol always remember the following:

1. The first part of the structure is derived from the parent alkanoic acid while the second part is derived from the parent alkanol.
2. The bridging ‘-O-’, in the functional group, comes from the alkanol parent, whereas the ‘-C=O’ comes from the alkanoic acid parent.
3. An H is removed from the alkanol and an -OH is removed from the alkanoic acid, and the remaining parts join together to form the alkanoate.

Exercise 9.2

Draw the structures of the alkanoates that are formed from the following alkanoic acids and alkanols:

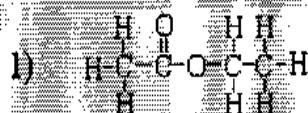
1. pentanoic acid and propanol.
2. methanoic acid and methanol.
3. octanoic acid and nonanol.
4. heptanoic acid and methanol.

Deducing the reactants of esterification given the structure of the product

Remember that the first part of the structure of an alkanoate is derived from the parent alkanoic acid while the second part is derived from the parent alkanol. Therefore, the reactants of an esterification can be deduced from the structure of the product of the reaction by identifying the name of the alkanoic acid from which the first part of the structure is derived and also the name of the alkanol from which the last part of the name is derived.

Exercise 9.3

Name the alkanoic acids and the alkanols from which the following alkanoates are formed:



9.3: Sources of alkanoates

Alkanoates are produced naturally by both plants and animals. The production of alkanoates by natural means is a slow process. However, alkanoates can also be faster, by reacting alkanoic acids with alkanols artificially using catalysts.

Natural sources of alkanoates

1) Fruits and flowers of plants

Alkanoates are produced naturally in fruits and flowers of plants such as nuts, avocados and olives. The flavours of fruits and sweet smells of flowers are due to the alkanoates. Vegetable oils are naturally occurring alkanoates which are used as energy storage compounds for the plants (Figure 9.1).

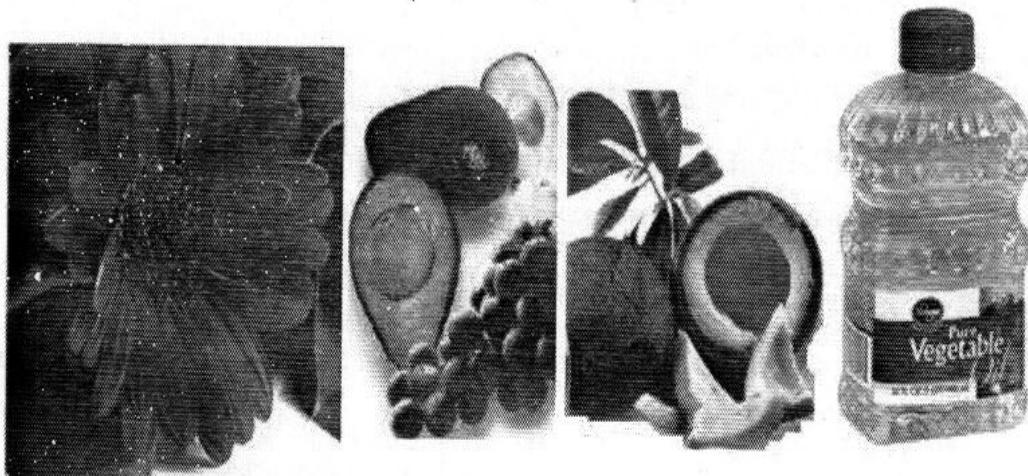


Figure 9.1: Examples of alkanoates and their sources

2) Animals

Alkanoates are also produced naturally in bodies of animals (Figure 9.2). They are used as energy storage compounds for the animals. The alkanoates produced by animals are mainly in form of fats

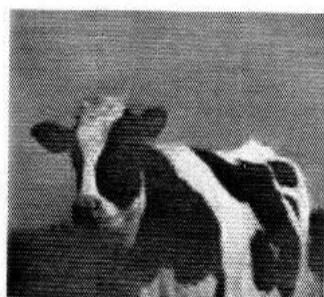


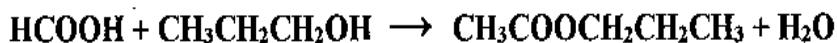
Figure 9.2: Animals such as these are also sources of alkanoates

Artificial sources of alkanoates

Alkanoates can be manufactured artificially by reacting alkanoic acids and alkanols as follows:



An example is the formation of propyl ethanoate from the reaction of methanoic acid and propanol:



The reaction between an alkanol and an alkanoic acid is called **esterification**. It is also called a **condensation reaction** because a water molecule is released during the reaction from the two reactant functional groups.

9.4: Properties of alkanoates

Physical properties of alkanoates

Alkanoates have the following physical properties:

- 1) They have pleasant sweet smell.
- 2) They are not readily soluble in water.
- 3) They have lower melting points and boiling points than alkanoic acids and alkanols of similar molecular masses.

Chemical properties of alkanoates

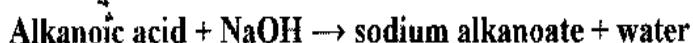
Alkanoates undergo hydrolysis reactions with sodium hydroxide. Hydrolysis of alkanoates with sodium hydroxide takes place in two stages:

Stage 1: Reaction of alkanoate with water in sodium hydroxide

When an alkanoate is boiled in an aqueous solution of sodium hydroxide, it reacts with water to form the parent alkanoic acid and alkanol:



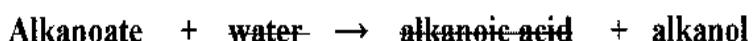
Stage 2: Sodium hydroxide neutralises the alkanoic acid formed in stage 1 to produce a sodium alkanoate salt:



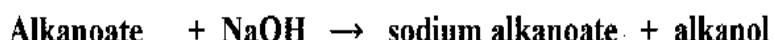
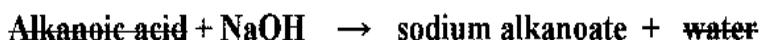
Overall reaction

The overall reaction is that an alkanoate reacts with sodium hydroxide to produce a sodium alkanoate salt and an alkanol as follows:





+



The alkanol can be obtained from the mixture by distillation or evaporation.

Exercise 9.4

1. Suggest reasons for the following physical properties of alkanoates by referring to intermolecular forces in them:
 - a) they are not readily soluble in water.
 - b) they have lower melting points and boiling points than alkanoic acids and alkanols of similar molecular masses.
2. Using balanced chemical equations, describe the stages in the hydrolysis of the following alkanoates with sodium hydroxide. In each case name the overall products.
 - a) Methyl ethanoate.
 - b) Ethyl methanoate.
 - c) Propyl ethanoate.

9.5: Uses of alkanoates

We have seen that alkanoates are volatile and they produce strong and pleasant smell. Because of this, alkanoates are used in a number of ways in our everyday life.

1. As solvents

Alkanoates are used as solvent. For example, ethyl ethanoate is used in nail polish remover.

2. As flavourings

Alkanoates are used to flavour some food and beverages. For example, some biscuits and soft drinks are flavoured with alkanoates. Pineapple flavours are produced from ethanol and butanoic acid, while banana flavour is produced from pentanol and ethanoic acid.



Figure 9.3: Examples of food stuffs that are flavoured with alkanoates

3. As fragrances

Because of their strong and pleasant smells, alkanoates are used as fragrances in perfumes and other beauty products.

9.6: Saponification

Fats and oils

We have seen that fats and oils are natural alkanoates produced in animals and plants. These alkanoates are called **triglycerides**. They are produced from a reaction of an alkanol (propan-1,2,3-triol) and alkanoic acids with long carbon chains (see Figure 9.4). The alkanol is known as glycerol (or glycerine), while the alkanoic acids with the long carbon chains are called fatty acids. Fats are the triglycerides which are solid at room temperature while oils are the triglycerides which are liquids at room temperature.

Exercise 9.5

By referring to the structure of the molecule of the tryglyceride in Figure 9.4:

- 1) suggest why fats and oils qualify to be called alkanoates.
- 2) how many alkanoic acids are in each triglyceride molecule?
- 3) draw the structures of the propan-1,2,3-triol and alkanoic acids which joined to make the triglyceride molecule.
- 4) why are fats and oils insoluble in water?

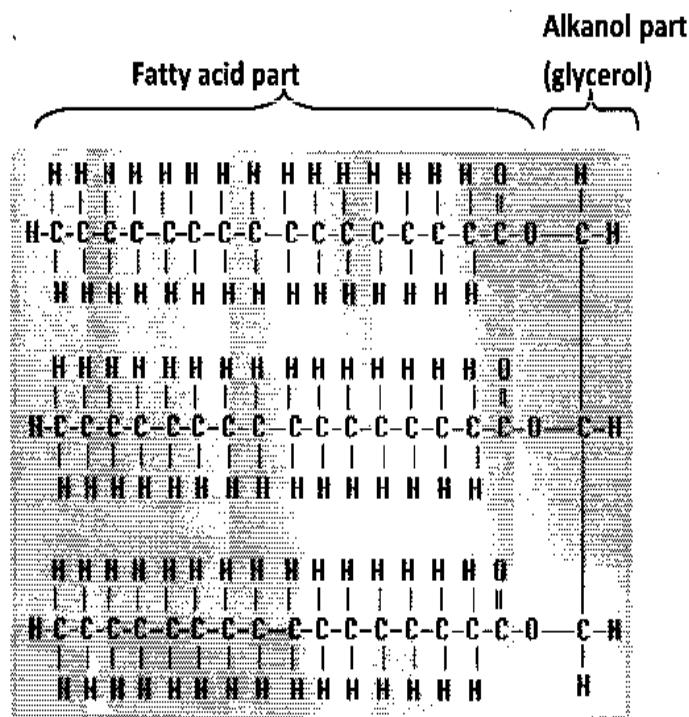


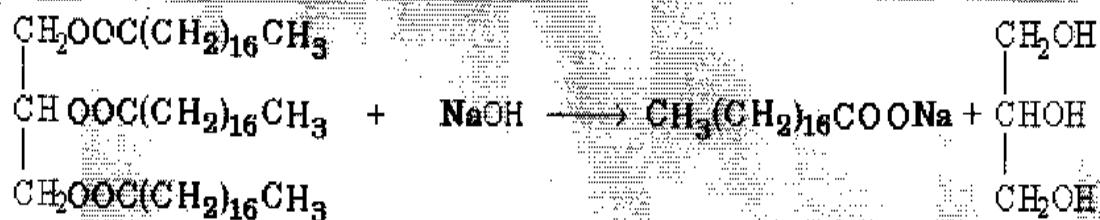
Figure 9.4: A triglyceride (an alkanoate)

Saponification

We have seen that alkanoates undergo hydrolysis with aqueous solutions of sodium hydroxide to produce sodium alkanoate salts. If the alkanoate used is an animal fat or vegetable oil the salts produced are soaps. Soap is a sodium or potassium salt of fatty acid. The reaction of fats or oils with aqueous sodium hydroxide to form soap is known as saponification.

Exercise 9.6

The following equation represents the hydrolysis of an animal fat (prop-1,2,3-triyl triooctadecanoate) with sodium hydroxide (NaOH).



1. Balance the equation.
2. Name the soap molecule that is formed during this reaction.
3. How many soap molecules can be formed from each molecule of triglyceride (the animal fat)?
4. Describe how soap is produced.

Experiment 9.1: Investigating the reaction of fats and oils with sodium or potassium hydroxide (Adapted from Childs, A (2000) P. 447)

Materials

- A beaker (250cm³)
- Measuring cylinder (10cm³)
- Evaporating basin
- Source of heat
- Tripod stand and wire gauze
- Water
- Cooking oil
- Sodium hydroxide solution (6M)
- Saturated salt solution (sodium chloride)
- Potassium hydroxide (6M)
- Animal fat

Safety precautions!

Concentrated sodium and potassium hydroxide is corrosive. If spilt onto skin, wash with plenty of water. It is necessary to wear eye protection. If sodium/potassium hydroxide is spilt into the eye, wash immediately with plenty of water over a few hours.

Procedure

1. Set the apparatus as shown in Figure 9.5.
2. Measure 25cm³ of sodium hydroxide and 10cm³ of oil into the evaporating basin.
3. Heat for 30 minutes.
4. Allow the solution to cool.
5. Add 10cm³ of saturated sodium chloride solution.
6. Separate soap from the mixture by filtration.
7. Boil the crude soap curds in water and reprecipitate the soap with sodium chloride and filtering again.
8. Repeat step 7 until pure soap is obtained.
9. Repeat steps 1 to 7 with animal fat.
10. Repeat steps 1 to 9 with potassium hydroxide and saturated potassium chloride solution.

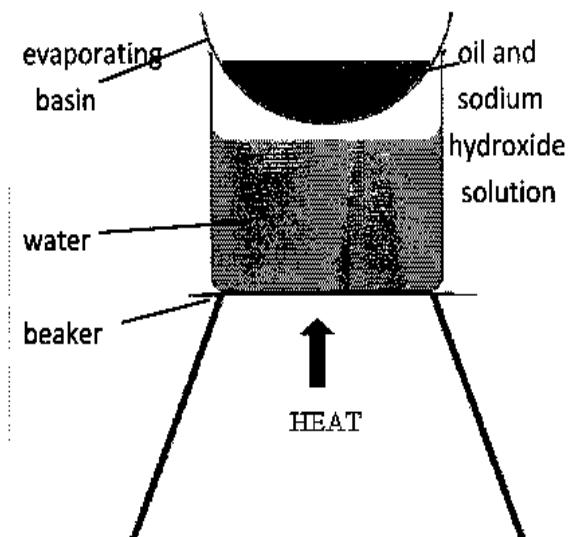


Figure 9.5 : Apparatus for preparation of soap

Discussion

- 1) What is the difference between soap made by reacting triglyceride with sodium hydroxide and the soap made by reacting potassium hydroxide with triglyceride in terms of hardness?
- 2) What, do you think, is the reason for adding saturated salt solution to the mixture after saponification reaction?
- 3) What is the colour if the soap that you made?
- 4) How do the colours and the fragrances in some commercial soaps come about?
- 5) Write a laboratory report.

Topic review

1. a) What is esterification?
b) How is esterification important us?
2. Name the alkanols and alkanolic acids that combined to form the following alkanoates:
 - a) $\text{CH}_3\text{COOCH}_3$
 - b) HCOOCH_3
 - c) $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOCH}_2\text{CH}_2\text{CH}_3$
3. Name alkanoate in (2.) above.
4. What is meant by the following terms:
 - a) triglyceride?
 - b) saponification?
5. Discuss the economic importance of saponification in our life.
6. Why is it important for us to study alkanoates?

Topic 10: Identification of unknown organic compounds

Different organic compounds have different structures and functional groups. The different functional groups make the compounds to have different properties. In this topic we will learn different ways of identifying organic compounds.

By the end of this topic you must be able to:

1. deduce the family and structural formula of an unknown organic compound.
2. distinguish organic compounds basing on their properties.

10.1: Deducing family and structural formula of an unknown organic compound

There are a lot of families of organic compounds in nature, but starting from our JCE chemistry up to now we have looked at seven of them. These are: **alkanes, alkenes, alkanols, alkanals, alkanones, alkanoic acids and alkanoates**. It is now possible to deduce the family and the structural formula of an unknown organic compound given relevant information such as:

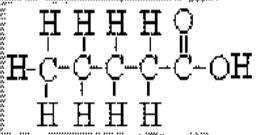
1. structural formula
2. general formula
3. product of their reactions
4. physical properties
5. chemical properties

Using structural formula of organic compounds to deduce family of an unknown organic compound

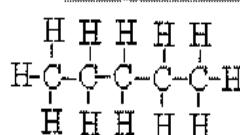
Given the structural formulae of organic compounds, you can tell its family by identifying the functional group present in the structural formulae. It is, therefore, necessary to be familiar with the functional group of each organic compound.

Exercise 10.1

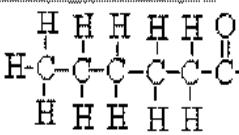
1. Write the functional group of each of the following organic compound.
 - a. alkanes.
 - b. alkenes.
 - c. alkanones.
 - d. alkanoic acids
 - e. alkanoates.
 - f. alkanol.
 - g. alkanals.
2. Name the family of the organic compounds to which each of the following structural formulae of organic compounds belong. Explain your answer in each case.



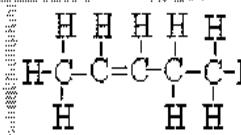
(a)



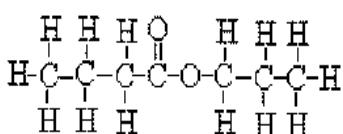
(b)



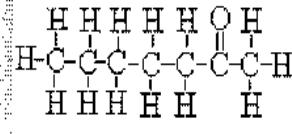
(c)



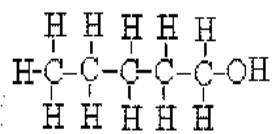
(d)



(e)



(f)



(g)

Using general formula of organic compounds to deduce family of an unknown organic compound

Each family of organic compounds has its general formula which can be used to work out the molecular formulae of the members in that family. General formula can therefore, be used to deduce the family and the molecular formula of organic compounds.

Exercise 10.2

1. Name two families of organic compounds that have the same general formula.
2. Given below are the general formulae of some families of organic compounds represented by letters P, Q, R, S and T.

$$P: C_n H_{2n}$$

$$Q: C_n H_{2n+1} COOH$$

$$R: C_n H_{2n+1} OH$$

$$S: C_n H_{2n+2}$$

$$T: C_n H_{2n} O$$

- a. Name the family of organic compound (homologous series) represented by the letters P and R.
- b. Which general formulae represent hydrocarbons?
- c. Which class of the organic compounds does the general formula $C_n H_{2n+1} COOH$, represent?

- d. In family Q, draw the structure of a compound
- when $n = 4$ in homologous series Q.
 - with 4 carbon atoms
 - with 5 carbon atoms
- e. Name the compound drawn in 2d i) above.

Using properties of organic compounds and tests to deduce families of unknown compounds

Families of organic compounds have unique properties which are determined by the nature of their functional groups. Organic compounds can be distinguished from each other by carrying out tests on their functional groups using such properties. These properties include the following:

1. Solubility in water.
2. Effects on:
 - a) acid-base indicators.
 - b) solution of Brady's reagent.
 - c) solution of Tollen's reagent.
 - d) solution of Fehling's reagent.
 - e) solution of acidified potassium dichromate (IV).
 - f) colour of bromine solution.

Solubility in water

Alkanes and alkenes are insoluble in water. Smaller alkanols, alkanoic acids, alkanals, and alkanones are soluble in water. But as the number of carbon atoms in these compounds increases they become insoluble in water.

Effect on acid-base indicators

Alkanoic acids turn blue litmus paper red and turn the pink colour of phenolphthalein indicator in alkaline solutions colourless. The other carbon compounds do not affect the acid-base indicators.

Effect on Brady's reagent

Alkanals and alkanones form yellow-orange precipitate in Brady's reagent. The other organic compounds do not react with the Brady's reagent.

Effect on Tollen's reagent or Fehling's test

- Alkanals produce a dark red precipitate when added to the blue solution of Fehling's reagent. Alkanones and the other organic compounds produce no colour change in the solution of Tollen's reagent.
- Alkanals produce a grey precipitate of silver or a silver mirror on the sides of the test tube when added to the colourless solution of Fehling's reagent. Alkanones and the other organic compounds produce no colour change in the solution of Fehling's reagent.

Effect on acidified potassium dichromate (IV)

Alkanols and alkanals turn the colour of the solution of acidified potassium dichromate (IV) from orange to green. The other organic compounds do not affect the colour of the solution of the acidified potassium dichromate (IV).

Effect on the colour of bromine solution

Bromine solution is red-brown in colour. When an alkene is added to the bromine solution, it turns colourless. All other organic compounds do not change colour of bromine solution.

10.2: Construction of flow diagrams

A flow diagram is a picture-way of step-by-step identification and separation of unknown organic compounds using special tests for their functional groups.

Exercise 10.3

1. The flow diagram in **Figure 10.1** can be used to distinguish the given organic compounds. The letters (A) to (J) represent expected results after each test.
 - a) Write the expected results after each test as represented by the letters.
 - b) Modify the flow diagram so that it should include a solubility test.
2. You are given the following organic compound: pentan-1-ol, hexene, butane, heptanoic acid, hexan-2-one and octanal in bottles labelled A, B, C, D, E and F, not in any order.
 - a) Construct a flow diagram that you can use to distinguish these organic compounds from each other. (Remember that solubility of some organic compounds decreases with increase in molecular sizes).
 - b) Use your flow diagram and tests for each functional group to identify the organic compounds in each bottom.

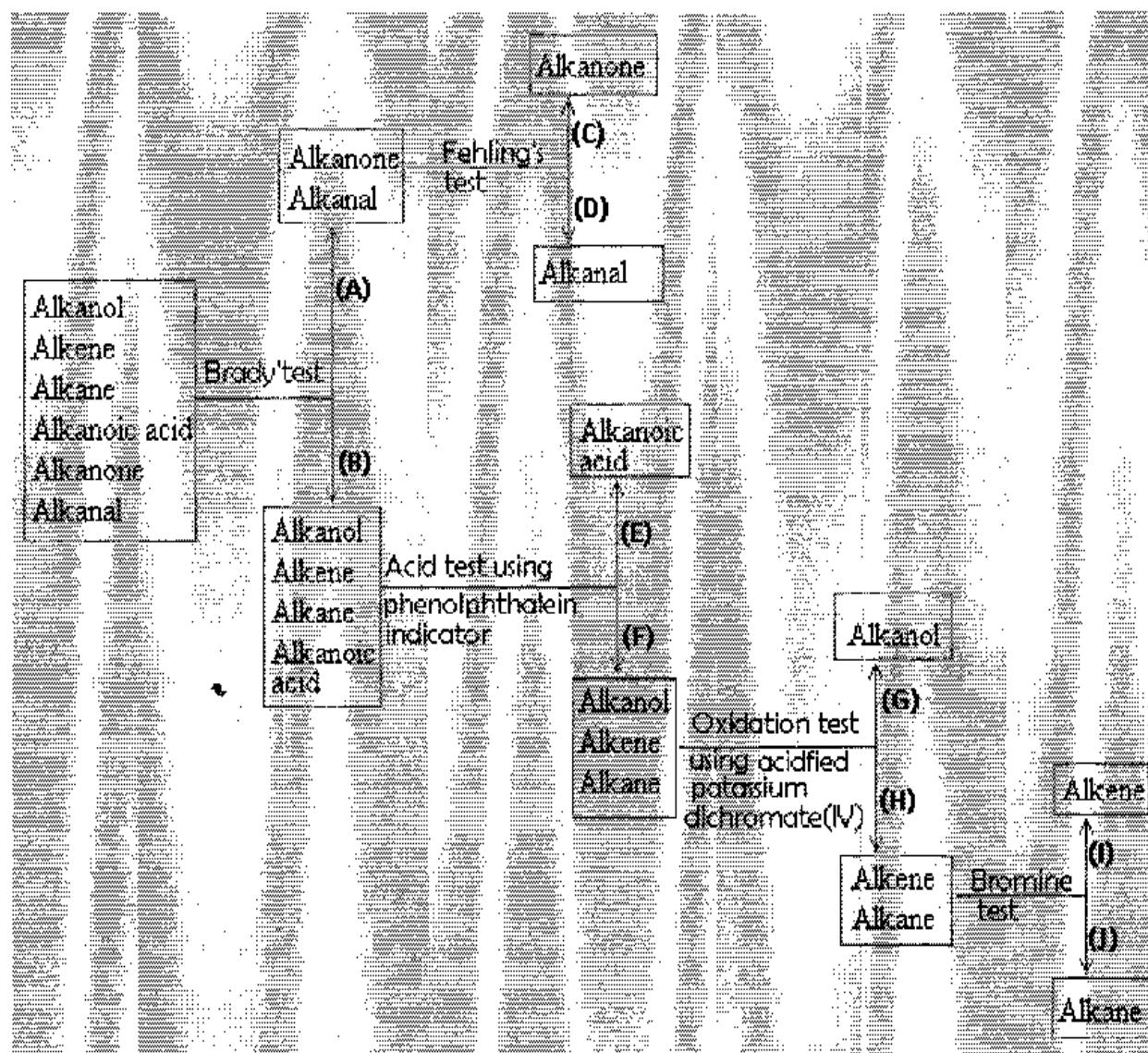


Figure 10.1: A flow diagram for identifying organic compound

References

- Alcohol's effects on the body. *National Institute on alcohol abuse and alcoholism(NIAAA)*. Retrieved January 16, 2016 from <http://www.niaaa.gov/alcohol-health/alcohols-effects>
- Chang, R. (2010). *Chemistry, 10th edition*. Newyork .The McGraw-Hill Companies, Inc.
- Childs, A. (2000). *Macmillan Secondary Chemistry*. Oxford. Macmillan education.
- Earl, B. and Wilford, L. (2001). *GCSE Chemistry. (Second edition)*. London.John Murray (Publishers) Ltd.
- Gallagher, R.and Ingram P. (2000). *Complete Chemistry*. New York. Oxford.
- GSCE CHEMISTRY. What is Test for Hydrogen Gas? –How di I test for hydrogengas?
Retrieved December 15, 2015 from <http://www.gscescience.com/itesthydrogen.html>
- McDuel, B. (Editor). (2000). *Teaching secondary chemistry*. London. John Murray (Publishers)Ltd.
- Michael, R and Colen, R. (n.d.). 10 overcounter medicines abused by teens. Retrieved Decamber 15, 2015 from
www.philly.com/blogs/healthycare/10-over-the-counter-medicines-abused-by-teen.html
- Mwanza, A et al. (2012). *Senior Physical science. Form 3*. Nairobi.Longhorn publishers.

Appendix 1: The Structure of a Laboratory Report

Title

Aim

Hypothesis

Materials

- _____
- _____
- _____
- _____

Procedure

1. _____
2. _____
3. _____
4. _____
5. _____

Observations

(data, results)

Data analysis

(tables or graphs)

(comparing hypothesis with results)

Conclusion



Appendix 2: The Periodic Table

^1_1H															^4_2He		
2.1															---		
^7_3Li	^9_4Be																
1.0	1.5																
$^{23}_{11}\text{Na}$	$^{24}_{12}\text{Mg}$																
0.9	1.2																
$^{39}_{19}\text{K}$	$^{40}_{20}\text{Ca}$	$^{45}_{21}\text{Sc}$	$^{49}_{22}\text{Ti}$	$^{51}_{23}\text{V}$	$^{52}_{24}\text{Cr}$	$^{55}_{25}\text{Mn}$	$^{56}_{26}\text{Fe}$	$^{59}_{27}\text{Co}$	$^{59}_{28}\text{Ni}$	$^{64}_{29}\text{Cu}$	$^{65}_{30}\text{Zn}$	$^{70}_{31}\text{Ga}$	$^{73}_{32}\text{Ge}$	$^{75}_{33}\text{As}$	$^{79}_{34}\text{Se}$	$^{80}_{35}\text{Br}$	$^{84}_{36}\text{Kr}$
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.8	1.8	1.9	1.6	1.6	2.0	2.4	2.8	3.0	
$^{86}_{37}\text{Rb}$	$^{88}_{38}\text{Sr}$	$^{89}_{39}\text{Y}$	$^{91}_{40}\text{Zr}$	$^{93}_{41}\text{Nb}$	$^{96}_{42}\text{Mo}$	$^{(98)}_{43}\text{Tc}$	$^{101}_{44}\text{Ru}$	$^{103}_{45}\text{Rh}$	$^{106}_{46}\text{Pd}$	$^{108}_{47}\text{Ag}$	$^{112}_{48}\text{Cd}$	$^{115}_{49}\text{In}$	$^{119}_{50}\text{Sn}$	$^{122}_{51}\text{Sb}$	$^{128}_{52}\text{Te}$	$^{127}_{53}\text{I}$	$^{131}_{54}\text{Xe}$
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5	2.6
$^{133}_{55}\text{Cs}$	$^{137}_{56}\text{Ba}$	$^{140}_{57}\text{Lu}$	$^{179}_{72}\text{Hf}$	$^{180}_{73}\text{Ta}$	$^{184}_{74}\text{W}$	$^{186}_{75}\text{Re}$	$^{190}_{76}\text{Os}$	$^{192}_{77}\text{Ir}$	$^{195}_{78}\text{Pt}$	$^{197}_{79}\text{Au}$	$^{201}_{80}\text{Hg}$	$^{204}_{81}\text{Ti}$	$^{207}_{82}\text{Pb}$	$^{209}_{83}\text{Bi}$	$^{210}_{84}\text{Po}$	$^{210}_{85}\text{At}$	$^{(222)}_{86}\text{Rn}$
0.7	0.9	1.3	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.8	1.9	2.0	2.2	2.4
$^{223}_{87}\text{Fr}$	$^{226}_{88}\text{Ra}$	$^{227}_{89}\text{Ac}$															
0.7	0.9	--															

Key

