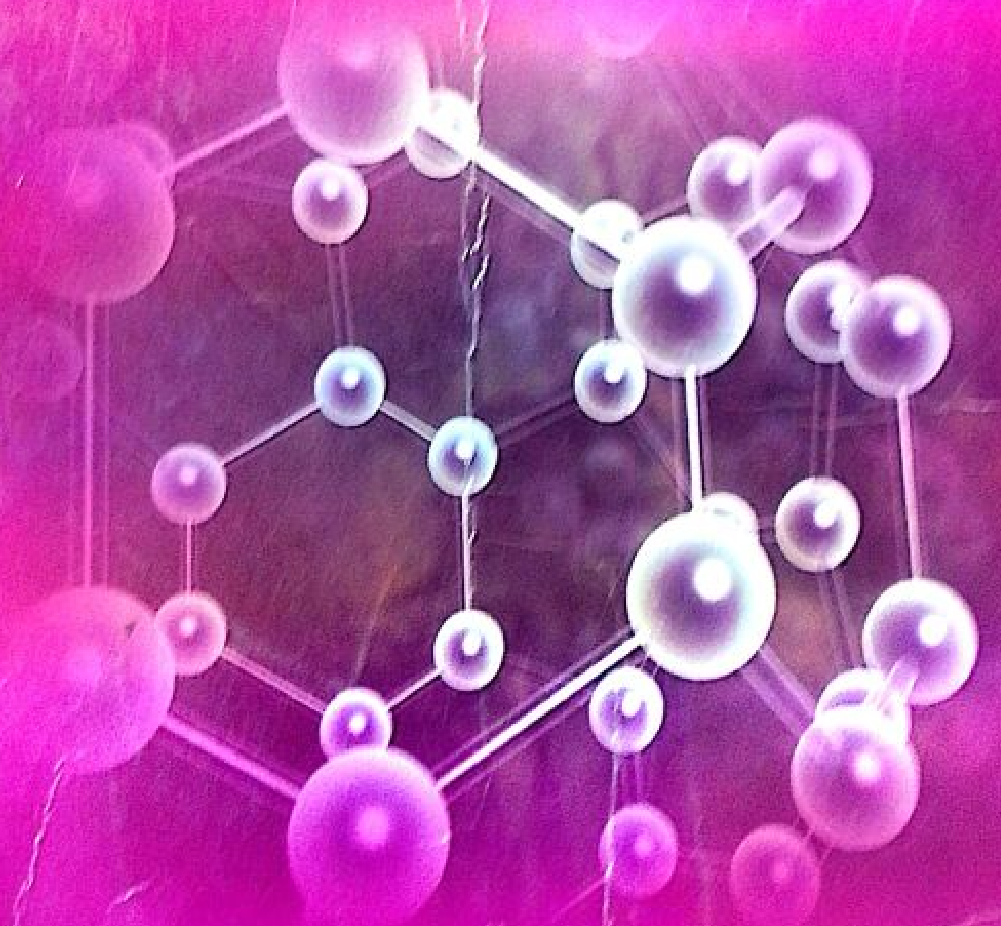
Baroque

New Lower Secondary School Curriculum

CHEMISTRY

Approved by NCDC and MoES

New Generation Books



BOOK 2

Senior Two Learner's Book

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Contents

Baroque Senior Two Chemistry Learner's Book has been developed in response to the new competence-based ***Lower Secondary Curriculum*** for Uganda. The curriculum was developed by the Ministry of Education and Sports (MoES) under the National Curriculum Development Centre (NCDC) and launched in 2020.

The book is a result of extensive research from several credible Chemistry resources and input from experienced teachers and experts.

Baroque Senior Two Chemistry Learner's Book entails;

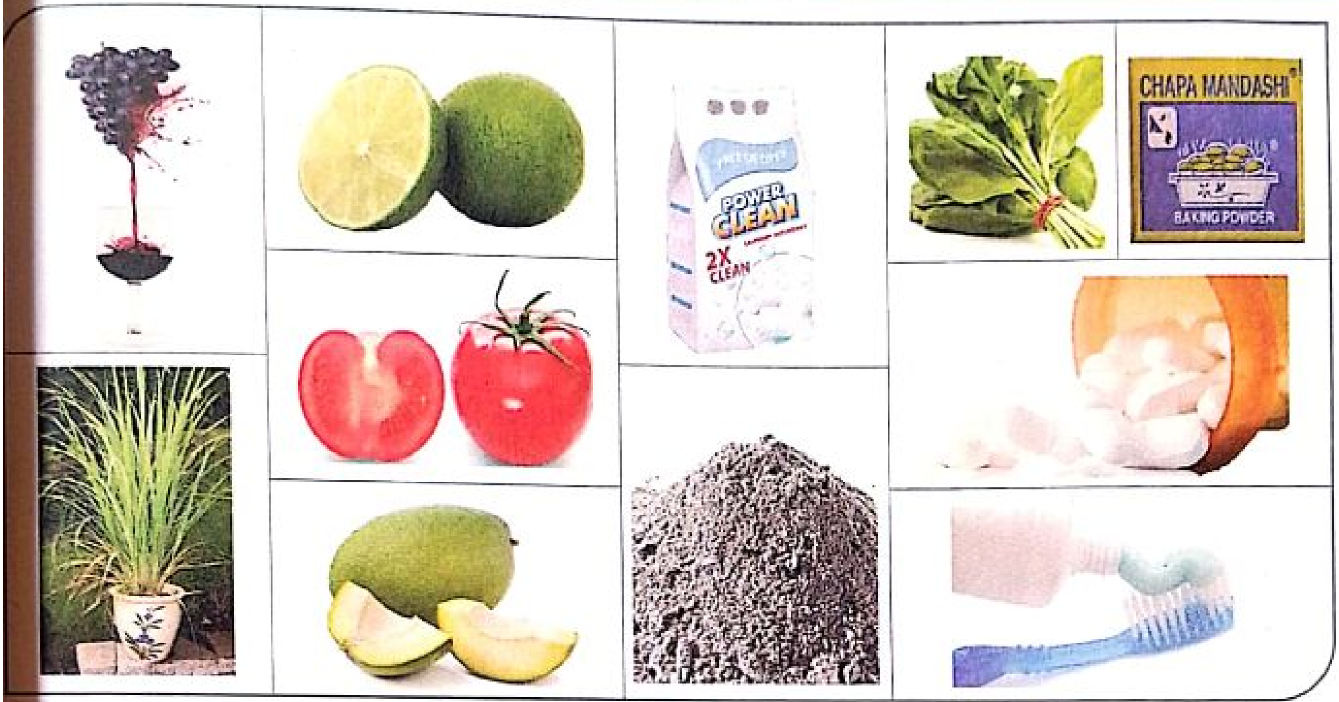
- an active competence-based and learner-centred approach,
- appropriate and accurate content,
- adequate and relevant activities and projects that trigger discovery, critical thinking, creativity, problem-solving and interactivity,
- acceptable, appropriate, standard and grammatically correct English, which encourages vocabulary development as well as correct representation of technical terms,
- accurate, relevant, clear, and adequate illustrations that enhance learning,
- intuitive methods, illustrations, activities, and projects that have been, explored to instill the principles of Chemistry.

In pursuit of a knowledge-based society, there is need for new generation learning books that are learner-centred, sufficiently researched and innovatively developed.

Baroque Senior Two Chemistry Learner's Book lays a firm foundation for learners who would like to pursue a career in Chemistry-related fields and seeks to equip all learners with the ability to apply Chemistry knowledge in day-to-day activities.

CHAPTER 1

Acids and Alkalis



Keywords

- acids
- alkalis
- bases
- neutralisation
- pH
- proton
- reaction
- precipitate
- lime water

By the end of this chapter, you should be able to:

a) recognise that locally available materials and substances are either acidic or alkaline.

b) understand the concept of ph as a measure of the strength of acids and alkalis.

c) understand the reaction between acids and alkalis.

Competency: You should be able to appreciate the properties and importance of acids, bases and salts in everyday life.

Introduction

Have you ever eaten raw mangoes, oranges or lemons, or tasted aloevera, okra or neem leaves? What did they taste like?

To many people, the word "acid" is associated with the liquid used in car batteries, that is known to recharge the battery and cause skin burns. The word acid comes from the Latin word ***"acidus"***, which means ***sour.*** In this chapter, you will appreciate the properties and importance of acids, alkalis/bases and salts in our daily life.

Figure 1.1: *Biting a raw mango*

Figure 1.2: *A boy eating okra*

1.1 Common Sources of Acids and Alkalis in our Daily Life

Figure 1.3: *Common sources of acids and alkalis*

Acids and alkalis are found in our day to day life. They are found in our environment, and some are man-made.

Some acids that occur naturally in plants and animals include; methanoic acid in bee and ant stings, citric acid in raw oranges and lemons, lactic acid in sour milk and carbonic acid in rain water. Sometimes, we add vinegar to food as a preservative because it contains ethanoic acid. Vinegar is obtained from rotting fruits; for example, pineapples.

Proteins are made up of amino acids, while apples contain vitamin C which is an acid, called ascorbic acid.

All acids contain at least one hydrogen atom, which is released as hydrogen ions when the acid is dissolved in water. It is these hydrogen ions in an aqueous solution that make the substance acidic.

Most vegetables, such as spinach, *Hibiscus canabinus* (locally known as malakwang), and amaranthus (dodo), contain alkalis. Toothpaste is an example of an alkali. It is used for brushing teeth because it neutralises the acid produced by bacteria that remain in the teeth.

All alkalis release hydroxide ions when dissolved in water. It is the hydroxide ions that make a substance behave like an alkali in an aqueous solution.

Note

An ***alkali*** is a soluble base.
A ***base*** is a substance that reacts with an acid to form salt and water only.

Activity 1.1(a): Classifying substances as either acids or alkalis / bases

In this activity, you will work in groups.

What you need

- lemon
- beakers
- raw mangoes
- sour milk
- okra
- baking soda
- African eggplant
- aloe vera
- a knife
- ash from banana peelings
- chapa mandashi

What to do

1. Wash the raw mangoes, African eggplant, aloe vera and okra provided.
2. Bite into each of the fruits provided and taste them. Touch to feel them.
3. Record the taste and feel in *Table 1.1*.

Table 1.1: Identifying acids and alkalis

(Put a tick (✓) for the correct verification and a cross (×) for the wrong verification)

Substance	Tastes sour	Tastes bitter	Feels slippery	Acid / Alkaline
Raw mangoes	✓	×	×	Acid
Lemon juice				
African eggplants				
Okra				
Banana peelings				
Chapa mandashi				
Tomato juice				
Aloe vera				

4. Cut the lemon with the knife provided, and taste it. Extract juice from the lemon as shown in *Figure 1.5.*

Figure 1.4: *Washing mangoes*

Figure 1.5: *Extracting lemon juice*

Note Wash your mouth after tasting each fruit or juice.

5. Dissolve the ash from banana peelings in water and touch the solution.
6. Repeat procedure 5 using baking soda, and chapa mandashi.
7. Classify the substances with the same taste and same feeling when touched, as either acids or alkalis.
8. Copy and complete *Table 1.1.*
9. Present your responses to the rest of the class.

Discussion Question

What is the general name for substances that:

(a) have a sharp and sour taste (b) are bitter and slippery?

Conclusion

After carrying out *Activity 1.1(a),* you have found out that substances which contain acids have the same physical properties in terms of taste. Substances which contain alkalis have the same taste and feel, which is different from those of acids.

For your understanding

Acids and bases definitely serve important functions, both inside and outside the scientific laboratory. In everyday life, acids and bases play a role in almost everything; from digestion of the food you eat, to the function of the medicine you take, and even to the cleaning products you use. Without acids and bases, many of the products in your home today would not be useful.

Indicators

DID YOU KNOW? Cars have several indicators with specific uses.

The indicators include: brake light, headlight, reversing light, hazard-warning light and the car horn. Find out their uses.

In Chemistry, there are also indicators which show specific colour changes when added to either acidic or alkaline solutions.

Figure 1.6: Motorcycle indicators

What is an indicator?

Since some substances are poisonous, it is not safe to taste them using a tongue. To know whether such substances are acidic, alkaline or neutral, we test them using indicators.

An indicator has one definite colour in an acidic solution and another definite colour in an alkaline solution.

Plant extracts are simple acid-base indicators. It is possible to use flower extracts to group different substances into acids or bases.

Such plant materials include; red or purple cabbage, red bougainvillea, hibiscus flower and blue morning glory.

Activity 1.1(b): Obtaining an indicator from plant extracts

In this activity, you will work in groups.

Safety precaution:

If it is necessary to heat the extract in order to concentrate the indicator, do not heat those containing the flammable solvents like ethanol or propanol, over an open flame. Use a water bath or a hotplate.

What you need:

- a mortar and a pestle
- beakers
- a dropper
- propanol or ethanol or warm water
- a measuring cylinder
- petals of morning glory / bougainvillea / hibiscus / red cabbage

What to do:

1. Cut as many petals as possible from one type of flowers, like red or purple cabbage, into small pieces.

2. Grind as many petals as possible in a mortar using a pestle to obtain a paste as shown in *Figure 1.7*.

Figure 1.7: *Preparation of a plant extract indicator*

3. Using a measuring cylinder, add 5 cm^3 of ethanol and continue grinding until a deep colour of the flower extract appears.
4. Sieve the mixture into a small beaker and cover it with a filter paper. This is your indicator, keep it safely.

Discussion Questions

Figure 1.7 shows the process of preparation of an extract.

1. In your groups, discuss why:
   a) propanol or ethanol is used in *Activity 1.1(b)*.
   b) coloured flowers are preferred.
   c) the solution was covered after sieving.
2. What is the importance of grinding the red cabbage leaves?

Conclusion

The red cabbage juice extracts work as indicators. They show different colour changes in acidic, basic and neutral media.

Activity 1.1(c): Classifying substances as acids or bases using a plant extract indicator

What you need:

- soap solution
- vinegar
- lemon / orange juice
- baking powder solution
- solution of banana ashes
- tea leaves
- distilled water
- test tubes
- sodium chloride (common salt) solution
- a test tube rack
- a measuring cylinder

What to do:

1. Using a measuring cylinder, measure 2 cm^3 of lemon juice and pour it into a test tube.
2. Add 3 drops of the indicator that you prepared in *Activity 1.1(b)*.
3. Observe and record the colour change in *Table 1.2*.

Table 1.2: Effect of the flower extract on different substances

Substance	Colour change when the indicator is added	Conclusion (acidic, basic or neutral)
Lemon juice		
Vinegar		
Soap solution		
Tea leaves		
Baking powder solution		
Solution of banana ashes		
Sodium chloride solution		
Distilled water		

4. Repeat procedures 1 and 2 using, vinegar, soap solution, tea leaves, baking powder solution, solution of banana ashes, sodium chloride solution, and distilled water.
5. Record your results in *Table 1.2* and conclude on the nature of the given solutions.

Discussion Questions

In groups, discuss the following:

1. What was the colour of the flower extract you used?
2. Which substance(s) turned the colour of the flower extract the same way that lemon did?
3. Which substance(s) turned the colour of the flower extract the same way that the baking powder solution did?
4. What name can be given to the flower extract. Why?
5. What name can be given to those substances that had an effect on the flower extract in the same way as;

   a) lemon juice b) distilled water c) baking powder solution?

Conclusion

The flower extracts showed different colour changes in acidic and basic solutions. It is reasonable to suggest plant flowers as an alternative source of indicators. They are simple, economical, and readily available for identifying acids and bases.
Therefore, Chemists should exploit the importance of brightly coloured flowers within their environment. This shows flexibility of the natural environment in science teaching.

For your understanding

Acids and bases give different colour changes when the same flower extract is added to them.

Commercial acid-base indicators

When indicators prepared from plant materials are stored in a laboratory for future use, they keep changing in composition and properties as time goes on. These changes give inconsistent results. For this reason, commercial indicators are prepared because they give more consistent results. Examples of commercial indicators include: methyl orange, litmus, phenolphthalein and universal indicator as shown in *Figure 1.8.*

Figure 1.8: *Commercial indicators and their effects on acidic and alkaline solutions*

The universal indicator gives precisely more accurate results.

In *Activity 1.1(d),* you will find out how litmus paper behaves as an indicator.

Activity 1.1(d): Investigating the effect of different substances on litmus paper

In this activity, you will work in groups.

What you need

- litmus paper (red and blue)
- a beaker
- a dropper
- carbonated water
- diluted liquid soap
- tea leaves
- distilled water
- soda ash solution
- diluted bleach
- a measuring cylinder
- wood ash solution
- vinegar
- lemon juice

What to do

1. Using a measuring cylinder, measure 2 cm^3 of lemon juice and pour it into a beaker.
2. Dip blue and red litmus papers into the lemon juice in the beaker, one at a time.
3. Record your observations in *Table 1.3.*
4. Repeat your activity with vinegar, diluted bleach, diluted liquid soap, tea leaves, wood ash solution, soda ash solution, carbonated water, and distilled water.
5. Individually, copy and complete *Table 1.3.*

Table 1.3: Effect of different substances on litmus paper

Substance	Colour change when blue litmus paper is dipped	Colour change when red litmus paper is dipped	Conclusion (acidic, basic or neutral)
Lemon juice			
Vinegar			
Diluted bleach			
Diluted liquid soap			
Tea leaves			
Solution of wood ash			
Soda solution			
Carbonated water			
Distilled water			

6. Present your findings to the class.

Conclusion

Neutral substances have no effect on litmus paper.

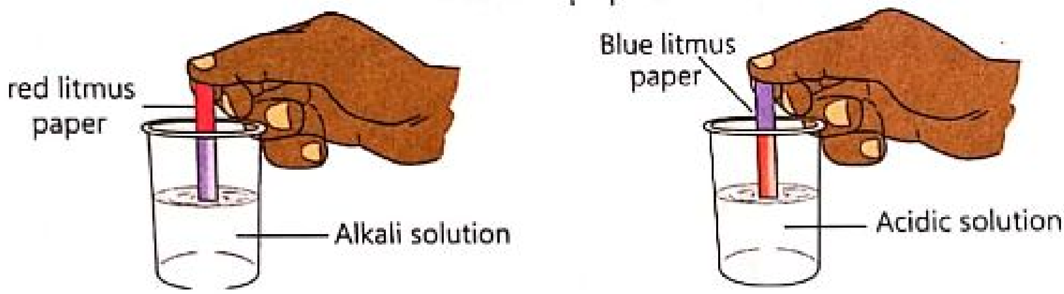


Figure 1.9: Effect of acids and alkalis on litmus paper

In acidic solutions, the blue litmus paper turns red; while in basic solutions, the red litmus paper turns blue, as shown in *Figure 1.9*.
You should also carry out the following experiment.

Using phenolphthalein as an indicator

Is it magic?

Pour about 100 ml of water into a beaker and add 3 drops of phenolphthalein indicator. You will obtain a colourless solution, just like water, as shown in *Figure 1.10*.

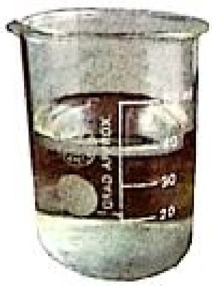


Figure 1.10: The colour after addition of phenolphthalein indicator

Phenolphthalein indicator is colourless.
With a dropper, add some drops of ammonia solution.

Safety precaution:
To avoid the irritating vapour of ammonia, carry out this experiment in an open space with open windows. Keep a distance of at least 2 metres away from the container of ammonia solution.

After adding some drops, you will see the solution turning pink as shown, in *Figure 1.11*. This situation is quite surprising because all the liquids are colourless and transparent, like water.

Figure 1.11: After addition of ammonia solution

Now, to the pink fluid you have obtained, add some drops of vinegar. What happens?
The liquid becomes colourless again, as shown in *Figure 1.12*.

Figure 1.12: After addition of vinegar

Phenolphthalein becomes pink when the pH of the solution is alkaline, and colourless when the solution is neutral or acidic. Ammonia solution is alkaline, water is neutral, while vinegar is acidic.

Assignment 1.1

1. In groups, mention the colour changes that occur when phenolphthalein and methyl orange indicators are separately added to;
   a) lemon juice
   b) soda ash solution
2. What general conclusion can you derive from your observations?

You will observe that different indicators exhibit different colours in acidic and basic media, as shown in *Table 1.4*.

Table 1.4: Colours exhibited by different indicators in acidic and alkaline media

Name of indicator	Colour in acidic solution	Colour in alkali
Litmus	Red	Blue
Methyl orange	Red	Yellow
Phenolphthalein	Colourless	Pink

1.2 Strength of Acids and Alkalis

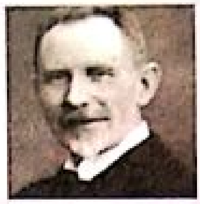
It is safe to taste lemon juice, but not sulphuric acid found in car batteries (since it corrodes and burns the skin). The degree of acidity in sulphuric acid is much higher than that in lemon juice. In the same way, sodium hydroxide solution is more basic than baking soda solution.

Sulphuric acid is a mineral acid and other examples of mineral acids are given in *Figure 1.13*. The degree of alkalinity or acidity of a solution is called the ***pH*** of that solution.

Figure 1.13: Mineral acids

DID YOU KNOW?

Soren Sorensen, a Danish Chemist, in 1909 introduced the concept of pH as a convenient expression of acidity or alkalinity of a given solution.



To measure the pH of a solution accurately, a scale of numbers ranging from 0 to 14 was developed; which is called a ***pH scale.***

The degree of basicity or acidity of a solution is represented on the pH scale by a number and a colour.

The indicators so far used in the previous activities and assignments do not tell the pH of a solution. The ***pH*** of a solution can be obtained using a ***universal indicator.***

The universal indicator is a mixture of other indicators and displays a range of colours, each of which, matches with a shade of colours and a number on the pH scale, as shown in *Figure 1.14.*

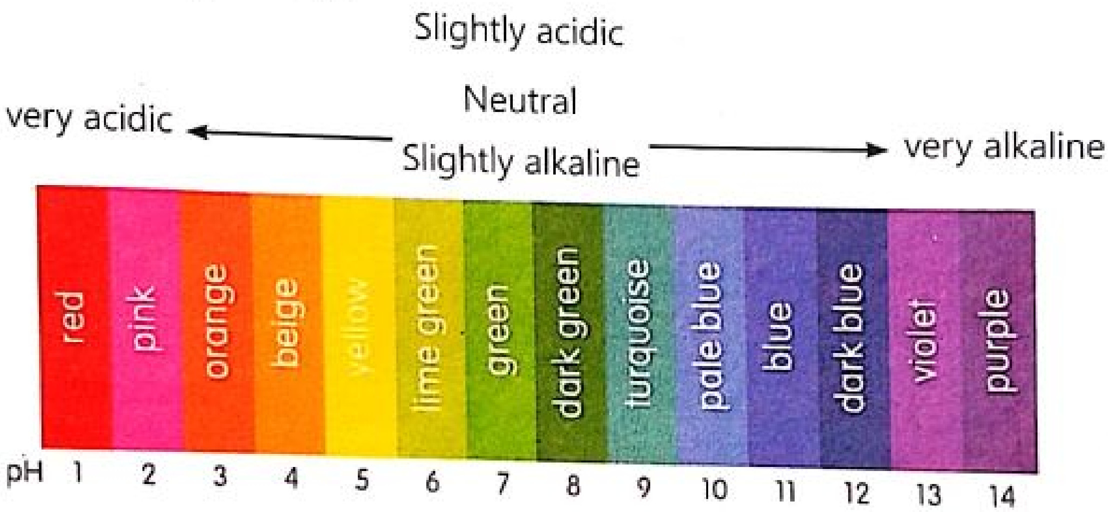


Figure 1.14: *The pH scale*

The pH scale reflects the relative amount of free hydrogen and hydroxide ions in a solution. The more the hydrogen ions, the more acidic and the smaller the pH. The greater the amount of hydroxide ions, the greater the alkalinity and the higher the pH. The middle number on the scale, 7, is used to describe substances which are neutral. For example, pure water has a pH of 7.

Strong acids have pH values which are less than 4 and are mainly commercial acids, like sulphuric acid, hydrochloric acid and nitric acid. They are referred to as mineral acids because they are obtained from minerals or non-living things. The rest of the acids are weak and these are mainly the naturally occurring acids, such as citric acid, and ethanoic acid.

Strong bases like sodium hydroxide and potassium hydroxide have pH values which are above 11. Weak bases have a pH value above 7 but below 11, they include: soap solution, soda ash solution, okra, and cabbage solution, among others.

Determination of pH

pH is the most common of all analytical measurements in industrial processing. Most solutions require a specific pH for certain reactions to take place, especially in the food and beverage industries, cosmetics and pharmaceutical sector. In industries setting pH measurements determines the quality of a product and monitors the chemical reaction. In *Activity 1.2,* by use of the universal indicator and the pH chart you will get the pH values of various solutions.

Activity 1.2: Determining the pH of a solution

In this activity, you will work in groups.

What you need:

- universal indicator
- test tubes
- a test tube rack
- distilled water
- dilute hydrochloric acid
- droppers
- soda ash solution
- lemon juice
- diluted bleach
- sodium hydroxide solution
- diluted liquid soap
- wood ash solution
- universal indicator
- a measuring cylinder

What to do:

1. Using a measuring cylinder, measure 2 cm^3 of dilute hydrochloric acid and pour it into a beaker.
2. Add 2 drops of universal indicator.
3. Note the colour of the solution.
4. Using the pH scale in *Figure 1.14,* record in *Table 1.5,* the pH corresponding to the colour of the solution.
5. Repeat procedures 1, 2 and 3 using sodium hydroxide solution, lemon juice, diluted liquid soap, soda ash solution, diluted bleach, wood ash solution, and distilled water.
6. Present your results in *Table 1.5.*

Table 1.5: Colour change and the corresponding pH of different substances

Substance	Colour change	Corresponding pH
1. Dilute hydrochloric acid		
2. Sodium hydroxide solution		
3. Lemon juice		
4. Diluted liquid soap		
5. Soda ash solution		
6. Diluted bleach		
7. Wood ash solution		

7. Present your results to the rest of the class.

Discussion Questions

1. What is the range of pH values of acids?
2. What is the range of pH values of bases?
3. What is the pH value of a neutral solution?

The universal indicator is a mixture of several indicators and it shows a range of colours in acids and bases, depending on the degree of acidity or alkalinity. Some acids are more acidic than others, while some bases are more basic than others. pH can be measured directly by using a *pH meter* as shown in *Figure 1.15.*

***Figure 1.15:** Measuring the pH of a solution*

1.3 Reaction Between Acids and Alkalis

Acid-base reactions form a large part of Chemistry, both in the biological and the Physical Chemistry fields. Such reactions control the pH, which is one of the important aspects a chemist uses in Biochemistry and Physical Chemistry processes.

Activity 1.3(a): Reacting an acid and a base

In this activity, you will work in groups.

What you need:

- lemon juice
- universal indicator
- sodium carbonate solution
- test tubes
- sodium hydroxide solution
- dilute hydrochloric acid
- a measuring cylinder

What to do:

1. Using a measuring cylinder, measure 2 cm^3 of lemon juice and pour it into a test tube.
2. Test the solution with a red litmus paper.
3. Add soda ash solution, a drop at a time, and keep testing the solution with litmus paper until it just turns blue.
4. Add 2 drops of universal indicator to the resultant solution in 3 above, and note the colour.
5. Repeat procedures 1, 2, 3 and 4 using sodium hydroxide solution and dilute hydrochloric acid.

Discussion Questions

1. What do you observe when sodium carbonate solution is added to lemon juice?
2. What conclusion can you make regarding the colour of the resultant solution and its pH?
3. What is the colour change when the universal indicator is added to the resultant solution?

Conclusion

The reaction between an acid and an alkali is called a ***neutralisation reaction.*** This is because the products are neutral in terms of pH. An acid reacts with a base to give a salt and water only.

Activity 1.3(b): How can too much acid in the stomach be reduced?

What you need:

- dilute hydrochloric acid
- universal indicator
- baking powder solution
- test tubes
- measuring cylinder

What to do:

1. Using a measuring cylinder, measure 5 cm^3 of dilute hydrochloric acid and pour it into a test tube.
2. Add 2 drops of the universal indicator.
3. Add baking powder solution, a little at a time, until the colour turns green.

Discussion Question

In groups, discuss the effect of baking powder solution on dilute hydrochloric acid solution.

Conclusion

Dilute hydrochloric acid in our stomach provides a medium for the digestion of food. However, when the stomach secretes excess acid, it causes indigestion. The acid burns the lining of the stomach, thereby causing peptic ulcers. Alkaline substances neutralise the excess acidity in the stomach. Some substances that are used to neutralise acidity in daily life are shown below.

Health benefit!	
	Milk, sodium bicarbonate and magnesium hydroxide are anti-acids used to relieve heartburn and indigestion, or to make your blood or urine less acidic in certain conditions. This is because they are alkaline in nature.
	Bee stings are acidic. They are also neutralised by rubbing a solution of banana peelings or baking powder on the site which the bee stung.

Properties of acids and bases

To investigate the properties of acids and bases, we shall consider the action of acids and bases on different substances.

Activity 1.3(c): Investigating the action of dilute acids on metals

This activity is to be done in groups.

What you need:

- wood splints
- zinc granules
- copper metal
- measuring cylinder
- heat source
- iron wool
- magnesium ribbon
- dilute hydrochloric acid
- lead metal

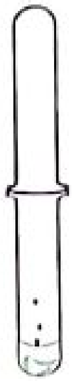


Figure 1.16: Reaction of magnesium with dilute hydrochloric acid

What to do:

1. Put 2 to 3 pieces of magnesium ribbon in a test tube.
2. Using a measuring cylinder, measure 2 cm^3 of hydrochloric acid and add it to the test tube, as shown in *Figure 1.16*.
3. Test the gas produced using a burning splint as shown in *Figure 1.17*.
4. Repeat procedures 1, 2 and 3 using zinc granules, iron, lead, and then copper.
5. Record your observations in *Table 1.6*.

***Figure 1.17:** Testing for the gas produced*

Table 1.6: Action of dilute hydrochloric acid on metals

Metal	Observations
Magnesium	
Zinc	
Iron	
Lead	
Copper	

Discussion Questions

1. Write a general word equation for the reactions that occured.
2. Name another acid which can be used, instead of hydrochloric acid.

Conclusion

Most metals react with dilute acids by displacing hydrogen from the acid, resulting into the evolution of hydrogen gas. The metals react with the dilute acids at different rates to form corresponding salts and hydrogen gas. For example, magnesium reacts very fast with dilute hydrochloric acid to form magnesium chloride (a salt) and hydrogen gas. Lead metal reacts very slowly with the same acid to form lead(II) chloride (a salt) and hydrogen gas.

Activity 1.3(d): Investigating the action of dilute acids on metal oxides

This activity is to be done in groups.

What you need:

- test tubes
- copper(II) oxide
- calcium oxide
- lead(II) oxide
- zinc oxide
- magnesium oxide
- dilute hydrochloric acid
- a measuring cylinder

What to do:

1. Using a measuring cylinder, measure 2 cm^3 of dilute hydrochloric acid and pour it into a test tube.
2. Add copper(II) oxide powder, a little at a time, to the acid in the test tube.
3. Warm the mixture gently.
4. Record your observations in *Table 1.7.*
5. Write a word equation for the reaction that took place.
6. Repeat procedures 1 to 4 using magnesium oxide, zinc oxide, calcium oxide, and lead(II) oxide.
7. Copy and complete *Table 1.7*, and write down your observations.

Table 1.7: Action of dilute hydrochloric acid on metal oxides

Metal Oxide	Observations
Copper(II) oxide	
Magnesium oxide	
Zinc oxide	
Calcium oxide	

Discussion Questions

1. Which two other acids react in the same way as dilute hydrochloric acid?
2. What is the name given to a reaction between a metal oxide and an acid?

Conclusion

From your observations, you realise that ,when warm dilute acids are added to metal oxides, they form a salt and water.

Equation:
Dilute acid + metal oxide → salt + water

In some cases, a clear solution was formed as the product. This means that the salt formed is soluble in water.

A cloudy solution, as a product of the reaction, means that the salt formed is insoluble in water. Such a cloudy solution is called a ***precipitate.***

For your understanding

A base is a metal oxide, hydroxide or carbonate that reacts with an acid to form a salt and water only.

Activity 1.3(e): Investigating the action of dilute acids on solid carbonates and hydrogen carbonates

This activity is to be done in groups.

What you need:

- test tubes
- delivery tube fitted with a cork
- calcium carbonate
- dilute hydrochloric acid
- lime water
- sodium carbonate
- sodium hydrogen carbonate
- copper(II) carbonate
- lead(II) carbonate
- a measuring cylinder

What to do:

1. Put 2-3 pea-sized pieces of marble chips in a test tube.
2. Using a measuring cylinder, measure 2 cm^3 of dilute hydrochloric acid, and add it to the carbonate. Assemble the apparatus as shown in the *Figure 1.18*.

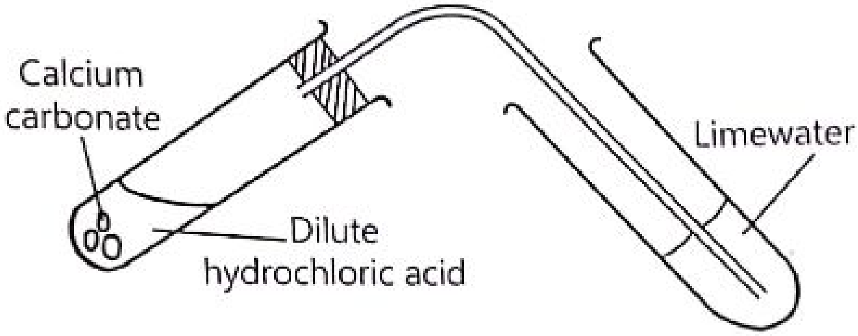


Figure 1.18: Action of dilute acids on solid carbonates and hydrogen carbonates

3. Pass the gas evolved through lime water.
4. Put a burning splint in the gas produced.
5. Record your results in 2, 3 and 4 above in *Table 1.8*.
6. Repeat the experiment using sodium carbonate, sodium hydrogen carbonate, copper(II) carbonate, and lead (II) carbonate; in the place of calcium carbonate.

Table 1.8: Action of dilute hydrochloric acid on carbonates and hydrogen carbonates

Carbonate / Hydrogen carbonate	Observations
Calcium carbonate	
Sodium carbonate	
Sodium hydrogen carbonate	
Copper(II) carbonate	
Lead (II) carbonate	

Discussion Question

Write a word equation for the reaction between dilute hydrochloric acid and calcium carbonate.

Acids react with metal carbonates and metal hydrogen carbonates to produce a salt, water, and carbon dioxide.

Equation:

Acid + carbonate → salt + water + carbon dioxide

Carbon dioxide forms a white precipitate with lime water. This is because carbon dioxide reacts with limewater to form calcium carbonate, which is an insoluble salt. The chemical name of ***limewater*** is ***calcium hydroxide solution.***

Assignment 1.2

In groups, research about the uses of acids and bases, and present your findings to the rest of the class.

For your understanding

Tooth decay starts when the pH of the mouth is lower than 5.5. The tooth enamel, made up of calcium phosphate, is the hardest substance in the body. It does not dissolve in water, but corrodes when the pH in the mouth is below 5.5. Bacteria present in the mouth produce acids by degradation of sugars and food particles which remain in the mouth after eating. The best way to prevent this is to clean the mouth after eating food, using toothpaste which is alkaline or basic and thus, neutralising the acid, thereby preventing tooth decay (*shown in Figure 1.19*).

***Figure 1.19:** Tooth decay*

For your understanding

Alkaline solutions contain hydroxide ions which neutralise hydrogen ions in acidic solutions.

Activity 1.3(f): Applying neutralisation reactions

In groups, research and discuss the other applications of neutralisation reactions in everyday life.

ICT Activity

a) In groups, use the internet or library source to research about the examples of strong acids, strong bases, weak bases, and weak acids. Using a suitable software, design a table to show those examples.

b) Present your findings to the rest of the class, using a suitable presentation software.

Sample Activity of Integration

Farmers from two villages in Kabale District, a place popular for growing vegetables which grow well in neutral soils, are experiencing low yields of vegetables. On close observation, one village is faced with the spread of *Lantana camara* which is shown in *Figure 1.20*, an indicator that the soils are acidic. The other village is experiencing the spread of *Lavandula spica* (Lavender) shown in *Figure 1.21*, which is an indicator that the soils are alkaline. The Environmental Officer has been contacted, and has mobilised a community meeting of the two villages.

Figure 1.20: Lantana Camara

Figure 1.21: Lavandula spica (Lavender)

Task: As a student of Chemistry, She/he has asked you to develop and present a speech to the farmers, on the state of their land and how their yields of vegetables can be improved by using the available inputs. Write a speech that you will present.

Chapter Summary

In this chapter, you have learnt that:

- An acid tastes sour, has a pH below 7, turns blue litmus paper red, and dissolves in water to produce hydrogen ions as the only positively charged ions.
- Some acids are found naturally in plants and animals.
- Bases have a bitter taste, are slippery, turn red litmus blue and have a pH above 7
- Alkalis are soluble bases and produce hydroxide ions when dissolved in water.
- An indicator distinguishes between acids and bases.
- Indicators can be prepared from plant extracts using flower petals.
- Commonly used commercial indicators include methyl orange, litmus paper and phenolphthalein indicator.
- The pH scale is used to show the degree of acidity or basicity of a solution, and it ranges from pH 0 to pH 14.
- The universal indicator shows the exact pH of a solution by use of colours and numbers on the pH scale.
- Acids react with bases to form a salt and water only, and the process is called neutralisation.
- Hydrogen gas explodes with a 'pop' sound when treated with a burning splint.
- Carbon dioxide turns lime water milky.
- Hydrochloric acid and nitric acid are very strong mineral acids.

Review Exercises:

1. a) Describe what an acid is.
   b) Name the acid found in each of the following substances:
      i) Tea
      ii) Lemon juice
      iii) Urine
      iv) Digestive juice in the stomach
   c) Describe the uses of acids in everyday life.
2. a) Describe what an alkali is.
   b) A senior two student added some dilute hydrochloric acid, a little at a time, to an alkaline solution of sodium hydroxide in a conical flask.
      i) Write the equation for the reaction between hydrochloric acid and sodium hydroxide.
      ii) Name the products formed.
   c) Describe the uses of bases in everyday life.
3. You are required to carry out an experiment that requires an acid -base indicator. In case your school laboratory does not have an acid-base indicator, explain how you would improvise one, to use in your experiment.
   a) State one importance of an acid-base indicator.
   b) Name two indicators commonly used in the laboratory.
4. Discuss the importance of pH and the pH scale in our daily life.
5. a) Phenolphthalein is an indicator. What does it mean?
   b) What does the pH value tell you about the pH of a solution?
6. The diagram below shows the pH scale. The scale has been developed to measure and compare the acidity or alkalinity of a solution.

pH number	**0**	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**

   a) i) How can you determine the pH of a solution using the pH scale above?
      ii) Pure water has a pH of 7 or it is said to be neutral. Why?
   b) i) Explain why it is recommended to use lemon grass and / or neem to clean your teeth after a meal.
      ii) The taste of milk can be described as sour but in certain proportions, a mixture of milk and livestock salt (also known as "magadi", "amakadi", "kihonde", or "kisura") gives a neutral solution. Suggest the pH value of livestock salt.

CHAPTER 2

Salts

Keywords

- equation
- neutralisation
- salt
- solubility

By the end of this chapter, you should be able to:

a) be familiar with and be able to carry out neutralisation reactions to prepare salts

b) know and appreciate the uses of common salts in everyday life.

Competency: You should be able to appreciate that acids and alkalis form salts.

Introduction

Figure 2.1(a) : Fertilisers in use

Figure 2.1(b): Use of common salt in preserving fish

We always have common salt added to our food to give it taste. Just imagine eating food without salt! Common salt or table salt is, actually, sodium chloride. Salts are used in the manufacture of fertilisers and preservation of fish as shown in *Figures 2.1 (a) and (b).*

What do we really mean (in Chemistry) when we refer to something as a salt? The word "salt" has a wider meaning in Chemistry. Salts are a large family of compounds, many of which are quite different in colour. Some salts are sour, others are bitter, and others are poisonous. Some salts are soluble in water, some are sparingly soluble, while others are insoluble.

Fun joke:
What do you call it when salt says hello to pepper? It is season greetings.

2.1 Preparation of Salts

Some salts exist naturally and can be extracted from their rocks using different methods. For example, sodium chloride exists naturally in Lake Katwe and Lake Magadi. Other salts occur naturally and are found in plants and animals, and in small traces, such as calcium phosphate.

Naturally existing salts contain a lot of impurities and it becomes very costly to obtain them in pure form. Pure salts, therefore, are manufactured in industries by the action of dilute acids on either a base, or a metal, and are referred to as ***inorganic salts.***

a) Action of an acid on a carbonate

In Chapter 1, you learnt that acids react with carbonates and hydrogen carbonates to form a salt, water, and carbon dioxide. In *Activity 2.1 (a)*, you will prepare a salt from sodium bicarbonate.

Activity 2.1(a): Reacting baking soda and vinegar

In this activity, you will work in groups.

What you need:

- a plastic water bottle or a test tube
- vinegar
- a spatula
- a balloon
- baking soda (sodium bicarbonate)

Figure 2.2: Baking powder

Figure 2.3: Vinegar

What to do:

1. Half fill a plastic water bottle or a test tube with vinegar.
2. Put one spatula endful of baking soda into a balloon.
3. Fix the balloon over the mouth of the bottle and carefully empty the baking soda into the vinegar. Shake the contents gently so as to mix.
4. Observe and report on the reaction.

Discussion Questions

1. What did you see happening to the balloon?
2. What else was formed when baking soda reacted with vinegar?

Conclusion

Sodium bicarbonate reacted with ethanoic acid in vinegar to form carbon dioxide, sodium ethanoate and water.

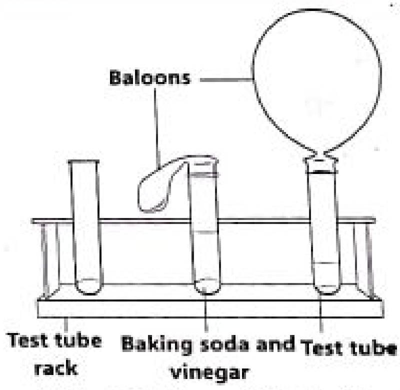


Figure 2.4: A balloon filled with gas

b) Action of an acid on a hydroxide

Previously, you learnt about the reactions of acids and bases. In Activity 2.1 (b), you will prepare a salt from the reaction of an acid and a base.

Activity 2.1(b): Reacting sodium hydroxide and dilute hydrochloric acid

In this activity, you will work in groups.

What you need:

- 2 conical flasks
- a burette
- 2 beakers
- a funnel
- an evaporating dish
- sodium hydroxide solution
- a measuring cylinder
- dilute hydrochloric acid
- a tripod stand and a wire gauze
- phenolpthalein indicator

What to do:

1. Using a measuring cylinder, measure 25 cm^3 of dilute sodium hydroxide solution pour it into a conical flask.
2. Add 2 drops of phenolpthalein indicator and shake the mixture.
3. Note the colour of the resultant solution.
4. Add dilute hydrochloric acid into the burette and note its initial volume.
5. Run dilute hydrochloric acid from the burette into the mixture while shaking, until the solution just changes to colourless.
6. Record the volume of the acid added to the alkali.
7. Repeat procedures 1, 4 and 5 using the same volume of the alkali and the acid used in the first experiment.
8. Transfer the contents of the conical flask to an evaporating dish and evaporate to dryness, as shown in *Figure 2.6*.
9. Record your observations.

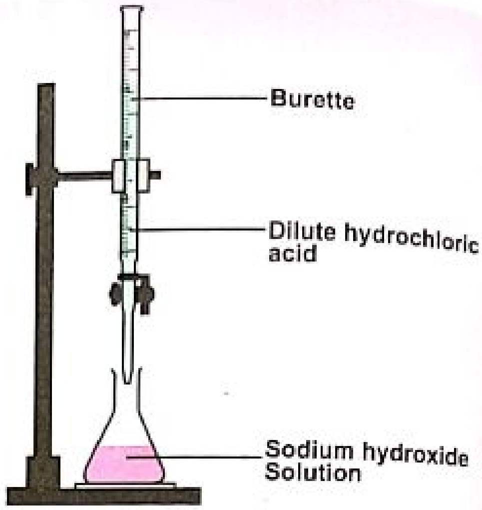


Figure 2.5: Preparation of sodium chloride

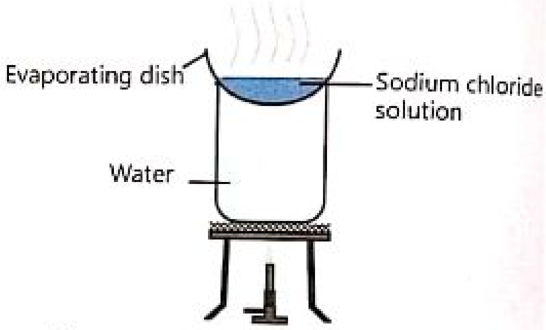


Figure 2.6: Evaporating to dryness

The reaction between sodium hydroxide solution and dilute hydrochloric acid is used to prepare sodium chloride salt in laboratories. Sodium chloride is referred to as ***common salt***.

This method is used to prepare salts of sodium, potassium and ammonia. However, sodium chloride occurs naturally in Lake Katwe in Kasese district and it is mined as shown in *Figure 2.7*. Lake Katwe also has the potential to produce other useful salts like sodium carbonate, sodium sulphate, and potassium sulphate, which are removed as impurities in the process of refining salty water to produce common salt.

Equation:

sodium hydroxide + hydrochloric acid → sodium chloride + water

Figure 2.7: Salt mining at lake Katwe

Salts of potassium, sodium and ammonia are prepared by the action of a dilute acid on a hydroxide solution in presence of an indicator which detects the point of neutralisation, since their corresponding hydroxides and oxides are very soluble in water. On noting the volumes to be used, the same quantities of the acid and the base are mixed without the indicator, and evaporated to obtain the salt.

Assignment 2.1

1. Write a word equation for the reaction between sodium hydroxide solution and dilute hydrochloric acid.
2. What term is used to refer to the reaction between sodium hydroxide solution and hydrochloric acid?

c) Action of an acid on an oxide

Magnesium sulphate occurs naturally as a mineral kieserite, $MgSO_4$. This is dissolved in hot water and the purified sulphate is obtained by crystallisation from the solution. You can make magnesium sulphate crystals in the laboratory by reacting magnesium oxide with dilute sulphuric acid.

Activity 2.1(c): Preparing magnesium sulphate from magnesium oxide and dilute sulphuric acid

In this activity, you will work in groups.

What you need:

- a source of heat
- a measuring cylinder
- dilute sulphuric acid
- magnesium oxide
- a glass rod
- an evaporating dish
- 2 beakers
- filter papers
- a filter funnel
- a spatula
- a tripod stand
- a wire gauze
- a conical flask

What to do:

1. Using a measuring cylinder, measure 50 cm^3 of dilute sulphuric acid and transfer it into a glass beaker.
2. Gently heat the solution to 60°C.
3. Using a spatula, add magnesium oxide powder, a little at a time, to the acid, and with constant stirring, keep adding the powder, until the unreacted magnesium oxide begins to settle at the bottom of the beaker.
4. Filter the mixture that is formed in the beaker.
5. Transfer the filtrate, which is magnesium sulphate solution, into an evaporating dish, and heat gently to evaporate off the water and form a concentrated solution.
6. Allow the solution to cool in order for magnesium sulphate crystals to form, and then filter.

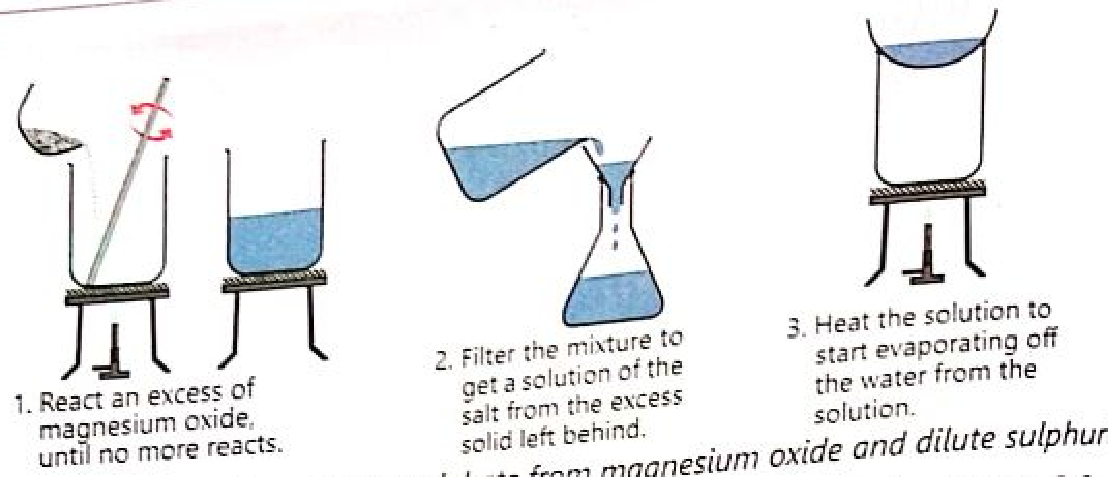


Figure 2.8: Preparation of magnesium sulphate from magnesium oxide and dilute sulphuric acid

7. Dry the crystals formed, between two filter papers as shown in *Figure 2.9.*

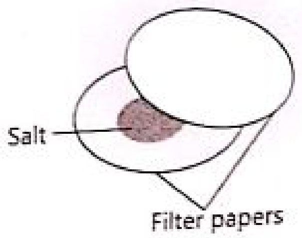


Figure 2.9: Drying crystalline substances between filter papers

Discussion Questions

Using *Activity 2.1 (c)*, answer the following questions.

1. What did you observe throughout the experiment?
2. Individually, discuss and explain to the rest of the class why;
   a) the acid was warmed
   b) magnesium oxide was added until in excess.
   c) the solution was cooled.
3. Write a word equation for the above reaction.
4. What do we call the reaction in (3) above?

Other salts which can be similarly prepared by the action of dilute sulphuric acid on their corresponding oxides include Copper(II) sulphate and Zinc sulphate.

d) Action of an acid on carbonates

In *Activity 2.1 (d)*, you are going to react dilute nitric acid and lead(II) carbonate.

Activity 2.1(d): Investigating how dilute nitric acid reacts with lead(II) carbonate

In this activity, you will work in groups.

What you need:

- a source of heat
- a measuring cylinder
- dilute nitric acid
- lead(II) carbonate
- a glass rod
- an evaporating dish

- 2 beakers
- filter papers
- a filter funnel
- a spatula
- a tripod stand
- a wire gauze
- a conical flask

Safety precaution:

You should keep in mind that lead(II) nitrate is very toxic, soluble, and can be absorbed through the skin!

What to do:

1. Using a measuring cylinder, measure 50 cm^3 of dilute nitric acid and transfer it into a glass beaker.
2. Using a spatula, add lead(II) carbonate powder, a little at a time, to the acid, and with constant stirring, keep adding the powder until the unreacted lead(II) carbonate begins to settle at the bottom of the beaker.
3. Filter the mixture that is formed in the beaker.
4. Transfer the filtrate into an evaporating basin and heat gently until the solution is saturated, as shown in *Figure 2.10*.

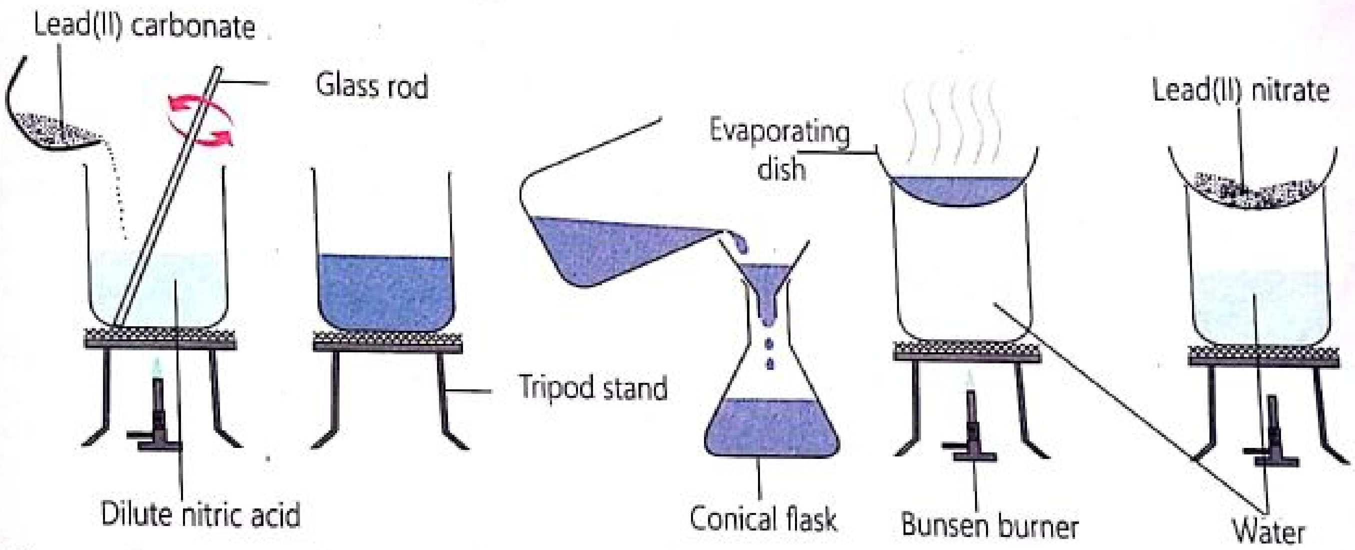


Figure 2.10: *Set up of the apparatus for the reaction of dilute nitric acid and lead(II) carbonate*

5. Allow the solution to cool in order for lead(II) nitrate crystals to form, and then filter.
6. Dry the crystals formed, between two filter papers as shown in *Figure 2.11*.

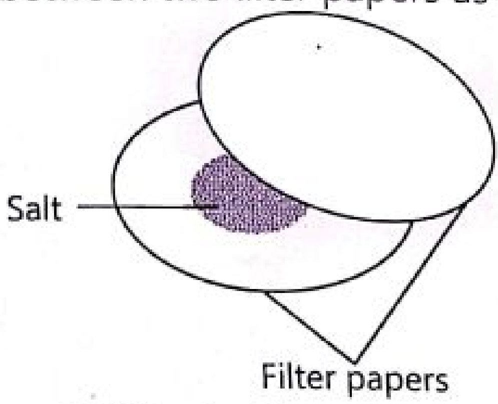


Figure 2.11: *Drying the crystals formed*

Discussion Questions

1. Which gas is formed, and how can it be tested for?
2. Why is it not advisable to evaporate the solution completely to obtain the salt?
3. Why is a water bath used?

Conclusion

Carbonates react vigorously with acids, releasing carbon dioxide gas. Some of the other salts that can be similarly prepared by the action of a dilute acid on a carbonate are:

- copper(II) sulphate
- copper(II) nitrate
- magnesium sulphate
- zinc sulphate
- calcium chloride
- calcium nitrate
- copper(II) chloride

e) Action of an acid on a metal

In chapter 1, you looked at how acids react with metals to form a salt and hydrogen.

Equation:
acid + metal $\rightarrow$ salt + hydrogen

Activity 2.1(e): Preparing zinc chloride from zinc powder and dilute hydrochloric acid

This activity will be done in groups.

Using your prior knowledge, prepare zinc chloride from zinc powder and dilute hydrochloric acid as shown in *Figure 2.12.*

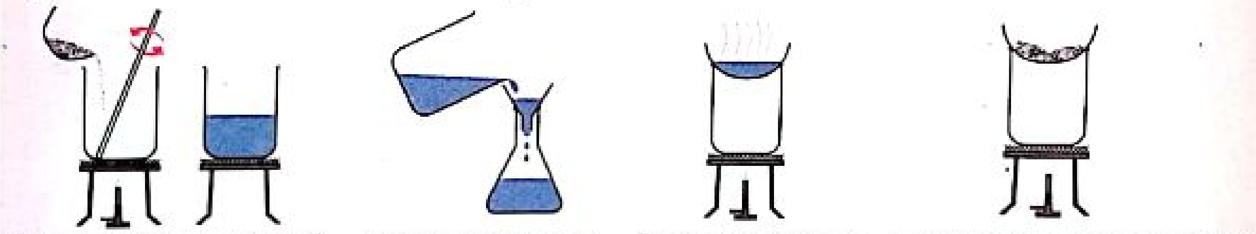


Figure 2.12: Preparation of zinc chloride from zinc and dilute hydrochloric acid.

Discussion Questions

1. Explain;

   a) why zinc powder was used instead of zinc granules.

   b) why it was necessary to filter the mixture before the evaporation was done.

2. Write the word equation for the reaction that took place.

This method is appropriate to prepare salts of metals which are moderately reactive; for example, salts of magnesium, zinc, and iron.

From the activities you have done so far in this chapter, you have found out that dilute acids react with metalic oxides or hydroxides to form salts and water only. They also react with carbonates to form salts, carbon dioxide, and water, and with metals to form salts and hydrogen. In all reactions, a salt is formed, which is concentrated by evaporating some of the water (in the filtrate) and crystallising on cooling the solution.

Insoluble salts

Some salts are insoluble in water. These salts continue to exist as solids, rather than dissolving in a liquid to form a solution. If an insoluble salt forms in water, it is called ***a precipitate***. The method used to prepare insoluble salts is referred to as ***precipitation*** or ***double decomposition***, since it involves the exchange of ions. In this reaction, an acid reacts with a soluble salt in a solution to form a soluble and an insoluble salt. The soluble salt formed remains in the solution and the insoluble salt settles at the bottom. The insoluble salt is then filtered off, washed in water, and dried.

Activity 2.1(f): Preparing lead(II) sulphate from dilute sulphuric acid and lead(II) nitrate

This activity will be done in groups.

What you need:

- dilute sulphuric acid
- beakers
- filter funnel
- lead(II) nitrate solution
- dilute hydrochloric acid
- filter papers
- silver nitrate solution
- measuring cylinder
- a stirrer

What to do:

1. Using a measuring cylinder, measure 50 cm^3 of dilute sulphuric acid and pour it into one beaker, and 50 cm^3 of lead(II) nitrate solution into another beaker.
2. Mix the two solutions and stir.
3. Record your observations in *Table 2.1*.

	TEST	Observations
1	To 50 cm^3 of lead(II) nitrate solution in a beaker, add dilute sulphuric acid.	
2	To 50 cm^3 of silver nitrate solution in a beaker, add dilute hydrochloric acid.	

4. Filter the precipitate formed, wash it several times with distilled water, and dry it using filter papers.
5. Write a word equation for the reaction.
6. Repeat procedures 1, 2, 3 and 4 using dilute hydrochloric acid and silver nitrate solution.

Dilute sulphuric acid reacts with lead(II) nitrate solution to form nitric acid which is soluble and lead(II) sulphate which is insoluble and is formed as a precipitate.

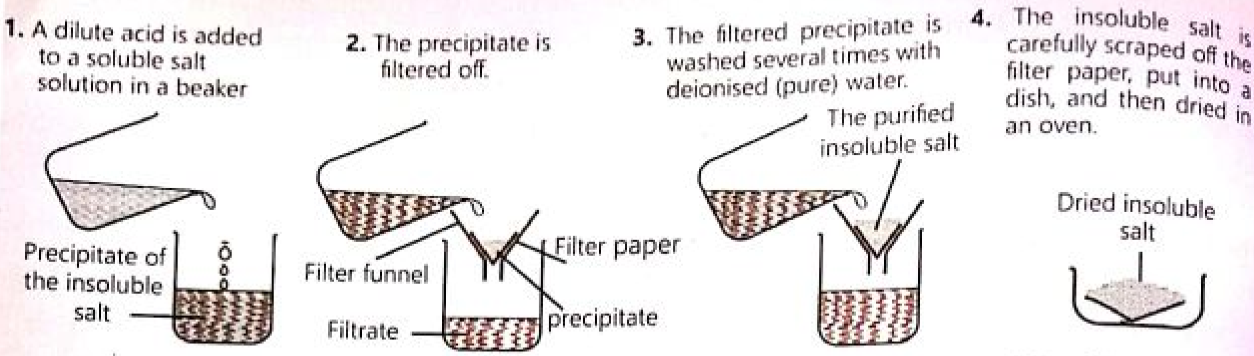


Figure 2.13: Preparation of lead(II) sulphate from dilute sulphuric acid and lead(II) nitrate

DID YOU KNOW? Now you can prepare any insoluble salt; for example, calcium carbonate (using sodium carbonate and calcium chloride solution). This involves reacting two soluble salts to form a soluble salt and an insoluble one.

Assignment 2.2

Discuss other examples of salts that can be prepared by the double decomposition method.

2.2 Importance of Common Salts

Salts play an important role in our lives. They have health, economical, agricultural and traditional values.

They provide minerals, such as sodium, which is essential for the proper functioning of nerves and muscles and is involved in the regulation of fluids in the body. Sodium also plays a role in the body's control of blood pressure and volume. Calcium, from salts, is important in strengthening of our bones and teeth.

Health Check

i) Many people know about the health risks associated with taking sodium chloride in excess, but reducing its intake is easier said than done.

ii) Even if you do not have hypertension or high blood pressure, it is still a good idea to cut down on your salt intake in order to maintain a normal blood pressure.

Activity 2.2(a): Uses of common salts

1. Research and discuss the uses of some common salts.
2. Copy and complete *Table 2.2.*

Table 2.2: Uses of salts

Name of the salt	Importance in our homes	Importance in manufacturing industries
Sodium chloride		
Sodium hydrogen carbonate		
Calcium carbonate		
Potassium nitrate		
Lead chloride		
Potassium carbonate		
Barium sulphate		
Sodium phosphate		
Ammonium chloride		

In our daily lives, everyone uses salts. Although meat, plants, and dairy products are considered key sources for mineral salts, the essential sources of mineral salts are the mines and sea water. Additionally, in certain industries, including food and beverages, pharmaceuticals and agriculture sectors, salts are also used as materials to manufacture other products which are used locally or exported for economic gains.

DID YOU KNOW?

Salts are used in making fireworks. Potassium nitrate is one of the most important components of gunpowder. This is what propels the fireworks into the sky.

A fuse is used to light the gun powder, which ignites to send the fireworks towards the sky. Once the fireworks are in the air, more gunpowder inside them causes them to explode with a bang.

Activity 2.2(b): How salts are used in fireworks

In this activity, you will work in groups.

1. Salts are used in fireworks and the colour of the fireworks depends on the metals present in the salt, as shown in *Figure 2.14*. Name some of the salts used to make fireworks, other than potassium nitrate.
2. Research about the Chemistry behind fireworks and explain your findings to the rest of the class.

Figure 2.14: Fireworks

Chemical formulae of salts

In Senior One, you learnt about chemical symbols of different elements.

Chemical formulae of compounds are derived from symbols of the elements and formulae of the radicals contained in the chemical names of those compounds.

A ***radical*** is a charged atom or a group of charged atoms that can not exist on their own. The charge on a radical corresponds to its valency. ***Valency*** is the combining power of an element or a radical. Valences of elements and radicals are constant.

In Chapter 3, you will learn about how valences of elements and radicals are determined.

Table 2.3: Some of the elements and radicals, with their valences

Radical	Symbol	Valence	Element	Symbol	Valence
Phosphate	PO_4	3	Sodium	Na	1
Nitrate	NO_3	1	Chlorine	Cl	1
Nitrite	NO_2	1	Lead	Pb	2
Ammonium	NH_4	1	Potassium	K	1
Hydrogen carbonate	HCO_3	1	Aluminium	Al	3
Carbonate	CO_3	2	Calcium	Ca	2

Writing chemical formulae

To write the chemical formulae of a compound, first identify the elements which are present in the compound.

Step 1: Write the symbol(s) of the element(s) or radical present in the compound.

Step 2: Indicate the valencies at the top right corners on the symbols of the element(s) or radical. For example, to write the chemical formula of sodium chloride (NaCl).

We shall write Na^1 and Cl^1, since each element present has valency one.

Step 3: Interchange the valences between the elements or element and the radical.

$Na^1 \times Cl^1$

Step 4: Place the valences at the bottom right corners of those symbols.

$Na_1Cl_1 \rightarrow NaCl$, which is the chemical formula of sodium chloride.

Salts are compounds. A compound is a substance formed when two or more elements are chemically combined together. Normally, salt ingredients occur as chemical compounds which are produced by living things and are crystalline in nature.

In *Activity 2.2(c)*, you will write the chemical formulae of some common salts.

Activity 2.2(c): Chemical formulae of salts

In this activity, you will use the internet and Chemistry textbooks to come up with the chemical formulae of some common salts.

Copy and complete *Table 2.4*.

Table 2.4: Elements found in some salts, and their chemical formulae

Salt	Elements found in the salt	Formula
Sodium chloride	Sodium and chlorine	NaCl
Sodium hydrogen carbonate		
Calcium carbonate		
Potassium nitrate		
Lead(II) chloride		
Potassium carbonate		
Barium sulphate		
Sodium phosphate		
Ammonium chloride		

Assignment 2.3

In groups, discuss, using the previous knowledge, how writing word equations of the common salts explains how these salts might have been made.

Solubility

Have you ever wondered why some salts dissolve in water while others do not? Have you ever wondered why table salt is usually added to hot food and not cold food? Have you ever dissolved common salt and sugar, separately, in water? If yes, which one dissolves faster? After carrying out some activities, you will find out that there are several factors which determine the solubility of a substance in water.

Activity 2.2(d): Factors which affect solubility of a substance

In groups, carry out the following activity.

What you need:

- sugar
- table salt
- a stirrer
- beakers
- water
- a heat source
- a tripod stand
- a wire gauze
- a weighing scale
- a measuring cylinder

What to do:

1. Using a measuring cylinder, measure 100 cm^3 of water which is at room temperature and pour it into a beaker.
2. Weigh 50 g of common salt and add it to the beaker in (1) above. Stir it up.
3. Repeat procedures 1 - 2 using 50 g of sugar.
4. Compare the solubilities of common salt and sugar at room temperature.
5. Repeat procedures 1, 2 and 3 using hot water at 60 ^{0}C.
6. Compare the solubilities of the two substances in cold and hot water.
7. From your observations, deduce the factors that affect the solubility of a substance.

From your observations, you have found out that solubility depends on the nature of a substance. Sugar dissolves in water more readily than common salt. It is also noted that when the temperature of water was increased, the solubility of the sugar also increased. More sugar or salt can dissolve if the amount of water increases.
At a given temperature, salt dissolves until the water can no longer dissolve anymore of it. At this point, the solution is said to be ***saturated.*** A saturated solution is a solution which can not dissolve any more solute at a particular temperature.

Solubility: is the amount of a salt in grams that is required to saturate 100 g of water, at a given temperature.

Activity 2.2(e): Finding the solubility of potassium nitrate in water at room temperature

What you need:

- water
- stirring rod
- watch glass
- thermometer
- heat source
- measuring cylinder
- boiling tube
- evaporating dish
- weighing scale
- powdered potassium nitrate

What to do:

1. Using a measuring cylinder, measure 10 cm^3 of water, pour it into a boiling tube and warm at 30°C.
2. Add powdered potassium nitrate into a beaker, little at a time, while stirring with a glass rod until no more can dissolve.
3. Using a weighing scale, measure the weight of an empty, clean and dry evaporating dish or watch glass. Record it as (x) g.
4. Decant the solution carefully into the evaporating dish or watch glass.
5. Weigh the evaporating dish or watch glass with the solution. Record it as (y) g.
6. Evaporate the solution to dryness over a medium flame. Do not heat the solid strongly to decompose it.
7. Allow the evaporating dish or watch glass to cool and then weigh it with the solid, potassium nitrate. Record the weight as (z) g.
8. Repeat procedures 1-7 using sodium chloride, copper(II) sulphate, sodium nitrate, barium sulphate, potassium carbonate and lead(II) chloride instead of potassium nitrate.

Discussion Questions

1. Calculate the amount of potassium nitrate which would saturate 100 g of water.
2. What is the main use of the watch glass in the experiment?
3. What is the effect of changing the temperature of water (the solvent in this case) on the solubility of the salt (the solute)?
4. What are the physical and chemical properties that affect the solubility of a solute?

Conclusion

Solubility is the property of a solid, liquid or gaseous solute that dissolves in a solid, liquid or gaseous solvent.

The extent of solubility of a substance in a specific solvent is measured as the ***saturation concentration***, where adding more solute does not increase the concentration of the solution; instead, it begins to precipitate the excess amount of solute.

Solubility should not be confused with the ability to dissolve a substance, since a solution can also be formed as a result of a chemical reaction. For example, zinc dissolves (with effervescence) in dilute hydrochloric acid as a result of a chemical reaction, releasing hydrogen gas in a displacement reaction. The zinc ions are soluble in the acid.

The aspect of differences in the solubilities of salts at different temperatures is applied in extraction of salts; for example, sodium chloride from Lake Katwe.

Activity 2.2(f): Separating salts with different solubilities

What you need:

- a beaker
- a glass rod
- sodium chloride
- sodium carbonate
- a weighing balance
- measuring cylinder
- a heat source
- water

What to do:

1. Using a weighing balance, measure 50 g sodium chloride and 50 g of sodium carbonate.
2. Dissolve both salts in a glass beaker with 100 cm^3 of water measured using a measuring cylinder and heat the solution until all the salts have dissolved.
3. Cool the solution at room temperature.
4. Filter the crystals that have formed and dissolve them in just enough hot water and allow to cool so as to recrystallise.
5. Repeat steps 3 and 4 dissolving and recrystallising, 2 more times.
6. Dry the crystals between filter papers.

Conclusion

The filtrate contains a large amount of the substance that is more soluble (than the other) at room temperature. On continuing the process of dissolving and recrystallising, the least soluble substance is separated.

When two salts are dissolved in hot water and the solution is cooled, the salt with the higher solubility at high temperatures will crystalise out of the aqueous solution first. The process by which two soluble salts are separated from an aqueous solution using their differences in solubility is called ***fractional crystallisation.*** This is mainly what takes place at Lake Katwe.

DID YOU KNOW?

Save time and money! When buying soap, consider its solubility in water. The more soluble the soap is in water, the less soap you will require to finish the activity. This will save your money!

Project Work:

Locally, salt is obtained from the ash of banana peelings or papyrus. With the help of drawings, describe briefly how a sample of salt can be obtained from the ash.

Activity 2.2(g): What is the pH of salts?

What you need:

- sodium chloride
- ammonium nitrate
- potassium sulphate
- sodium carbonate
- ammonium carbonate
- potassium carbonate
- potassium chloride crystals
- sodium phosphate
- aluminium chloride
- a dropper
- a spatula
- water
- a test tube
- a universal indicator
- measuring cylinder
- the pH chart
- potassium nitrate
- ammonium chloride

What to do:

1. Using a measuring cylinder, measure 5 cm^3 of water and pour it into a test tube.
2. Add half a spatula endful of potassium chloride crystals and shake to dissolve.
3. Add 2-3 drops of the universal indicator and shake.
4. Using a pH chart, record the pH of the solution in *Table 2.5.*
5. Repeat steps 1-4 using the other salts.

Table 2.5: Determination of the pH of the resultant solution when the salts are dissolved in water

Salt	pH
Sodium chloride	
Ammonium nitrate	
Potassium sulphate	
Sodium carbonate	
Ammonium carbonate	
Potassium carbonate	
Aluminiun chloride	
Sodium phosphate	
Ammonium chloride	
Potassium nitrate	
Potassium chloride	

Salts are made from acids and alkalis. Acids and alkalis are classified as either weak or strong. When a salt is formed from a strong acid and a strong base, the resultant solution is neutral. The number of hydrogen ions is equal to that of the hydroxide ions, in the same solution. When a salt is formed from a weak acid and strong alkali, the resultant solution is alkaline. This is because there are more hydroxide ions (in that solution) than the hydrogen ions.

When a salt is formed from a strong acid and a weak alkali, the resultant solution is acidic. This is because the number of hydrogen ions is greater than that of the hydroxide ions in that solution.

ICT Activity

a) In groups, using the internet, research and watch videos about the production of salts from Lake Katwe, pointing out the factors which affect their production.

b) Make a summary of what you have observed, and use a suitable software to present your findings to the rest of the class.

Sample Activity of Integration

Lake Katwe is a historical source of common salt in Uganda. However, salt extraction from the lake is disorganised, as shown in *Figure 2.15*. The ore will also get used up with time, implying that the salt source is non-renewable.

The community now wonders how people will survive without salt, upon depletion of the ore in the lake.

Figure 2.15: Salt mining at Lake Katwe

Task: The area Member of Parliament has invited you for a talk show as a Chemistry student. You are to discuss alternative ways on how common salt can be obtained or made available. Describe the procedure followed to obtain the salt.

Chapter Summary

In this chapter, you have learnt that:

- Salts can be prepared by the reaction between acids and bases.
- Neutralisation is a reaction between an acid and a base to form a salt and water only.
- Salts get their names from the acids they are derived from.
- Some of the physical properties of salts include different colours, solubility in water.
- Some salts are soluble in water, while others are insoluble.
- All carbonates except those of potassium, sodium and ammonium are insoluble in water. It is important to note that anhydrous solid carbonates normally exist in powder form.
- All salts of potassium and sodium are readily soluble in water.

Review Exercise:

1. a) What is a neutralisation reaction?
   b) Give the products formed when the following compounds react, and write word equations for the reactions:
   i) Dilute nitric acid and calcium carbonate
   ii) Dilute sulphuric acid and potassium hydroxide
   iii) Dilute hydrochloric acid and copper(II) oxide
   iv) Dilute hydrochloric acid and magnesium
2. Describe briefly how copper(II) sulphate crystals can be prepared in the laboratory using copper(II) oxide.
3. Describe how you can separate a mixture of barium sulphate and sodium chloride using the common school laboratory apparatus.
4. a) Discuss the importance of common salts in our daily lives.
   b) Common salt from Lake Katwe is obtained by fractional crystallisation. Explain the term;
   i) fractional crystallisation ii) crystal
   c) Identify one crystalline substance at home.
5. Name the salts obtained from the following reactions;
   a) zinc metal + dilute hydrochloric acid
   b) sodium carbonate + dilute sulphuric acid
   c) copper(II) oxide + dilute hydrochloric acid
6. Indicate whether each of the following statements is true or false.
   a) A saturated solution contains less solute than it can dissolve.
   b) The process by which a salt is obtained from a solution by first cooling the solution and then warming it is called recrystallisation.
   c) It is possible to prepare a salt from two salts.
   d) Two soluble salts can be used to prepare an insoluble salt.
   e) When carbonates are heated, the gas evolved is carbon monoxide.

CHAPTER 3 The Periodic Table

period \ group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

6	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
7	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Keywords

- atomic number
- electron
- group
- mass number
- neutron
- period
- proton

By the end of this chapter, you should be able to:

a) understand that elements can be grouped into metals and non-metals, and relate the physical properties of metals and non-metals to their uses.

b) know that the periodic table is a classification of elements according to their atomic numbers.

c) relate the arrangement of the first 20 elements in groups and their positions in the periodic table.

d) understand the relationship between the position of elements in groups and the charge on the ions that they form.

Competency: You should be able to investigate the diversity of elements in the Periodic Table.

Introduction

Table 3.1: A calendar of March, 2020

	Sun	Mon	Tue	Wed	Thu	Fri	Sat
Week 1	1	2	3	4	5	6	7
Week 2	8	9	10	11	12	13	14
Week 3	15	16	17	18	19	20	21
Week 4	22	23	24	25	26	27	28
Week 5	29	30	31				

Study the calendar in *Table 3.1*. What do you notice?

- The 31 days are not arranged in one straight line.
- The days are arranged in weeks. That after every 7 days, a new week begins. The cycle repeats until all the 31 days have elapsed and that is where the next month will start from.
- As you move from left to right on a calendar, you move from one day to another; for example, from Sunday to Monday, Monday to Tuesday, and so on.
- As you move downwards, you move from one week to another; for example from week 1 to week 2, and there is always a difference of 7 days.

In the similar way, days are arranged in a calendar, elements are also arranged in a table where each element occupies a unique position. The table that shows how elements are arranged is called the ***Periodic table.*** In this chapter, you will investigate the repetitive patterns formed by elements and their diversity in the Periodic Table.

3.1 What is a Periodic Table?

In Senior 1, you learnt about what an element is. You also learnt that elements are represented by chemical symbols and that each chemical symbol is unique to an element. Elements are indicated in the Periodic Table using their symbols. In *Activity 3.1(a)*, you will find out what a Periodic Table is.

Activity 3.1(a): What is a Periodic Table?

In this activity, you will work in groups.
What to do:

1. Observe the Periodic Table chart in your class or Chemistry laboratory.
2. Observe the numbers associated with each symbol of the element and take note of the pattern followed by them in columns and rows.
3. Discuss the arrangement of the elements.
4. Note the number of columns and rows.
5. Present your responses to the rest of the class.

The arrangement of elements, the numbers associated with the chemical symbols, and the trends they follow, as discussed in *Activity 3.1(a)*, will help you to explain further what the Periodic Table is.

In the Periodic table, the vertical columns of elements are called ***groups***, and are represented by capital roman numbers, for example, group I, group II, group III, and group IV up to group VIII. The horizontal rows of elements are called ***periods***, and are represented by numbers 1, 2, 3, up to 7.

Activity 3.1(b): Finding the meaning of terms used in the Periodic Table

What you need:

- Internet
- Chemistry textbooks

What to do:

In groups, research and write a report explaining:

1. The meanings of atomic number (proton number) and atomic mass (mass number).
2. The relationship between protons and electrons in a neutral atom.
3. The charges on protons, electrons, and neutrons.
4. Present your responses to the rest of the class.

Conclusion

Any atom of an element contains particles called protons, electrons, and neutrons.

Protons are positively charged, neutrons carry no charge, and electrons are negatively charged. The numbers of these particles are unique to a specific element.

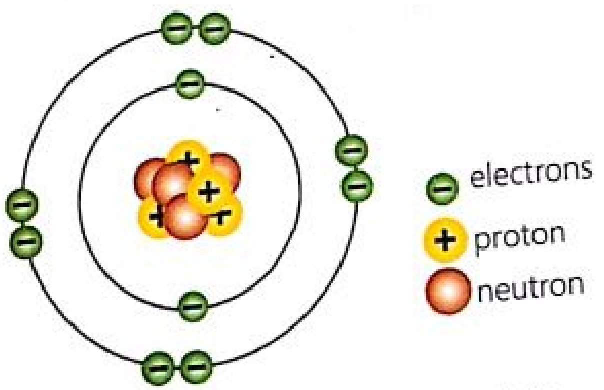


Figure 3.1: The structure of an atom

In a neutral atom, the number of protons is equal to the number of electrons. Therefore, the atomic number is also equal to the number of electrons. The ***atomic number*** is also known as the ***proton number***.

Classification of elements

From *Activity 3.1(a)*, you took note of the numbers on different chemical symbols. In *Activity 3.1(b)*, you found out the meanings of some of the terms used in the Periodic table. In *Activity 3.1(c)*, you will identify atomic numbers and atomic masses of elements and relate them to positions of the elements in the Periodic Table.

Activity 3.1(c): Relating the atomic number and the position of an element in the Periodic Table

In groups, carry out the following activity.

What you need:

- Internet
- Chemistry textbooks

What to do:

1. From the pattern you took note of in *Activity 3.1(a)*, state the general pattern of atomic numbers and atomic masses in columns and rows.
2. Note the atomic numbers of three elements that follow each other and make a conclusion on the arrangement of elements in the Periodic Table.

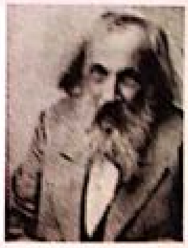
Conclusion

The Periodic Table is, therefore, defined as the tabular classification of chemical elements are arranged in order of increasing atomic number (proton number), such that elements with similar properties are in the same group (column). An element with one proton occupies the first position, implying that for instance, the 20th element has 20 protons.

DID YOU KNOW?

Science fact:

Dmitri Mendeleev, a Siberian-born Russian Chemist, was the first person to make a Periodic Table much like the one that we use today. By 1869, only 63 elements had been discovered and by 2019, they were 118 elements.



Assignment 3.1

Using the internet or any other resource materials, research and discuss the work of John Newlands, Dmitri Mendeleev, and Sir William Ramsay in classifying elements in groups.

a) Write a report about the contribution of Dmitri Mendeleev and other scientists to the development of the Periodic Table.

b) Present your findings to the rest of the class.

The original Periodic Table was modified and corrected several times to accommodate new elements that were discovered later.

Mendeleev arranged different elements using their atomic masses and tried to group the elements with similar chemical properties. He found out that elements recurred at regular intervals. Using his arrangement, Mendeleev left gaps in his Periodic Table as he predicted that new elements would be discovered which would fill those gaps. From the activities done so far in this chapter, you must have discovered that the Periodic Table has, mainly, 7 periods and 8 groups.

3.2 Arrangement of Electrons In Relation to the Position of an Element in the Periodic Table

You now know that atoms contain negatively charged particles called electrons. The arrangement of electrons in atoms will help you to understand further the positions occupied by different elements in the Periodic Table.

Activity 3.2(a): Investigating how electrons are arranged in an atom

What you need:

- Internet
- Chemistry textbooks

What to do:

In groups, carry out research on:

1. How electrons are arranged in atoms. The arrangement of electrons in energy levels around the atomic nucleus is known as ***electronic configuration***.
2. Write the electronic configurations of the first 20 elements.
3. Present your findings to the rest of the class.

Conclusion

For your level, you will restrict yourself to the electronic arrangements of the first twenty elements in the Periodic Table.

A simple model is used to show the arrangement of electrons. This is necessary if you are to visualise how the arrangements actually are.

Circles are used to represent energy levels around the nucleus and small balls to represent electrons. Each energy level can only hold a specific number of electrons. Electrons exist singly or in pairs in the energy levels. Energy levels are numbered 1, 2, 3, on words, starting from the ones which are nearest to the nucleus, as shown in *Figure 3.2*.

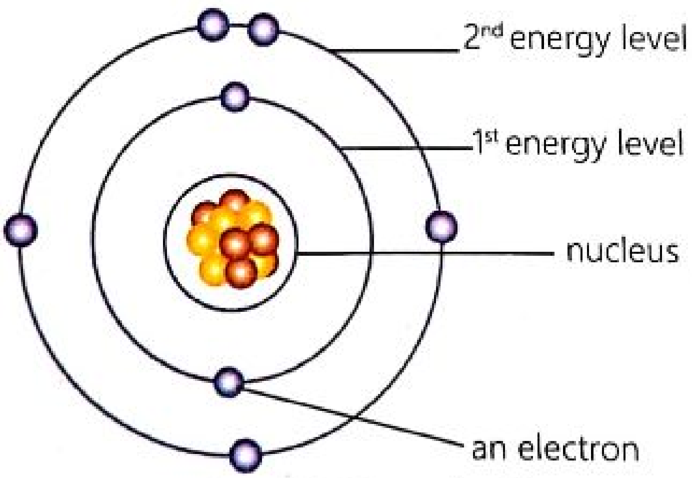


Figure 3.2: Energy levels

The first energy level is filled with a maximum of 2 electrons while the 2nd, 3rd and 4th energy levels are filled with a maximum of 8 electrons each.

From *Activity 3.2(a)*, you must have found out that electrons occupy the energy levels. From the electronic configurations you wrote down, you can tell the number of electrons in the outermost energy level of each atom and the number of energy levels.

In the following activity, you will relate the electronic arrangement to the position of an element in the Periodic Table.

Activity 3.2(b): Relating the electronic arrangement to the position of the element in the Periodic Table

In this activity, you will work in groups.

What you need:

- Internet
- library materials

What to do:

1. Look at the electronic configurations of group I elements. What do you notice?
2. Note the electronic configurations of atoms of elements in any other group. What do you observe?
3. Use the similarity noticed to suggest a general conclusion about the group to which an element belongs.
4. Look at the electronic configurations of atoms of the elements in period 3. Do you notice any similarity? If yes, what is it?
5. Observe the electronic configurations of atoms of the elements in period 4 and state the difference between them and those of the elements in period 3.
6. Observe any other two consecutive periods and draw a general conclusion about the period to which an element belongs.
7. Present your findings to the rest of the class.

From *Activity 3.2(b)*, you discover that as you move from one atom of the element to another, down a group, a complete energy level of electrons is added. You also discover that the number of energy levels corresponds to the period in which an element is found. Also, the number of electrons in the outermost energy level of an atom of an element corresponds to the group to which that element belongs.

When elements are grouped together, basing on ***the number of electrons*** in the outer most energy level of their atoms, they are said to belong to the same group. Group VIII elements are generally referred to as group O elements or noble gases. They are gases and are very unreactive.

Helium is placed in group O, even though its atom has two and not eight electrons in the outer most energy level. This is because its chemical properties resemble those of the elements in this group, whose energy levels are completely filled up.

Elements whose atoms have electrons occuppying the ***same number of energy levels*** belong to the same period.

DID YOU KNOW?

All substances which we eat or use are made up of at least one element of the Periodic Table.

You will now find out how the arrangement of the elements corresponds to their properties. In *Activity 3.2(c)*, you will find out the physical properties of elements and how similar they are, amongst elements of the same group.

Activity 3.2(c): Finding out the physical properties of elements in relation to their positions in the Periodic Table

In this activity, you will work in groups.

What you need:

- Internet
- library materials

What to do:

In groups,

1. Carry out research and find out the physical properties of the first 20 elements. Consider the melting points, boiling points, appearances, colour, smell, taste, states, densities, solubilities, electrical and heat conductivities of those elements.
2. Compare the properties of elements in one group, for instance, sodium and potassium, magnesium and calcium, fluorine and chlorine, and helium and neon.
3. Copy and complete in *Table 3.2*.

Table 3.2: Physical properties of the first twenty elements

Group of the elements	Elements	Physical properties
1. Group I / alkali earth metals		
2. Group II / alkali earth metals		
3. Group III		
4. Group IV		
5. Group V		
6. Group VI		
7. Group VII / halogens		
8. Group VIII or noble / rare / inert gases		

Atoms of elements in the same period have the same number of energy levels. Since they have different numbers of electrons in the outermost energy levels, elements in the same period have different chemical properties.

Atoms of elements in the same group have the same number of electrons in their outermost energy levels. They have varying numbers of the energy levels successively increasing down the group. Since they have the same number of electrons in the outermost energy levels of their atoms, elements in the same group have similar chemical properties.

3.3 Relationship Between the Position of an Element in Groups and the Charge on Ions they Form

In an atom, the number of protons is equal to that of the electrons. This means that an atom is ***electrically neutral.*** During chemical reactions, atoms of elements can lose, gain or share electrons in order to become stable. If an atom loses or gains an electron or electrons, it forms an ***ion.*** Ions are usually very stable and this is because their outermost energy levels are completely filled with electrons just like those of the noble gases.

Elements with more than four electrons in the outermost energy level find it easier to gain electrons, while those with less than four electrons find it easier to lose electrons. In *Activity 3.3(a)*, you will find out more about the formation of ions.

Activity 3.3(a): Finding out how positively charged ions (cations) are formed

You will do this activity in groups.

What you need

- Internet
- library text book

What to do

1. Write down the electronic configuration of magnesium.
2. Determine whether magnesium needs to gain or lose electrons in order to attain the electronic configuration of a noble gas. How many electrons will it lose or gain?
3. Find out the noble gas element whose electronic configuration is nearest to the one of magnesium.
4. Write the electronic configuration of the magnesium ion.
5. Determine the number of electrons and protons in the magnesium ion.
6. Determine the overall charge on the magnesium ion.
7. Find out the formula of the sodium ion. Explain why it carries that charge?
8. Comment on the electronic configurations of sodium and magnesium ions and the charges on them.
9. Present your findings to the rest of the class.

From *Activity 3.3(a)*, you realise that sodium and magnesium lose their outermost electrons to form positively charged ions. The sodium atom has 11 protons and 11 electrons. It is electrically neutral. When a sodium atom loses an electron from its outermost energy level, it remains with 11 protons and 10 electrons. The number of protons become greater than that of the electrons. Therefore, the sodium ion carries a positive charge. This is illustrated in *Figure 3.3*.

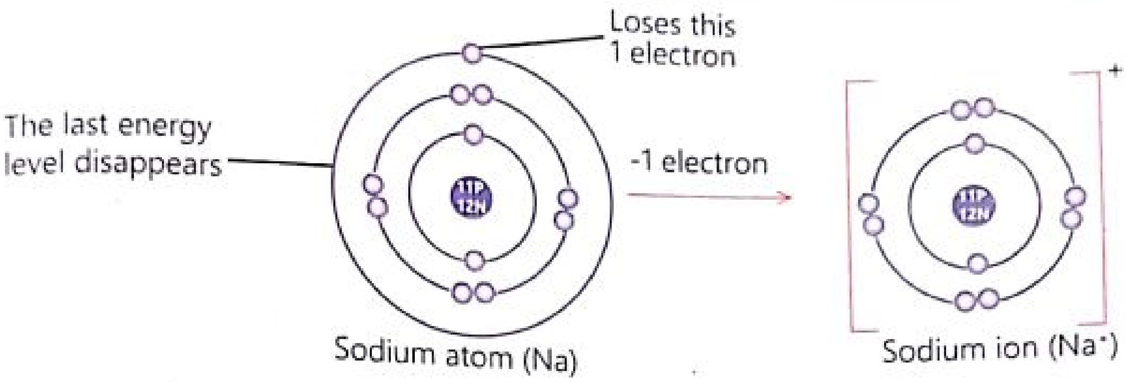


Figure 3.3: Forming the sodium cation

An ***ion*** is a charged particle which is formed from an atom or a group of chemically combined atoms, after gaining or losing one or more electrons. When atoms lose electrons from their outermost energy levels, they become stable. This is because they acquire a stable noble gas electronic configuration. Examples are atoms of group I, group II, and group III elements. When they lose one, two and three electrons respectively, they form positively charged ions called ***cations***, with the stable noble gass structures because the number of protons is greater than that of the electrons. The difference between the number of protons and electrons gives the charge of that ion. In *Activity 3.3(b)*, you will know how negatively charged ions (anions) are formed.

Activity 3.3(b): Finding out how anions are formed

You will do this activity in groups.

What to do:

1. Write down the electronic configuration of the sulphur atom.
2. Determine whether sulphur needs to gain or lose electrons in order to attain the noble gas structure. How many electrons will it lose or gain.
3. Which noble gas structure does sulphur attain when it becomes an ion?
4. Write the electronic configuration of the sulphur ion.
5. Determine the number of electrons and protons in the sulphur ion.
6. Determine the overall charge on the sulphur ion.
7. Find out the charge on the chloride ion.
8. Comment on the electronic configurations of the chloride and sulphur ions.
9. Present your findings to the rest of the class.

From *Activity 3.3(b)*, you realise that the chlorine and sulphur atoms gain electrons to form negatively charged ions with a stable noble gas structure.

The chlorine atom has 17 protons and 17 electrons. It is electrically neutral. When a chlorine atom gains an electron in its outermost energy level, it remains with 17 protons but 18 electrons. The number of protons is less than that of the electrons. Therefore, the chloride ion carries a negative charge, as shown in *Figure 3.4*.

Elements which form negatively charged ions are non-metallic in nature. Elements that become stable by gaining electrons in their outermost energy levels to acquire a stable noble gas electronic configuration, are those of group V, VI and VII. They form negatively charged ions called ***anions.***

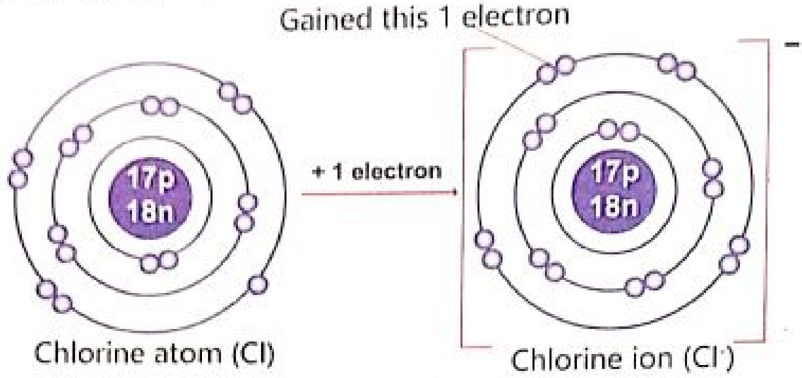


Figure 3.4: Forming the chloride anion

Having learnt about how atoms attain charges, you will now relate the group that an element belongs to in the Periodic Table to the charge on its ion. You found out that the sodium ion has a charge of $^+1$; and so do all other elements in group I. This is because all elements in group I lose the one electron in their outermost energy level so as to attain the stable noble gas structure.

Group II elements have a charge of $^+2$ because they lose the two electrons in their outermost energy levels. Group VI and VII elements form ions with charges of $^-2$ and $^-1$ respectively because they gain 2 electrons and 1 electron respectively.

Fun joke:

You must be made of Fluorine, Iodine and Neon, because you are so FINe (fine).

Most of the elements exist in the earth's crust. However, some of these elements are more abundant than the others.

Activity 3.3(c): Finding out the abundance of elements in the earth's crust

You will do this activity in groups.

What to do:

1. Study *Table 3.3* and answer the questions that follow.

Table 3.3: The percentage composition of elements in the earth's crust

No.	Element	Percentage abundance (%)
1.	Oxygen	46.1
2.	Silicon	28.1
3.	Aluminium	8.23
4.	Iron	5.63

5.	Calcium	4.15
6.	Magnesium	2.33
7.	Sodium	2.36
8.	Potassium	2.09
9.	Lithium	0.002
10.	Gold	0.0000004
11.	Lead	0.0014
12.	Silver	0.0000075

2. Write the proton number, chemical symbol, atomic mass, and the electronic configuration of the first 10 elements in *Table 3.3.*
3. According to *Table 3.3*, what is the most and least abundant element in the earth's crust.
4. Construct a pie chart to represent the data in *Table 3.3.*
5. Present your findings to the rest of the class.

3.4 Classification of Elements into Metals and Non-metals

Having learnt about a variety of elements in the Periodic Table, you can now identify the common elements that you come across in everyday life. These elements are broadly classified as metals and non-metals. In *Activity 3.4(a)*, you will group elements as either metals, or non-metals, and discuss their physical properties and uses.

Activity 3.4(a): Grouping elements into metals and non-metals

You will do this activity in groups.

What you need:

- environment materials
- library textbook

What to do:

1. Identify elements within your environment and classify them as metals and non-metals.
2. Discuss the properties of metals and non-metals.
3. Discuss how the properties relate to the uses of the elements.
4. Present your findings to the rest of the class.

Conclusion

Where do we find metals and non-metals in the Periodic Table? In *Activity 3.4(b)*, you will make conclusions about the arrangement of metals and non-metals in the Periodic Table.

Table 3.4: Physical properties of metals and non metals

Property	Metals	Non-Metals
Physical State	Metals are solid at room temperature except mercury.	Non metals, generally, exist as solids and gases, except bromine.
Melting and boiling points	Metals generally have high melting and boiling points.	Non-metals have low melting and boiling points, except diamond and graphite.
Density	Have generally high densities.	Have generally low densities.
Malleability and Ductility	Are malleable and ductile.	Are neither malleable nor ductile.
Electrical and thermal conductivity	Are good conductors of heat and electricity.	Are generally poor conductors of heat and electricity except graphite
Luster	Possess shining luster.	Do not have luster, except iodine.
Deep and ringing sound	Give deep and ringing sound when struck.	Do not give deep and ringing sound.
Hardness	Are generally hard except sodium and potassium.	Solid non-metals are generally soft except diamond.

Activity 3.4(b): Arrangement of metals and non-metals in the Periodic Table

You will do this activity in groups.

What to do:

1. Identify the groups to which the elements you identified in *Activity 3.2(a)* belong. (Both metals and non-metals)
2. Do you notice any pattern in the arrangement of the elements? If yes, what is it?
3. What conclusions can you draw from the arrangement?

As you move from left to right in the Periodic Table, metallic character decreases as non-metallic character increases. Elements from group I to group III are metals, while those from group IV to group VIII are non-metals. The metallic character increases as you move down the group in the first three groups.

Assignment 3.2

Use the Periodic Table to find the names of:

a) Three metals commonly used in our daily life.

b) Two non-metals found in the substance that humans breathe in.

Relating the physical properties of metals and non-metals to their uses

Table 3.5: Uses of metals and non-metals in relation to their physical properties

Elements	Property	Use
Metals	Good conductors of heat	Used to make pans for cooking.
	Hard and have a shiny appearance	Used to make jewellery.
	Good conductors of electricity, malleable and ductile	Used to make electrical cables.
	Strong	Used to build scaffolding and bridges.
	Sonorous	Used in making bells.
Non-metals	Poor conductors of heat	Used to make pan handles.
	Do not conduct electricity	Used as insulating material around wire cables.

ICT Activity

In groups, use the internet or any other resource materials to search and watch animations about the atom and how electrons are arranged within an atom. Present your findings to the rest of the class.

Sample Activity of Integration

Chemistry is one of the compulsory subjects for learners in secondary schools in Uganda. You are a S.2 learner and the school is organising a Parents' meeting for the S.2 class with their parents. The headteacher wants the parents and their children to appreciate the importance of learning Chemistry in our daily lives. *Figure 3.5* shows the different materials made from different elements.

Figure 3.5: Materials made from different elements

Task

As a S.2 student , the head of the Chemistry department has asked you to prepare a speech to these parents and the learners addressing the following:

i) The value of learning about the Periodic Table, with its components and arrangement.

ii) The importance of the knowledge of physical properties of the metals and non-metals in our daily lives.

Write a speech you will deliver in the meeting.

Chapter Summary

In this chapter, you have learnt that;

- Mendeleev organised the elements in order of increasing atomic mass and in groups of similar chemical behaviour. He also left spaces / gaps for not yet dicovered elements and used the patterns of his table to make predictions of the properties of these undiscovered elements.
- The elements in the modern Periodic Table are arranged in order of increasing atomic number.
- The columns in the Periodic Table are called groups or families, because they contain elements with similar chemical properties.
- All the members of a particular group have the same number of valence electrons and similar chemical properties.
- The number of valence electrons is important, because it determines the chemical properties of elements in a group.
- The rows in the Periodic Table are called periods. Elements in the same period have the same number of energy levels but different number of electrons.
- Elements are categorised as metals and non-metals.
- Metals are found on the left side while non-metals on the right side of the Periodic Table.
- Metallic character generally increases from top to bottom, down a group, among metals in the periodic table.
- Metals form positively charged ions while non-metals form negatively charged ions.

Review Exercise: no activity

1. a) Atoms of the elements D, A, E, and J have the atomic numbers 9, 12, 8 and 6 respectively.
   i) Write the electronic configuration of each element.
   ii) Which of these elements are;
      a) metals? b) non-metals?

   b) Elements A, D, and E form ions. Write the formula of each ion formed.
2. Iron is a commonly used metal. Give the properties of iron which make it suitable for use as;
   i) a roofing material.
   ii) a school gong / bell.
   iii) a clothes hanger.

CHAPTER 4

Carbon in the Environment

Keywords

- allotropes
- carbon
- environment
- greenhouse gases
- hard water
- renewable
- fuels

By the end of this chapter, you should be able to:

a) understand how and why carbon compounds are used as fuels.

b) know and appreciate the difference between renewable and non-renewable fuels and understand that non-renewable fuels are not sustainable.

c) know and appreciate the impact of burning carbon-based fuels on the environment.

d) understand the process of making charcoal.

e) know and appreciate the physical properties and uses of carbon dioxide.

f) understand how the increase in carbon dioxide in the air can cause the atmosphere and the oceans to get warmer.

g) understand what greenhouse gases are, where they come from and how they are affecting climate.

h) understand the origin of hard water in limestone areas and investigate how it can be softened.

i) understand how the properties and uses of allotropes of carbon relate to their structures.

Competency: You should be able to investigate the diversity of carbon compounds in the environment.

Introduction

Figure 4.1: Some of the substances that contain carbon

In Chapter 2, you looked at carbon as one of the elements in the Periodic Table. Is carbon a metal or a non-metal? What are some of the properties of carbon? Where in the environment can we find carbon?

Carbon is the fourth most likely abundant element in the universe and is the building block of life on earth. *Figure 4.1*, shows some of the substances that contain carbon. The rate of combustion of fossil fuels has, to some extent, increased the amount of carbon dioxide in the atmosphere. Other processes such as respiration and many others have also added some carbon dioxide into the atmosphere. All these processes release large amounts of carbon dioxide into the atmosphere. On average, the proportion of carbon dioxide in the air remains more or less constant that is about 0.03% by volume. This is an indication that there are ways through which carbon dioxide is removed from the atmosphere. All these processes and ways through which carbon dioxide is added to and removed from the atmosphere are known as the ***carbon cycle.***

The carbon cycle

Activity 4.0: Drawing a carbon cycle

In this activity, you will work in groups.

What you need:

- manila paper
- coloured pencils or markers
- a ruler
- sellotape or masking tape

What to do:

1. Illustrate the processes of the carbon cycle on manila paper.
2. Use arrows to indicate the interchange between carbon forms.
3. Use different colours and drawings to make your work appealing.
4. Compare your work with that from other groups.
5. Add missing parts and display your work using sellotape on the chalkboard.
6. What are the components of a carbon cycle?
7. What is the role of a carbon cycle?
8. Present your responses to the rest of the class.

Conclusion

The carbon cycle shows how carbon is recycled in nature. This is an important cycle that transforms carbon into compounds that are of important use to the lives of living organisms. It is through the carbon cycle that the components of the bodies of living organisms are made or removed.

DID YOU KNOW?

Did you ever wonder what your body is made of?
Sure, you have skin, bones and hair but do you know that all these are made of mainly carbon atoms!

4.1 Using Carbon Compounds as Fuels

Carbon compounds exist in fuels which are used by man for various purposes. In *Activity 4.1*, you will identify the uses of carbon based fuels.

Activity 4.1: Identifying uses of carbon-based fuels

In this activity, you will work in groups.

You are provided with different items which use carbon-based fuels in *Figures 4.2, 4.3, 4.4 and 4.5.*

Figure 4.2: *A lamp*

Figure 4.3: *A generator*

Figure 4.4: *Firewood*

Figure 4.5: *A motor car*

Study *Figures 4.2, 4.3, 4.4 and 4.5.* Identify:

1. The substances used for lighting and / or heating in each figure.
2. Which of the substances used are carbon-based fuels?
3. Which of the fuels burns faster?
4. Which of the fuels produce(s) soot?

Conclusion

Most of the substances that you use at home to produce light and heat contain carbon. These include; firewood, kerosene / paraffin, and natural gas / biogas etc.

These materials are used as fuels because they give a lot of heat or light when burnt in air. Fuels may be made from plants or animals. A fuel is any substance that we burn to provide heat or power.

Carbon-based fuels contain carbon. We convert the chemical energy in the fuel into other forms of energy. We burn most fuels to obtain their energy in the form of heat.

Assignment 4.1

Write a word equation for the burning of carbon-based fuels in:

a) limited oxygen supply (b) excess oxygen supply

4.2 Renewable and Non-renewable Fuels

Non-renewable fuels are formed naturally and are available in limited amounts. These can not be readily replaced due to the long time it takes for them to be formed. Renewable fuels are easily formed over a relatively shorter period of time.

Activity 4.2(a): Identifying renewable and non-renewable fuels

What you need:

- Internet or a library

What to do:

1. Discuss the examples of renewable and non-renewable fuels and explain where they are used.
2. Explain why it is better to use renewable fuels other than non-renewable fuels.
3. Present your findings through a discussion to the rest of the class.

Types of fuel	Advantage	Disadvantage
Renewable fuels	• Maintenance requirements are lower • Can be used over and over again. • Saves the environment by emitting little pollutants into air. • Reliable sources of energy due to availability.	• There is a high need for storage and storage technologies are expensive. • Unpredictable weather events that disrupt the continuity of these resources.
Non-renewable fuels	• Cheap in converting from one type of energy to another. • Easy to use whether in house or anywhere else. • Very cost effective (cheap). • Create jobs to many people in extracting, transporting, and refining. • Very easy to store.	• Take millions of years to form. • Can never be replaced once they have been used. • Release carbon dioxide into the atmosphere when burnt which causes global warming. • Release sulphur dioxide when burned which causes breathing problems. • Time consuming for example mining, inserting into pipes, and transportation.

My environment, my responsibility

We need to use non-renewable fuels sustainably and use more renewable fuels due to their sustainability and environmental friendliness.

Activity 4.2(b): Using fuel in your locality sustainably

What you need:

- Internet or a library

What to do: In groups:

1. Research about the idea of sustainability and write a report on how the use of fuel in your locality can be made more sustainable.
2. Present your work to the teacher.

In order to conserve deposits of fossil fuels, alternative energy sources (fuels) like biogas, biodiesel and alcohols are now being used. Biogas is made from the anaerobic digestion of waste materials, for example, animal excreta and dead plants.

For your understanding

Alcohol as a fuel

Sugarcane plants are one of the most efficient convertors of sunlight into chemical energy. Sugarcane juice can be used to prepare molasses, which are fermented to give alcohol (ethanol). Some countries now use alcohol as an additive in petrol, since it is a cleaner fuel that gives rise to only carbon dioxide and water on burning in sufficient air (oxygen).

4.3 Impact of Burning Carbon-Based Fuels on the Environment

When cooking at home or school with carbon-based fuels like charcoal, firewood or paraffin, smoke is always produced. Where does this smoke go after being produced? Which gas pollutant is highly abundant in that smoke? Which other activities lead to the production of smoke / fumes? What is the impact of too much smoke / fumes accumulating in the atmosphere?

Using the knowledge acquired from answering the questions above, you will discuss the effect of burning carbon-based fuels in *Activity 4.3*.

Activity 4.3: Effect of burning carbon-based fuels

In this activity, you will work in groups.

What you need:

- a library or Internet

What to do:

1. Discuss how;
   a) carbon-based fuels pollute the environment.
   b) the effect of carbon-based fuels on the environment can be minimised.
2. Present your finding to the rest of the class in a discussion.

In a closed room or a poorly ventilated kitchen, carbon burns in limited supply of oxygen to produce carbon monoxide.

Effects of burning carbon in limited supply of air

- It causes global warming
- Leads to suffocation
- Causes smog
- Leads to brain damage and heart diseases.
- Can lead to death in case of serious pollution.

4.4 Process of Making Charcoal

Figure 4.6: Local method of making charcoal from wood

Figure 4.7: Modernised method of making charcoal from wood

Which materials are used to make charcoal? Do you know how charcoal is made from these materials? Most of the charcoal that is used in homes and schools is got from wood and other biomass types. This means that trees are cut down in order to make charcoal. This is done by burning wood and other biomass in limited supply of air, after which charcoal and other byproducts are produced.

Trees should not be cut down to produce charcoal. Instead, other materials should be used, for example, dry agricultural waste and dry animal waste. Charcoal obtained from materials other than wood is called sustainable charcoal because these materials can be obtained on a daily basis.

My environment, my responsibility

Charcoal is the most common fuel used in homes but it has a significant impact on our environment. This is because it leads to deforestation. Alternative sources of fuel which are sustainable can be used instead.

Wood charcoal

This is formed when wood is heated in a limited supply of air (destructive distillation of wood). It is a black porous solid and a very good absorbent of gases. The use of charcoal as a fuel is cheap and efficient but it is sustainable only if it is made from wood that can be regrown easily.

Charcoal can also be made from waste organic materials and charcoal dust. These are called charcoal briquettes shown in *Figure 4.8*. These burn longer (5-6 hours) and are environmentally friendly, unlike charcoal from wood which burns for a shorter time (1-2 hours).

Figure 4.8: Charcoal briquettes

Activity 4.4: Making charcoal from waste organic materials and charcoal dust

In this activity, you will work in groups.

What you need:

- a library
- Internet

What to do:

1. Research and discuss how charcoal is made from waste organic materials.
2. Visit a local charcoal making place, and ask questions about the processes involved in making charcoal.
3. Present your findings to the rest of the class.

4.5 Physical Properties of Carbon dioxide and its Uses

DID YOU KNOW?

Carbon dioxide was first identified by a Scotish chemist and physician, Joseph Black in the 1750s. He found out that carbon dioxide could be produced by heating chalk which lost mass during the heating process, as shown in the equation.

Calcium carbonate → Calcium oxide and carbon dioxide.

In Senior 1, you looked at carbon dioxide as one of the components of air. How can you identify carbon dioxide in the air? The properties of carbon dioxide can be grouped into physical and chemical. Carbon dioxide is a chemical compound composed of one atom of carbon and two atoms of oxygen.

It is found in very small amounts in the atmosphere and accounts for only 0.03% by volume of the total volume of air in the atmosphere.

In *Activity 4.5(a)*, you will find out some of the physical properties of carbon dioxide. Carbon dioxide exists naturally in the atmosphere but can also be prepared in the laboratory, by reacting dilute hydrochloric acid and limestone.

Activity 4.5(a): Preparing carbon dioxide by reacting limestone with dilute hydrochloric acid

In this activity, you will work in groups.

What you need:

- limestone / calcium carbonate
- a wooden splint
- dilute hydrochloric acid
- a thistle funnel or dropping funnel
- concentrated sulphuric acid
- 2 wash bottles
- a flat bottomed flask
- delivery tubes
- a gas jar
- potassium hydrogen carbonate solution

What to do:

1. Place a sample of limestone / calcium carbonate in a flat-bottomed flask.
2. Using delivery tubes, connect the flat-bottomed flask to a wash bottle containing potassium hydrogen carbonate solution, then to another one containing concentrated sulphuric acid.
3. Arrange the apparatus as shown in *Figure 4.9.*

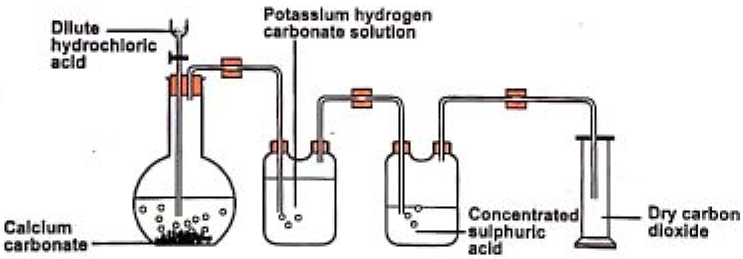


Figure 4.9: *Laboratory preparation of carbon dioxide*

4. Run dilute hydrochloric acid from the thistle funnel into the flat-bottomed flask.
5. Collect the gas using a gas jar.
6. Lower a lighted splint into a gas jar containing the gas.

Discussion Questions

1. Explain why the gas is;
   a) passed through potassium hydrogen carbonate solution.
   b) passed through concentrated sulphuric acid.
   c) collected by downward delivery.
2. Write a word equation leading to the formation of carbon dioxide.
3. Identify the physical properties of carbon dioxide.

Conclusion

The reaction between a metal carbonate and a dilute acid does not go to completion if the metal salt formed is insoluble or sparingly soluble in water.

When dilute sulphuric acid is added to calcium carbonate (marble chips), effervescence occurs for a few seconds and the reaction stops. The insoluble calcium sulphate formed coats the marble chips thereby stopping further contact between the acid and the carbonate, hence, stopping reaction.

Sulphuric acid + Calcium carbonate → Calcium sulphate + Water + Carbon dioxide

The reaction between dilute sulphuric acid and lead(II) carbonate / barium carbonate, dilute hydrochloric acid and lead(II) carbonate also do not go to completion because of the same reason.

Health Check

Exposure to carbon dioxide can produce a lot of health effects. These may include headache, dizziness, restlessness, difficulty in breathing, sweating, tiredness, increased heart rate, elevated blood pressure, and convulsions.

Uses of carbon dioxide

For a number of reasons, carbon dioxide is one of the most important gases on earth. Plants use carbon dioxide to make their own food and since humans and other animals depend on plants for food, carbon dioxide is necessary for survival on earth. *Figure 4.10* shows the uses of carbon dioxide in every day life.

Figure: 4.10: Uses of carbon dioxide

Activity 4.5(b): Uses of carbon dioxide

In this activity, you will work in groups.

What you need:

- a library
- Internet

What to do:

1. Identify the uses of carbon dioxide given in *Figure 4.10*.
2. Research and discuss other uses of carbon dioxide.

DID YOU KNOW?

Solid carbon dioxide (shown in *Figure 4.11*), often called dry ice, does not melt. At atmospheric pressure, it sublimes at -78.5°C. The gas is so cold that moisture in the air condenses on it, creating a dense fog used in stage and movie effects. Solid carbon dioxide is scattered in the clouds to precipitate rain.

***Figure:* 4.11:** *Solid carbon dioxide used to preserve perishable items*

4.6 Effects of Increasing Carbon dioxide Concentration in the Atmosphere and Oceans

When we drive, burn oil and gas for cooking, use electricity generated from coal and burn vegetation to clear land for farming, there is an increase of carbon dioxide in the atmosphere. Carbon dioxide is removed from the atmosphere naturally by plants during ***photosynthesis.***

Plants use sunlight to convert carbon dioxide into carbohydrates. Thus, plants provide food to organisms directly or indirectly. However, humans are destroying large areas of natural vegetation such as the forests. Therefore, the rates of photosynthesis are declining, leaving more carbon dioxide in the atmosphere.

The danger is increased if we burn vegetation since combustion releases more carbon dioxide into the air. In *Activity 4.6*, you will find out the impact of high levels of carbon dioxide in the atmosphere.

Activity 4.6: Effect of increased carbon dioxide on the atmosphere and the oceans

In this activity, you will work in groups.

What you need:

- a library
- Internet

What to do:

1. Research about and report on why the increase in carbon dioxide levels in the atmosphere can cause the atmosphere and the oceans to get warmer.
2. Discuss other human activities that increase the levels of carbon dioxide in the atmosphere.
3. Suggest and explain the possible control measures.
4. Present your findings to the rest of the class.

Conclusion

A number of activities contribute to emission of carbon dioxide into the atmosphere which leads to a variety of dangerous environmental and health effects.

4.7 Causes and Effects of Greenhouse Gases

In Uganda, crops are grown seasonally due to seasonal availability of the rains. This results into periods of food scarcity when supplies are low. However, today scientists have devised means of growing some crops throughout the year under controlled conditions in a ***greenhouse.*** A greenhouse allows some radiation in and traps some. The trapped radiation raises the temperature in a greenhouse to the necessary temperature for plant growth.

This is why greenhouses are used to grow plants in some areas where temperatures may be so low. Having known how greenhouses work, it is the same way in which some gases behave in the atmosphere. Gases like carbon dioxide trap heat in the atmosphere. The effect of these gases is known as the ***greenhouse effect.*** All gases that contribute to the greenhouse effect are referred to as the ***greenhouse gases.***

A greenhouse gas is a gas that absorbs and emits radiant energy within the thermo infra-red range. Greenhouse gases cause the greenhouse effect. The primary greenhouse gases in the earth's atmosphere are water vapour, carbon dioxide, nitrous oxide, ozone and chlorofluorocarbons as shown in *Figure 4.12.*

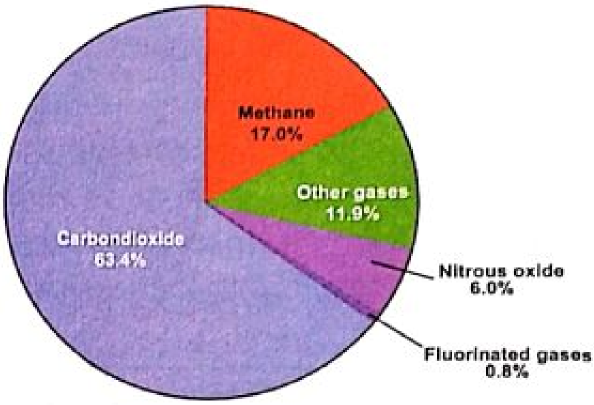


***Figure 4.12:** Greenhouse gases*

In *Activity 4.7*, you will find out the greenhouse gases, their sources, effects and possible methods of controlling them.

Activity 4.7: Greenhouse gases, their sources, effects and possible control methods

In this activity, you will work in groups.

What you need:

- a library
- Internet

What to do:

1. Discuss about greenhouse gases, their sources, effects on the environment and possible methods of control.
2. Copy and complete *Table 4.1.*

Table 4.1: Greenhouse gases, sources, effects and possible control methods

Pollutant	Source	Effect	Possible methods of control
Sulphur dioxide (SO_2)			
Nitrogen oxides (NO, NO_2, N_2O)			
Carbon dioxide			

Carbon monoxide			
Hydrocarbons			
Smoke			
Chlorofluorocarbons			

The effect of a greenhouse gase depends not only on its concentration in the atmosphere but also on its ability to absorb radiation energy from the sun. The absorption-emission process keeps the heat close to the earth's surface. To reduce the rate of greenhouse gas production, scientists have come up with many solutions. These include; planting trees, the use of alternative energy sources, carbon capture and storage facilities, and the use of decarbonised fuel produced by reforming natural gas into a mixture of hydrogen and carbon dioxide.

The enhanced greenhouse effect results into the overall increase in the temperature of the whole world, a situation referred to as ***global warming.*** Global warming changes the climate which affects the distribution of vegetation and causes seasonal changes. This results into longer dry seasons and drought in some areas and flooding in others.

My environment, my responsibility

Every person is capable of reducing the effect of global warming worldwide, but this can only be done if we all collaborate in using environmental friendly fuel.

Assignment 4.2

In groups:

a) Discuss the human activities that cause global warming.

b) Suggest possible measures to minimise on the effects of these activities.

4.8 Hard Water

Figure 4.13: Different sources of water

Have you ever washed clothes or utensils using water collected from different sources such as, rain water, borehole water, spring water or tap water? *Figure 4.13* shows the different sources of water used in every day life.

You realise that washing using water from different sources takes different amounts of soap to form lather.

Water from some sources such as rain water readily forms lather with soap, while water from other sources does not easily form lather. Water which forms lather readily with soap is referred to as ***soft water.*** Water that does not form lather readily with soap is referred to as ***hard water.***

Activity 4.8(a): Sources of soft and hard water

In this activity, you will work in groups.

What to do:

1. Discuss about the different sources of water.
2. Which of the different sources discussed, provide;
   a) hard water b) soft water

Conclusion

There are various sources of water, including underground and surface water.

Causes of hardness of water

From the examples of hard water, you can realise that water hardness is associated with the ground. Most underground and surface rocks contain limestone.

Magnesium and calcium ions in limestone are responsible for the hardness of water. When washing, soap will first react with magnesium and calcium ions before it forms lather. In *Activity 4.8(b)*, you will find out the origin of water hardness.

Activity 4.8(b): How rain water becomes hard

In this activity, you will work in groups.

What to do:

1. Research, discuss and explain how rain water becomes hard as it soaks through limestone.
2. Write word equations of the reactions that occur.
3. Explain how hard water is related to the amount of soap used.
4. Present your findings to the rest of the class.

Conclusion

Rain water is naturally soft, when it falls on porous ground, it penetrates and dissolves magnesium and calcium ions along the way making it hard.

How can hard water be softened?

The method of water softening depends on the type of water hardness. Water hardness can be permanent or temporary.

Temporary hardness of water can be removed by boiling while permanent hardness of water cannot be removed by boiling.

Activity 4.8(c): Removing hardness of water using washing soda

In this activity, you will work in groups.

What you need:

- washing soda (sodium carbonate)
- freshly collected rain water
- spring water
- a spatula
- a measuring cylinder
- test tubes
- a soap solution
- a dropper / stirrer

What to do:

1. Using a measuring cylinder, measure 50 cm^3 of rain water. Add soap solution drop by drop while shaking until a permanent lather is formed.
2. Record the amount of soap used.
3. Repeat procedures 1 and 2 using spring water.
4. Which of the two samples of water used more soap to form a permanent lather?
5. To 50 cm^3 of rain water, add 2 spatula endfuls of sodium carbonate, followed by soap solution drop by drop while stirring, until a permanent lather is formed.
6. Record the amount of soap used.
7. Repeat procedures 5 and 6 using spring water.
8. Compare the amount of soap solution used by each sample before and after the addition of sodium carbonate.
9. Comment on your observations and discuss with the rest of the class.

Hard water can be softened through either physical or chemical means. You have realised that at one point, the addition of washing soda makes lather formation easy while washing.

sodium carbonate + magnesium ions → magnesium carbonate + sodium ions

Washing soda dissolves in water and reacts with calcium ions or magnesium ions which are removed as insoluble carbonates.

Sodium carbonate + calcium ions → calcium carbonate + sodium ions

In *Assignment 4.3*, you will research about other means of water softening. Hard water is also removed when water is passed through permutit which is a complex substance. In this process, calcium or magnesium ions are removed from water through the exchange of ions producing insoluble calcium or magnesium salts of the permutit. The water obtained does not contain magnesium and calcium ions.

Assignment 4.3

In groups:

1. Research about how boiling, distillation, addition of calcium hydroxide solution and ammonia solution removes hardness of water.
2. Present your findings to the rest of the class.

Distillation is a physical method that removes both temporary and permanent hardness of water. Water is vaporised and condensed into distilled water (deionised), which is pure and soft. This method is costly.

For your understanding

Hard water forms solids or 'fur' in kettles and pans. The ***'fur'*** is a layer of a precipitate of calcium or magnesium carbonate. ***'Fur'*** forms when temporary hard water is boiled in kettles and pans. Heat decomposes soluble hydrogen carbonates into insoluble carbonates.

As a result, ***'fur'***, which is a bad conductor of heat, is formed and leads to wastage of fuel. *Figure 4.14*, shows **fur** formed in a kettle and a tap.

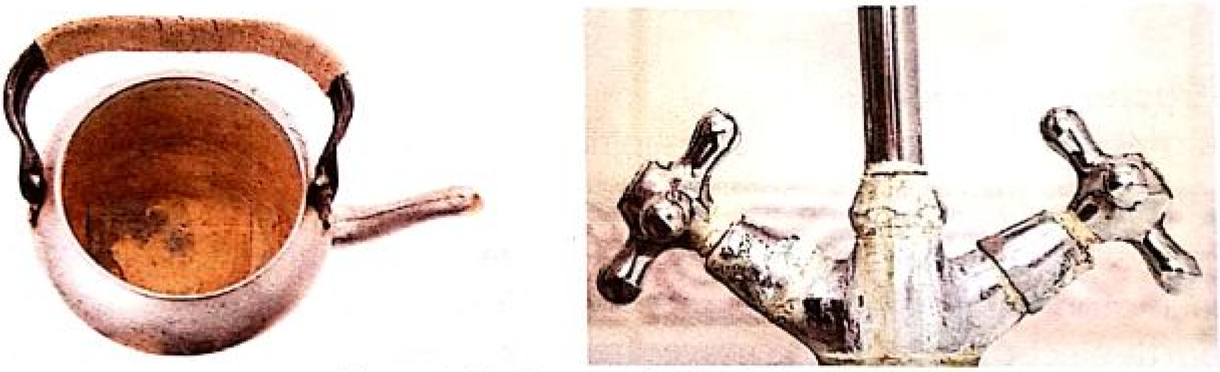


Figure 4.14: *Fur on a kettle and taps*

4.9 How Properties and Uses of Allotropes of Carbon Relate to their Structures

Most elements exist in solid form, by which they can be identified. However, some elements exist in more than one physical form of solid. The situation by which an element can exist in more than one physical form without change of its state is called ***allotropy.*** The different physical forms are allotropes.

Carbon is one of those elements. Carbon is capable of forming many allotropes. Some allotropes of carbon are ***graphite*** and ***diamond.***

The structure of diamond

Diamond is the purest form of carbon, in which each carbon atom is covalently bonded to four other carbon atoms, forming a tetrahedral structure. All its outermost four electrons are used in bonding. The structure is extended throughout the substance in three dimensions. Diamond is a poor conductor of heat and electricity.

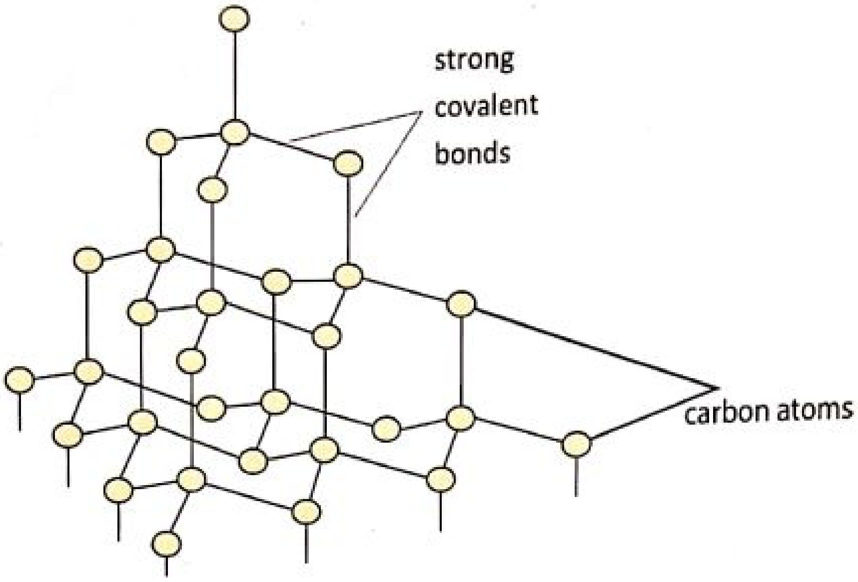


Figure 4.15: Structure of diamond

DID YOU KNOW?

Diamond is the hardest substance ever known on earth and yet it is a non-metal.

Structure of graphite

In graphite, each carbon atom is joined to the other three carbon atoms, forming a hexagonal structure as shown in *Figure 4.16.*

Each carbon atom is at the corner of the hexagon layer and the layers are joined to each other by weak forces of attraction allowing them to slide over each other. The fourth electron in graphite moves freely because it is not attached to any carbon.

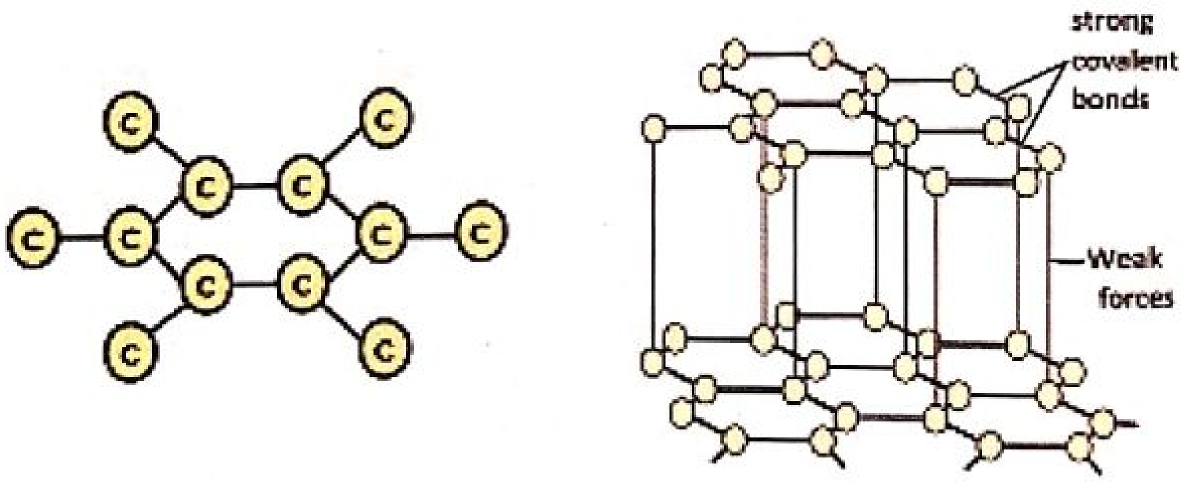


Figure 4.16: Structure of graphite

Fun joke:

How did the carbon boy turn into a diamond?

Activity 4.9: Identifying how the physical properties and uses of the allotropes of carbon relate to their structures

In this activity, you will work in groups.

What to do:

1. Make models of the structure of graphite and diamond using locally available materials.
2. Use their structures to identify the properties of diamond and graphite.
3. Use their properties to determine the uses of carbon structures such as carbon fibres and graphite.
4. Present your work to the rest of the class.

Like diamond, graphite also has a very high melting point. If you burn diamond or graphite in oxygen, then only carbon dioxide gas is formed and nothing is left behind. This shows that graphite and diamond are made up of carbon only and since they are made up of carbon atoms only, their symbol is taken to be **C.**

Graphite has delocalised electrons which are free to move between layers and hence, conducting electricity.

Other allotropes of carbon

Amorphous carbon

- They don't have distinctive shapes like diamond and graphite.
- They are black, opaque and dull.
- They are found to contain few crystals of graphite joined together by a variety of impurities.
- They fairly conduct electricity because of the presence of graphite.

Examples of amorphous forms of carbon: coke and charcoal.

Use:

- electronic displays

Fullerenes

- Has a spherical structure.
- Comprises of 60 carbon atoms arranged as 10 hexagons and 12 pentagons.
- Inter molecular forces are weak and the melting point is low.
- Has a lot of free electrons inside hence conducts electricity.

Use:

- Carrier for gene and drug delivery system.

Note

Drug delivery systems are systems that control the rate at which a drug is released and the location in the body where it is released.

- Used as anti oxidants.

ICT Activity

1. In groups, research on the internet or a library source for the digital models of the structures of diamond and graphite. The difference in their structures and the uses of each allotrope of carbon.
2. Present your findings to the rest of the class.

Sample Activity of Integration

There is a strong global movement on the issue of global warming and greenhouse gas effects. Uganda is part of this movement. The challenges of global warming and effects of greenhouse gases are caused by both urban environment as shown in *Figure 4.17* and local communities as shown in *Figure 4.18*.

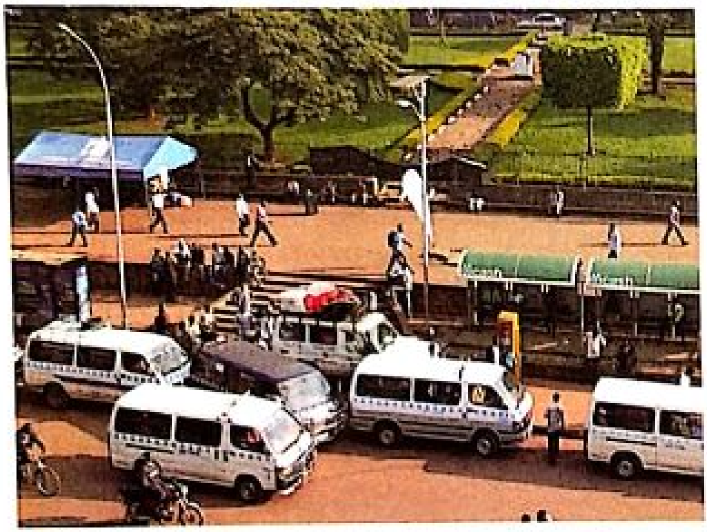


Figure 4.17: *Urban environment*

Figure 4.18: *Charcoal obtained from firewood*

Task:

As a learner of Chemistry, prepare a presentation for your community on a radio talk show about global warming and greenhouse effects.

In the presentation:

1. Give a simple explanation of the meaning of global warming and greenhouse gases.
2. Give an example of how urbanisation is greatly contributing to global warming and greenhouse gases.
3. Give an example of how your local community is also contributing to global warming.
4. Suggest the possible ways of reducing global warming or the effect of greenhouse gases.

Chapter Summary

In this Chapter, you have learnt that:

- On earth, carbon circulates through land, oceans and the atmosphere, creating what is known as the carbon cycle.
- Most carbon compounds are used as fuels because they give a lot of heat and light when burnt in air.
- Renewable fuels are fuels produced from readily available renewable resources while non-renewable fuels are made from resources that are easily depleted.
- Non-renewable fuels are not sustainable and their use should be minimised
- Wood charcoal is formed when wood is heated in a limited supply of air.
- Charcoal obtained from materials other than wood is called sustainable charcoal and it is environmentally friendly.
- Carbon dioxide is a chemical compound composed of one carbon atom and two oxygen atoms. It can be prepared in the laboratory by reacting limestone with dilute hydrochloric acid.
- Hard water is water which contains magnesium and calcium ions and does not easily form lather with soap.
- Hard water is formed when rain water dissolves limestone. Rain water is carbonic acid because it contains dissolved carbon dioxide.
- Boiling removes only temporary hardness of water while distillation and addition of washing soda removes both permanent and temporary hardness.
- Carbon is an allotropic element because it exists in different forms without change of state.
- Different forms are called allotropes of carbon, for example, diamond and graphite.

Review Exercises:

1. a) Why is carbon and its carbon compounds used as fuels?
   b) Describe the impact of burning carbon-based fuels on the environment?
2. a) Describe the differences between renewable and non-renewable resources.
   b) Explain why non-renewable fuels are not sustainable?
   c) Construct a summary table stating two advantages and two disadvantages of each type of renewable energy.
   d) What are the six ways in which an average household could reduce its annual energy consumption?

3. a) Diamond and graphite are allotropes of carbon. Name other allotropes of carbon.
   b) Discuss the differences between the physical properties of diamond and graphite.
4. a) Discuss the uses of charcoal in our daily life, other than being a fuel.
   b) Explain why a charcoal burner should not be used in a poorly ventilated room.
5. a) Discuss how hard water is formed.
   b) When washing clothes using water from a well, more soap is used than when washing with rain water.
      i) What type of water does the well contain?
      ii) Name the ions that are responsible for the type of water given in b(i).
   c) Suggest one physical and one chemical method used to remove ions responsible for the type of water given in (b).
   d) Discuss one;
      i) advantage of the type of water given in (b)(i).
      ii) disadvantage of the type of water given in (b)(i).
6. The graph in *Figure 4.19* shows the amount of carbon dioxide produced per continental Activity per year across the globe.

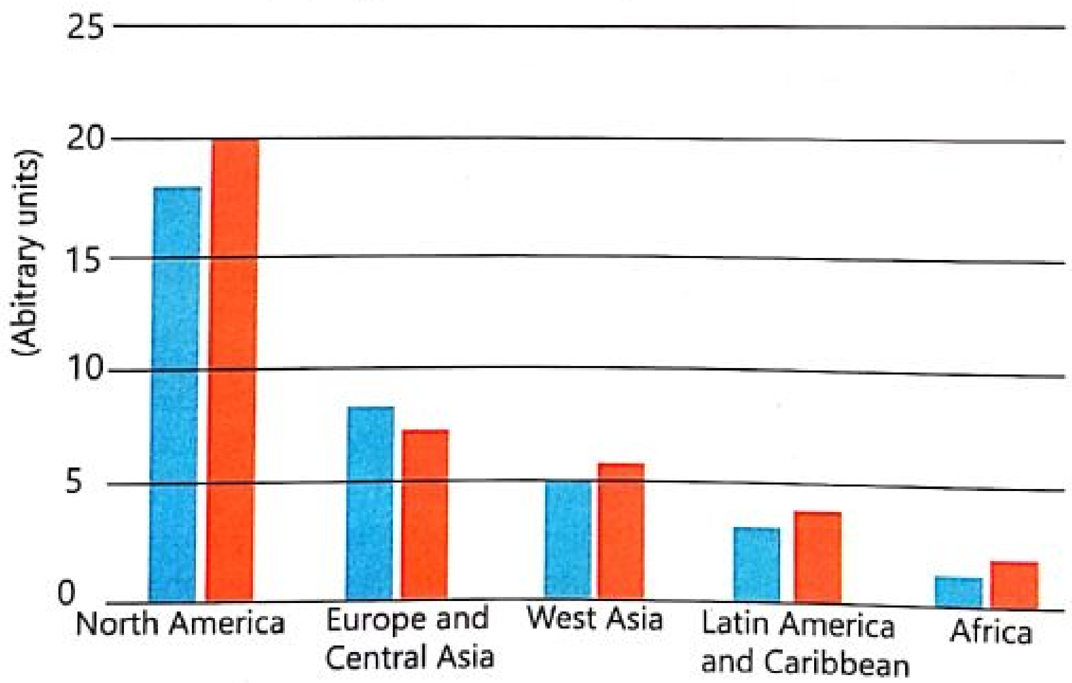
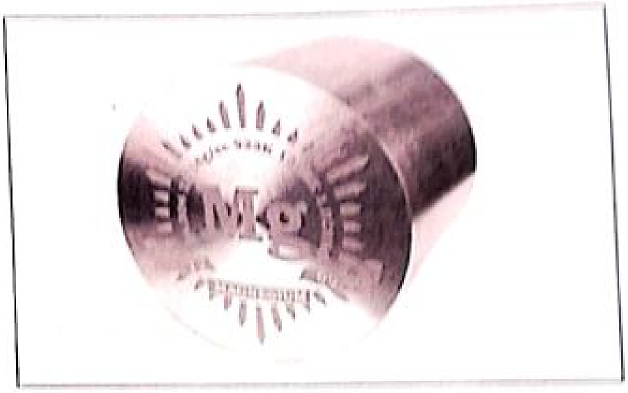


***Figure 4.19:** Amount of carbondioxide produced per continental activities per year across the globe*

   a) Comment on the graph.
   b) Explain why the amount of carbon dioxide generally changed.

CHAPTER 5 The Reactivity Series



Keywords

- alloys
- metals
- non-metals
- reactivity
- reactivity series
- displacement
- electrolysis
- corrosive

By the end of this chapter, you should be able to:

a) appreciate that metals vary in their chemical reactivity and can be arranged in a reactivity series.

b) understand that alloys are mixtures of a metal with other metals and / or non-metals and compare the properties of common metals with their alloys.

Competency: You should be able to evaluate data on reactivity in order to arrange metallic elements according to their reactivity.

Introduction

You have previously learnt that elements are divided into metals and non-metals from the Periodic Table in Chapter 3. At this level, we shall look at the reactivity of these elements.

Some of the different metals used in every day life are shown in *Figure 5.1*. The chemical reactivity of a metal is related to its tendency to form positive ions. From quite simple experiments, metals can be arranged in order of their reactivity. In this chapter, you will evaluate data on reactivity in order to arrange metallic elements according to their reactivity.

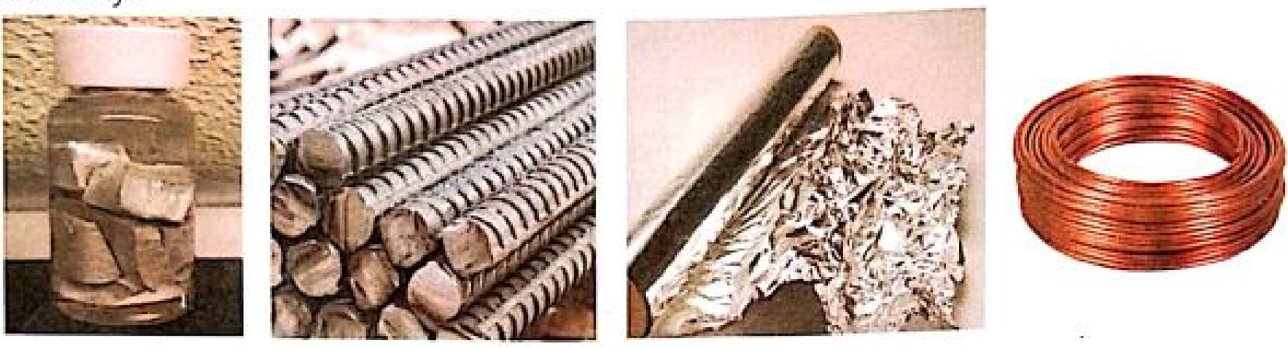


Figure 5.1: *Different metals*

5.1 What is the Reactivity Series?

Placing metals in their order of reactivity is termed as the ***reactivity series.***
The non-metals, carbon and hydrogen, are often included in the reactivity series, and this is quite important when considering the method by which the metal can be extracted from its ore. You should also note that a more reactive metal will displace a less reactive metal from its compounds.

How metals vary in their reactivity

In daily life, when you are stung by a bee, or cut with a razorblade, or pierced by a needle or a nail, you react differently. Some shout, others cry and others are rushed to the nearby health centres. Similarly, metals react differently when subjected to different conditions.

Reaction of potassium and sodium metals with cold water

Sodium and potassium are both group one elements but react differently with water. In *Activity 5.1(a)*, you will find out the reaction of potassium and sodium metals with water.

Safety precaution:

- Use gloves while handling potassium and sodium metals.
- Safety goggles should be worn before and during the experiment.
- Observations should be made at a reasonable distance.
- Potassium and sodium react vigorously that they catch fire if kept in the open. Therefore, to protect them and to prevent accidental fires, they are kept immersed in kerosene or oil.

Activity 5.1(a): Finding out what happens when potassium and sodium react with cold water

In this activity, you will work in groups.

What you need:

- litmus papers (red and blue)
- potassium metal
- sodium metal
- 2 filter papers
- water
- a trough
- a pair of tongs
- a knife

What to do:

1. Pour water in the trough.
2. Use a dry knife to cut a tiny piece (pea size) of sodium metal and use a filter paper to remove the oil from the metal.
3. Use a pair of tongs to drop it into a trough of cold water.
4. Carefully observe what happens to the metal.
5. Record your observations in *Table 5.1.*
6. Test the liquid in the trough using blue and red litmus papers and describe what happens to each paper.

Table 5.1: Reaction of potassium and sodium with cold water

Metal	Observation	Effect on litmus paper
Sodium		
Potassium		

7. Repeat procedures 1-6 using potassium metal.
8. Present your responses to the rest of the class.

Discussion Questions

1. Compare the vigour with which each metal reacts with water.
2. Which metal is more reactive?

Figure 5.2: Reaction of potassium with water

Figure 5.3: Reaction of sodium with water

Conclusion

From *Activity 5.1(a)*, you observe that potassium reacts violently with cold water to form potassium hydroxide and hydrogen gas.

Potassium + water → potassium hydroxide + hydrogen gas

You also observe that sodium reacts vigorously with cold water to form sodium hydroxide and hydrogen gas.

Sodium + water → sodium hydroxide + hydrogen gas

The reaction of potassium metal with cold water produces a lot of heat, which causes both potassium and hydrogen gas to burn producing a purple flame or a lilac flame.

Sodium and cold water also produce a lot of heat but not enough to burn sodium and hydrogen gas, showing that potassium is more reactive than sodium.

Reaction of other metals with cold water

Activity 5.1(b): Finding out what happens when other metals react with cold water

In this activity, you will work in groups.

What you need:

- a beaker
- a filter funnel
- a boiling tube
- water
- calcium metal
- magnesium
- aluminium
- zinc
- iron
- copper
- a splint
- a spatula

What to do:

1. Put a spatula endful of calcium at the bottom of the beaker containing some water.
2. Invert a funnel over calcium metal.
3. Add enough water to cover the whole funnel.
4. Quickly invert a boiling tube full of water and collect the gas as shown in *Figure 5.4*.
5. Carefully observe what happens to the metal.
6. When the boiling tube is full of hydrogen gas, close it with a cork while still under water.
7. Test the gas by placing a burning splint at the mouth of the test tube.
8. Test the liquid using blue and red litmus papers.
9. Record and discuss your observations.
10. A cloudy solution is formed when calcium reacts with cold water. Explain why this happens.
11. Repeat procedures 1-9 using magnesium, aluminium, zinc, iron and copper.

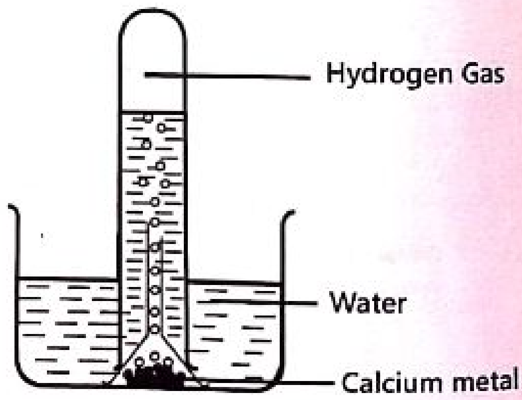


***Figure 5.4:** Action of water on calcium*

From *Activity 5.1(b)*, you observe that the reaction of calcium with water is less violent. The heat produced is not sufficient for the hydrogen gas to catch fire. Calcium sinks at the bottom of the beaker and reacts moderately with water, producing bubbles of hydrogen gas.

Magnesium reacts very slowly with cold water while aluminium, zinc, iron and copper do not react with cold water.

Metals which are very reactive, react with cold water to form their corresponding metal hydroxide solution and hydrogen gas.

From your observations, which is the most and least reactive metal among magnesium, sodium, potassium and calcium.

Activity 5.1(c): Investigating what happens when steam is passed over strongly heated magnesium, aluminium, zinc, copper and iron

In this activity, you will work in groups.

What you need:

- magnesium ribbon
- zinc
- iron
- a wet glass wool
- a heat source
- a boiling tube
- aluminium
- copper
- sand paper
- a cork with a glass tube

What to do:

1. Soak glass wool in water and insert it at the bottom of the boiling tube.
2. Clean a piece of magnesium ribbon using sandpaper.
3. Insert it in the middle of the boiling tube.
4. Stopper the boiling tube with a cork containing glass tubing.
5. Clamp the boiling tube as shown in *Figure 5.5*.

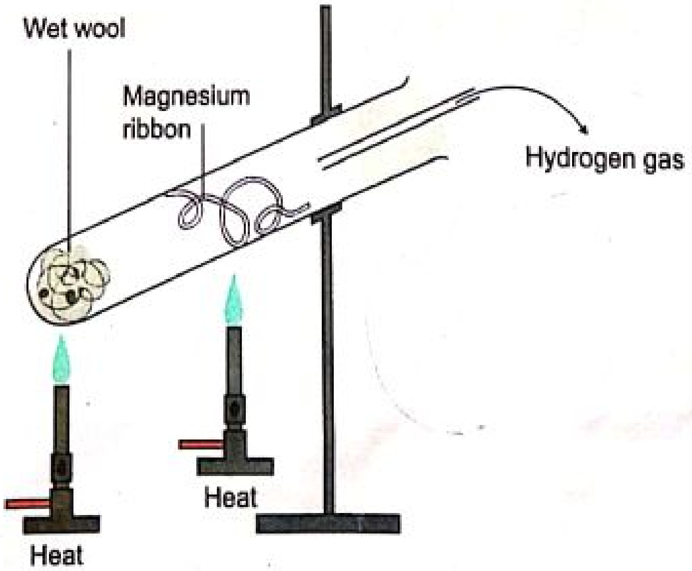


Figure 5.5: *Reaction between magnesium ribbon and steam*

6. Heat the magnesium ribbon strongly and then heat the wet glass wool gently.
7. Estimate the time for the magnesium ribbon to glow red hot.
8. Carefully record your observations.
9. Write a word equation for the reaction between magnesium and steam.
10. Repeat procedures 1-9 using aluminium, zinc, copper and iron filings respectively in place of the magnesium ribbon.
11. Present your responses to the rest of the class.

Discussion Questions

Using *Activity 5.1(c),* explain the following questions and present your responses to the rest of the class.

1. Why was the magnesium ribbon cleaned first?
2. Explain why the wet glass wool was heated.
3. Compare the reactivity of magnesium ribbon with cold water and with steam.

Conclusion

From your observations, you find that magnesium takes the shortest time to react with steam, while iron takes the longest time.

Metals which are moderately reactive do not react with cold water but react with steam to form corresponding metallic oxides and hydrogen gas. Those which are least reactive such as copper do not react with cold water or steam.

Using the observations in *activities 5.1(a), 5.1(b)* and *5.1(c),* complete *Table 5.2* by stating how each metal reacts with steam or cold water. Give the products of each reaction.

Table 5.2: Reaction of metals with cold water and steam

Metal	Reaction with cold water or steam	Product
Potassium		
Sodium		
Calcium		
Magnesium Aluminium		
Zinc		
Iron		
Copper		
Silver		
Gold		

The reaction of metals with water can, therefore, be used to produce the metal reactivity series from the most reactive to the least reactive. Using *Activities 5.1(a) 5.1(b)* and *5.1(c)* and results in *Table 5.2*, write the reactivity series of the metals.

Safety precaution:

Potassium and sodium are highly reactive and therefore are kept under oil as shown in *Figure 5.6* to prevent them from reacting with moist air.

Figure 5.6: Sodium stored under oil

The reactivity of metals with acids

Metals react with acids under different conditions. What are some of the conditions? Some metals react with some acids forming a salt but unlike the reaction between acids and bases, we do not get water. What then could be the second product formed from the reaction?

How quick the reaction goes depends on the metal used and its position in the reactivity series.

General equation:
Metal + acid → metal salt + hydrogen

Safety precaution:

It is not safe to add sodium or potassium to an acid in the laboratory. This is because such reactions are fast, explosive and produce a lot of heat energy which may break the apparatus.

Activity 5.1(d): Finding out what happens when dilute hydrochloric acid is added to calcium, lead, magnesium, zinc, iron and copper

In this activity, you will work in groups.

What you need:

- magnesium ribbon
- zinc pieces
- iron filings
- copper
- calcium
- a glass trough
- a filter funnel
- a thistle funnel
- water
- lead
- delivery tubes
- a gas jar
- dilute hydrochloric acid
- a round bottomed flask

What to do:

1. Put zinc pieces in a round bottomed flask.
2. Using a thistle funnel, add dilute hydrochloric acid into the flask.

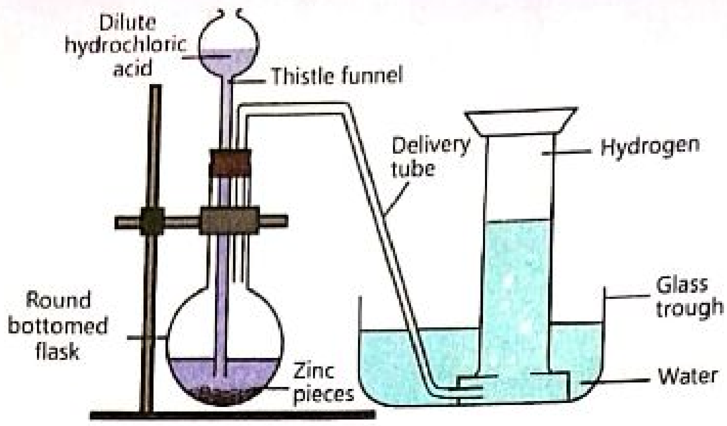


Figure 5.7: Laboratory preparation of hydrogen

3. Observe what takes place in the round bottomed flask and the rate at which the gas jar is being filled with hydrogen.
4. Repeat procedures 1, 2 and 3 using magnesium, iron, calcium, copper and lead.
5. Record your observations in *Table 5.3* and discuss your results with the rest of the class.

Table 5.3: Reaction of metals with hydrochloric acid

Metal	Rate of reaction	Products
Calcium		
Magnesium		
Zinc		
Iron		
Lead		
Copper		

Discussion Questions

1. From your discussions, arrange the metals in their order of reactivity.
2. Write a word equation for the reaction taking place when zinc metal reacts with hydrochloric acid.
3. Present your work to the rest of the class.

From *Activity 5.1(d)*, you observe that the rate of formation of hydrogen gas bubbles was fastest in calcium and that no bubbles were seen during reaction of copper and dilute hydrochloric acid.

Metals react with dilute acids to give a metal salt and hydrogen gas. Calcium and magnesium react rapidly with dilute acids. Metals like lead, zinc and iron react slowly with dilute acids, whereas copper does not react at all.

Water and acids contain hydrogen atoms that can be replaced and therefore when they react with metals, hydrogen atoms are displaced and combine to form hydrogen gas. Why do you think copper does not react with either an acid or water under all conditions?

Activity 5.1(e): Investigating the reaction between copper(II) oxide and hydrogen gas

This activity should be demonstrated by the teacher.

Safety precaution:

A mixture of hydrogen and air explodes when lit. All the air in the apparatus should be removed.

What you need:

- zinc granules
- dilute hydrochloric acid
- copper(II) oxide
- a combustion tube
- a tap funnel
- a rubber cork
- a heat source

What to do

1. Set up the apparatus as shown in *Figure 5.8*.
2. Put copper(II) oxide in a combustion tube and heat it strongly.
3. Add hydrochloric acid to the zinc granules and pass the dry gas produced through the apparatus until the pop sound stops.
4. Record what you observe in the combustion tube.
5. Write a word equation for the reaction that takes place in the combustion tube.
6. Present your observations to the rest of the class.

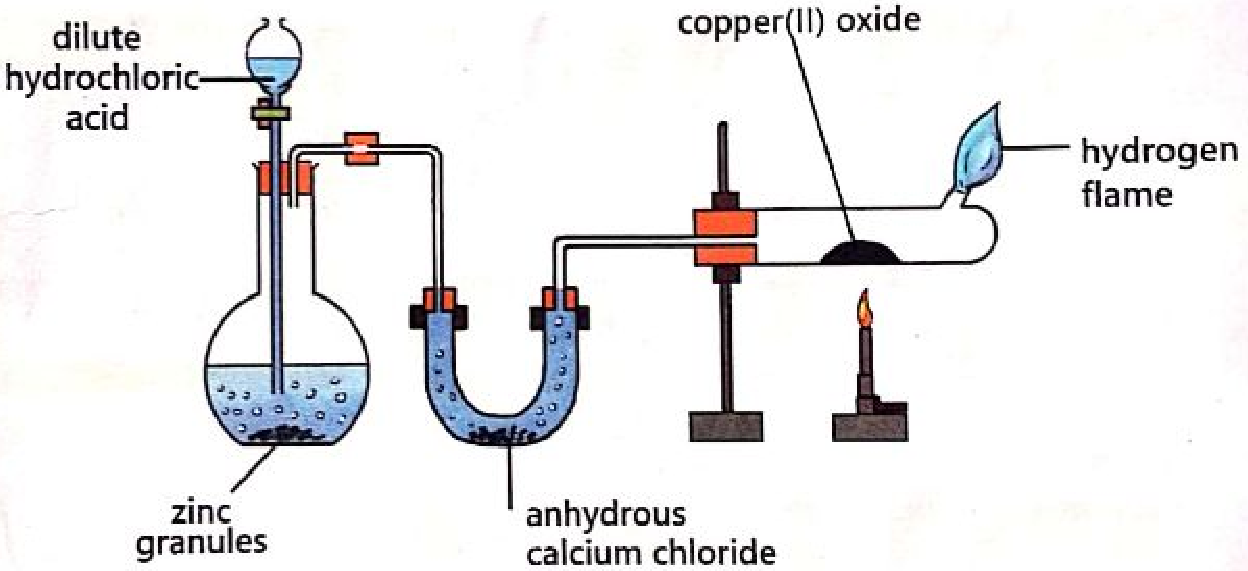


Figure 5.8: Reaction of copper(II) oxide and hydrogen gas

Conclusion

Hydrogen displaces copper atoms from their oxide which is black to form copper atoms which are brown. Hydrogen combines with oxygen to form water which is colourless. This shows that hydrogen is more reactive than copper.

Assignment 5.1

Discuss how you would measure and record the level of reactivity of the metals, and evaluate the limitations of the investigation.

For your understanding

When a metal reacts with water or hydrochloric acid, it displaces hydrogen atoms in the water or hydrochloric acid. This shows that the metal is more reactive than hydrogen. However, copper and silver do not react with water or acid. Therefore, they are less reactive than hydrogen.

Reaction of metals with oxygen

Oxygen supports burning and therefore, most of the metals when heated, react with oxygen to form corresponding oxides.

The vigour with which each metal reacts with oxygen depends on its position in the reactivity series. The reaction can be used to arrange elements in the order of their reactivity, beginning with the most reactive to the least reactive, since different metals react with oxygen at different rates.

The more vigorous a metal burns in oxygen, the higher its position in the reactivity series. There are different ways through which you can heat various metals to make them burn more efficiently in oxygen.

Figure 5.9: Burning of sodium in oxygen

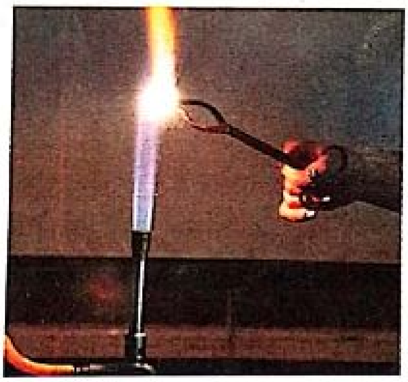


Figure 5.10: Burning of magnesium in oxygen

Figure 5.11: Burning of potassium in oxygen

Figure 5.12: Burning of calcium in oxygen

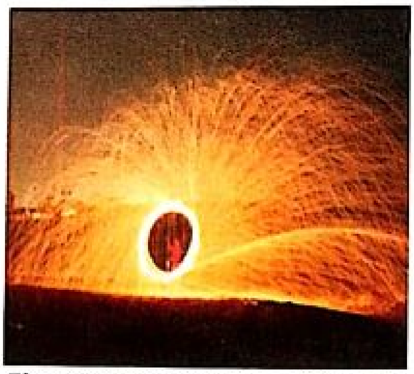


Figure 5.13: Burning of iron in oxygen

Figure 5.14: Burning of zinc in oxygen

Activity 5.1(f): Investigating the reaction of metals with oxygen

In this activity, you will work in groups.

What you need:

- sodium metal
- a deflagrating spoon
- water
- potassium
- magnesium ribbon
- calcium
- an oxygen gas jar
- litmus papers (red and blue)

What to do:

1. Ignite sodium metal on a deflagrating spoon and quickly lower it into a gas jar of oxygen.
2. Add a little water to the product and shake.
3. Carefully observe and record your observations.
4. Test the resultant solution with blue and red litmus papers and state the effect of the solution on each litmus paper.
5. Write a word equation for the reaction between the metal and oxygen.
6. Repeat the activity with calcium, potassium and magnesium ribbon.
7. Discuss your observations in your groups.
8. Present your responses to the rest of the class.

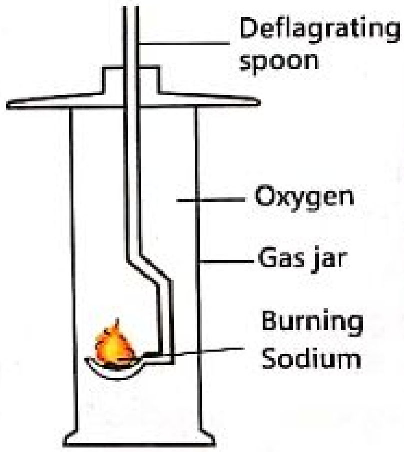


Figure 5.15: Burning of sodium in oxygen

Conclusion

Highly reactive metals burn in oxygen with a characteristic flame, forming a metal oxide.

General equation:
Metal + oxygen → metal oxide

Some less reactive metals do not burn in oxygen. They only react moderately even when heated strongly. In *Activity 5.1(g)*, you will find out the reaction of other metals when heated in oxygen.

Activity 5.1(g): Finding out what happens when other metals are heated in oxygen

In this activity, you will work in groups.

What you need:

- iron wire or steel wool
- a deflagrating spoon
- an oxygen gas jar
- zinc
- copper
- aluminium
- a heat source
- a centimetre ruler

- a measuring cylinder
- red and blue litmus papers
- a pair of tongs
- magnesium ribbon

What to do:

1. Using a centimetre ruler, measure 4 cm of magnesium and clean it using steel wool.
2. Hold the ribbon using a pair of tongs and heat it over a Bunsen flame.
3. When it starts to burn, lower it into a gas jar of oxygen but do not drop it in.
4. Measure and add 5 cm^3 of distilled water to the product.
5. Test the resultant solution with red and blue litmus papers.
6. Attach a small piece of iron wire to the end of a deflagrating spoon.
7. Heat until the wire is red hot.
8. Lower the glowing wire into the gas jar of oxygen.
9. Add 5 cm^3 of distilled water to the product and test the resultant solution with red and blue litmus papers.
10. Repeat procedures 6-9 using a strip of zinc, copper and aluminium.
11. What conclusion do you derive from the observation.

Conclusion

Aluminium, zinc and lead are covered with a thin protective layer of oxide which prevents the metal from further oxidation. Therefore, they must be cleaned to remove the layer of the oxide in order to allow them combine with the oxygen on strong heating . Aluminium, zinc, iron and copper do not burn with a characteristic flame but combine readily with oxygen when heated strongly to form their corresponding oxides.

From *Activities 5.1(f)* and *5.1(g)*, you found out that sodium and potassium readily burn in oxygen compared to other metals and that copper takes the longest time to react with oxygen. Some metals are inactive and do not react with oxygen at all, even under very strong heating, for example gold and silver.

Competition for oxygen and ions between metals

You saw that metals can be arranged in order of their reactivity, using their reactions with cold water, steam, dilute acids and oxygen. Now, you will see what happens when metals compete with each other, and with carbon to form a compound.

Activity 5.1(h): Competition for oxygen between magnesium and copper

In this activity, you will work in groups.

What you need

- copper(II) oxide
- magnesium powder
- magnesium oxide
- copper powder
- heat source
- a crucible
- a spatula

What to do:

1. Transfer one small spatula endful of magnesium powder in a crucible.
2. Add one small spatula endful of copper(II) oxide powder to the same crucible.
3. Mix the two powders together using a spatula.
4. Heat the mixture strongly.
5. What do you observe?
6. Repeat procedures 1-5 using a mixture of magnesium oxide and copper powder.

Figure 5.16(a): *Copper(II) oxide powder*

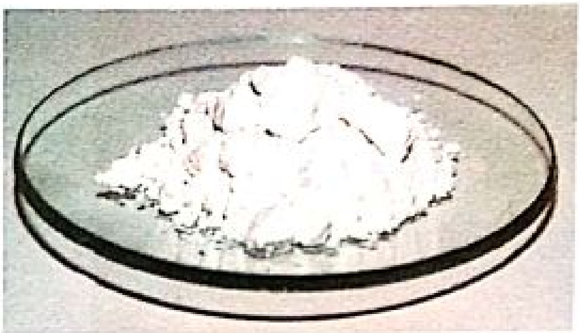


Figure 5.16(b): *Magnesium oxide*

Discussion Questions

1. From your observation, which of the two elements is more reactive than the other.
2. Explain the reasons for your answers.

Competition for oxygen between aluminuim and iron

You have realised that a more reactive metal has a greater tendency to form ions by losing electrons than a less reactive metal. Therefore, if a more reactive metal is heated with an oxide of a less reactive metal, it will remove the oxygen from it. Aluminium is more reactive than iron. Therefore, when iron(III) oxide is mixed with aluminium and the mixture is heated using a magnesium fuse, a very violent reaction occurs as the competition between the aluminium and the iron for the oxygen takes place.

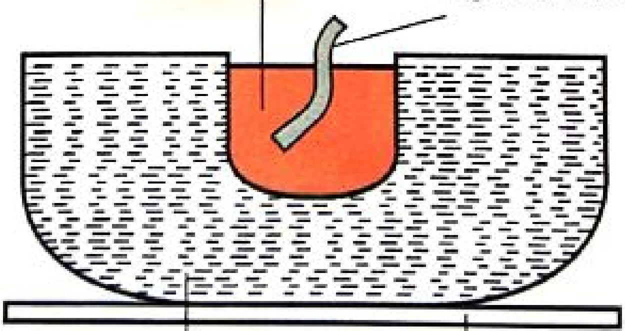


Figure 5.17: *Heating aluminium and iron(III) oxide using a magnesium fuse*

Aluminium, being the more reactive metal, removes oxygen from the less reactive iron. It is a very ***exothermic reaction.*** When the reaction is over, a solid lump of iron is left along with a lot of white aluminium oxide powder.

This particular reaction is known as the ***Thermic reaction.*** Since large amounts of heat are given out and iron is formed in a molten state, this reaction is used to weld together damaged railway lines. It is also used in making explosive bombs.

Figure 5.18: *Molten iron used for welding*

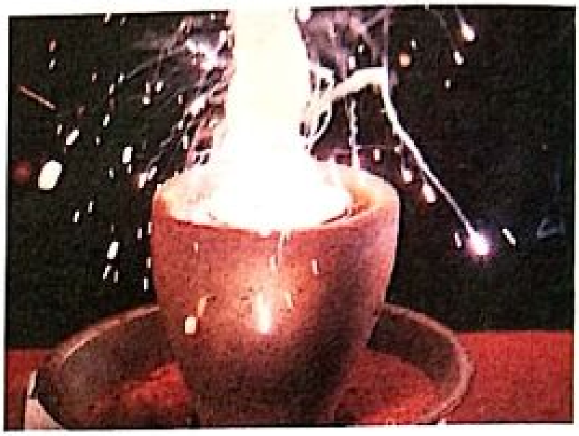


Figure 5.19: *Molten iron used in making explosive bombs*

For your understanding

Aluminium is more reactive than iron. However, we can use it for television aerials, satellite dishes and ladders without protecting it as shown in *Figures 5.20 (a)* and *5.20 (b)*. Why?

Figure 5.20(a): *Aluminium ladder*

Figure 5.20(b): *Satellite dish made of aluminium*

Aluminium protects itself! It reacts rapidly with oxygen, forming a thin coat of aluminium oxide that you cannot see. This sticks tight to the metal, acting as a barrier preventing further corrosion. Therefore, aluminium metal behaves as if it is nonreactive.

Activity 5.1(i): Investigating affinity for oxygen between magnesium, carbon and zinc

What you need:

- magnesium oxide
- a spatula
- zinc oxide
- carbon (charcoal powder)
- heat source
- a crucible

What to do:

1. Transfer one small spatula endful of magnesium oxide on a crucible.
2. Add one small spatula endful of charcoal powder (carbon).
3. Mix magnesium oxide and charcoal powder together using a spatula.
4. Heat the mixture strongly on a crucible.
5. What do you observe?
6. Repeat procedures 1-5 using a mixture of zinc oxide and charcoal powder (carbon).

Figure 5.21: *Reaction between charcoal and zinc oxide*

Discussion Questions

1. From your observation, identify which of the three elements is more reactive than the others.
2. Explain the reasons for your answers.

Conclusion

In each case, you should look for signs showing that the reaction has occurred. Where the reaction has occurred, the mixture will glow red hot, showing that carbon is more reactive than the metal. Where there is no observable change, it shows that carbon is less reactive than the metal in the oxide.

You will find out that when a mixture of magnesium oxide and carbon is heated, there will be no observable change but a heated mixture of zinc oxide and carbon changes from a white-black powder to grey.

Zinc oxide reacts with carbon to form zinc and carbon dioxide. This means that carbon is more reactive than zinc but less reactive than magnesium.

Electron structure and reactivity of elements

In Chapter 3, you found out that group I elements readily form positive ions compared to group II elements. This is because group I elements have one electron in their outermost energy.

For elements in the same period, the less the number of electrons required to be lost to form a stable ion, the more reactive the element is.

The size of an atom also determines the ease with which an atom loses an electron. The bigger the atom, the more reactive an element will be. When the atom is big, the electrons are far from the nucleus and therefore require less energy to be removed. The atom which easily loses electrons displaces those which do not easily lose electrons.

Assignment 5.2

Comparing the electronic structure of an element and its reactivity

a) Individually, research and draw the electronic structure of magnesium, aluminium and sodium atoms.

b) Arrange the structures in (a) beginning with the most reactive.

c) Explain your order in (b) above.

Activity 5.1(j): Investigating the reaction between iron and copper(II) sulphate

What you need:

- iron wool
- copper(II) sulphate solution
- beaker
- a measuring cylinder
- a stop watch

What to do:

1. Using a measuring cylinder, measure 100 cm^3 of copper(II) sulphate solution and pour it in a clean beaker.
2. Add iron wool to the copper(II) sulphate solution.
3. Leave it to stand for one minute.
4. What do you observe?
5. Suggest a reason for your observation.
6. Write a word equation for the reaction.
7. Repeat procedures 1, 2 and 3 using copper and silver nitrate solution.

Figure 5.22: *Iron wool reacting with copper(II) sulphate*

Conclusion

Displacement simply means moving something from its place. A displacement reaction is a reaction involving an element replacing another element from its compound in an aqueous solution or its oxide. The displacement reaction occurs when an element takes the place of another element in a compound.

When copper is added to a solution of silver nitrate, it reacts with silver nitrate to form hair like crystals of silver metal and a blue solution of copper(II) nitrate.

General equation:
Copper + silver nitrate → copper(II) nitrate + silver

Extraction of metals using charcoal

Generally, elements displace other elements below them in the reactivity series. Metals lower in the reactivity series are extracted from their oxides using carbon, which readily removes oxygen from their oxides to form a metal and carbon dioxide. For example, copper can be extracted from its oxide by heating copper(II) oxide with charcoal.

Activity 5.1(k): Investigating the reaction between carbon and copper(II) oxide

In this activity, you will work in groups.

What you need

- charcoal powder
- copper(II) oxide
- 2 test tubes
- a test tube holder
- lime water
- a heat source
- a spatula
- 2 petri dishes
- test tube holder
- heat source

What to do

1. Put a spatula endful of copper(II) oxide into a clean petri dish.
2. Add one spatula endful of charcoal powder to the copper(II) oxide.
3. Mix the two powders thoroughly using a spatula.
4. Transfer the mixture into a clean, dry test tube and set up the apparatus as shown in *Figure 5.23.*
5. Heat the mixture strongly until there is no further change.
6. Allow the test tube to cool on its holder.
7. Pour out the cooled mixture into a clean petri dish.
8. Carefully observe the changes that are taking place and record them in your notebook.
9. Present your observations to the rest of the class.

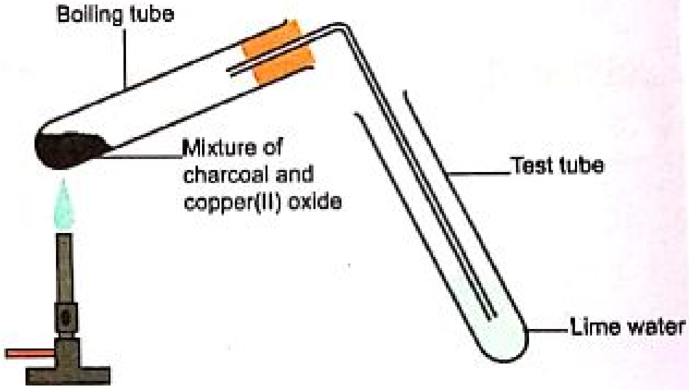


***Figure 5:23:** Set up of the reaction between carbon and copper(II) oxide*

Discussion Questions

1. Write a word equation for the reaction that takes place in the boiling tube.
2. Discuss other metals that can be extracted from their oxide by carbon.

Conclusion

If a metal is less reactive than carbon, it can be extracted from its oxide by heating the oxide with carbon. The carbon displaces the metal from its oxide by removing oxygen from the oxide, leaving the metal behind.

For your understanding

Because they are easily obtained from their ores, less reactive metals have been known and used for thousands of years. For example, copper has been in wide use for 6,000 years, and iron for 3,500 years.
However, the more reactive metals like sodium and potassium had to wait until the invention of ***electrolysis,*** in 1800, for their ***extraction.***

Assignment 5.3

1. Research and discuss why aluminium, group I and group II elements cannot be extracted on a charcoal block.
2. Gold occurs naturally in an uncombined state and can be recovered in a pure form from alluvial deposits. Explain this statement.

Activity 5.1(l): Arranging metals, carbon and hydrogen in order of reactivity series

1. Using *Activities 5.1(f), 5.1(g), 5.1(h), 5.1(i), 5.1(j)* and *5.1(k)* arrange the elements in order of reactivity starting from the most reactive.
2. To which group in the Periodic Table does the most reactive metal belong?

Note

Is the reactivity series of metals of any importance?

Relative reactivity of metals is important when considering the uses of metals. This is especially when they are exposed to conditions where they may corrode or when a metal is required to be isolated from a solution containing its ions. It is used to determine the products of single displacement reactions, whereby metal **X** will replace another metal **Y** from its solution if **X** is higher than **Y** in reactivity series.

Assignment 5.4

1. Research and discuss the application of reactivity series in daily life.
2. Present your responses to the rest of the class.

5.2 Alloys

Figure 5.24: Common items made of alloys in our daily life

Almost every material we could ever want is found somewhere on earth under our feet. From the gold we wear as jewellery and iron used in cutlery and roofing, to the oil products that are used as fuels.

Most of the naturally occurring elements are metals and they are useful. However, sometimes metals become less perfect for the purpose and the uses they are required to do. What are some of the examples of the inefficiencies shown by metals? Pure iron is a very strong metal but also rusts easily in damp air.

What about aluminium? Pure aluminium is very light and too weak to perform its purpose efficiently.

To improve on the ***efficiency, appearance*** and ***strength*** of metals, two or more metals or carbon are mixed to form ***an alloy.***

What is an alloy?

Alloys are found everywhere around you. Have you ever used an alloy in your everyday life? What are some of the examples of alloys that you know? Are alloys important in everyday life? Why do you think that people prefer alloys to pure metals? Let's take a closer look at what alloys are and why they are so useful.

Assignment 5.5

1. Discuss what an alloy is.
2. Identify examples of objects made using alloys and their corresponding uses in our daily lives.

The objects you have identified in *Assignment 5.5* are obtained by mixing elements of different metals or carbon. From your findings, you have discovered that most of the objects you use in everyday life are made up of two or more different metals and sometimes, contain a mixture of carbon.

DID YOU KNOW?

Save money and the environment!

Iron and steel are the most recycled metals in the world. They are easily recycled by melting them in a furnace at a very high temperature to produce a block or sheet which is sold to manufacturers as a metal product.

Old cars, fridges, washing machines, cookers can all get recycled. It costs far less to recycle them than obtaining new ore to extract the iron and it is better for the environment. It is easy to separate them from other scrap using a magnet as shown in *Figure 5.25*.

Figure 5.25: A magnet picking iron scrap

Discussion Questions

1. Identify the uses of alloys in our every day life.
2. Why are alloys preferred to pure metals.

Conclusion

Alloys are substances made by melting two or more metals or carbon. They crystalise upon cooling into a solid mixture. They have properties different from metals from which they are made. What are the compositions of different alloys?

Figure 5.26: Bronze is often used for making sculptures

Assignment 5.6

Finding out the composition and uses of alloys.

What to do:

1. Carry out a research on the composition and uses of the following alloys:
   a) brass
   b) bronze
   c) duralumin
   d) stainless steel
   e) steel
   f) nichrome
   g) solder
2. Record and present your findings to the rest of the class.

Activity 5.2: Identifying why alloys are preferred to metals

In this activity you will work in groups.

What to do:

1. Use the internet and other sources to discuss the properties of alloys.
2. Present your findings to the rest of the class.

ICT Activity

In groups, use the internet and other sources to research and watch videos about how different alloys are made. Record the composition of each alloy and use a presentation software to present your findings to the rest of the class.

Sample Activity of Integration

The job of blacksmiths, *Figure 5.27*, has existed for a long time while crafting low value items such as spear heads, knives and axes as shown in *Figure 5.28* from iron ore. However, this skill is declining due to low incomes from the low value items produced by the blacksmiths.

Figure 5.27: Iron smelting

Figure 5.28: Iron tools

Instruction

As a learner of Chemistry, you are tasked to promote a declining skill of blacksmiths. Write a message advising on how you can increase their earnings by making more valuable items from existing mineral elements.

Also in your message:

1. State the purpose of three different valuable products.
2. Explain how they can be produced.
3. Explain why they have to make high quality products.

Chapter Summary

In this chapter, you have learnt that:

- The more easily a metal atom can give up electrons, the more reactive the metal will be.
- Reactivity series is used to rank metals, hydrogen and carbon according to their reactivity.
- The most reactive metals are placed at the top and the least reactive metals are placed at the bottom.
- Reactivity series is used to investigate reactivity of metals with different substances.
- The more reactive the metal, the more stable its compounds are. Such compounds do not easily break down.
- The more reactive the metal, the more difficult it is to extract it from its ores, since these are stable compounds. For the most reactive metals, you need an advanced method of extraction such as electrolysis.
- The less reactive the metal, the less likely it is to form compounds. That is why copper, silver and gold are found as elements in the Earth's crust. Other metals are always found as compounds.

Metal	Symbol
Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Aluminium	Al
Carbon	C
Zinc	Zn

Metal	Symbol
Iron	Fe
Tin	Sn
Lead	Pb
Hydrogen	H
Copper	Cu
Silver	Ag
Gold	Au
Platinum	Pt

- Highly reactive metals react with cold water, forming a corresponding hydroxide and hydrogen gas.
- Moderate reactive metals react with steam to produce corresponding metal oxides and hydrogen gas.
- Metals such as copper, silver and gold do not react with cold water or steam.
- Metals higher than hydrogen in the reactivity series react with dilute acids except nitric acid to form a salt and hydrogen gas.
- Metals below carbon in the reactivity series are extracted from their metal oxides using carbon.

- Metals react with oxygen, forming metal oxides.
- An alloy is a uniform mixture of two or more metals and at times a non metal.
- Alloys are stronger compared to their constituent metals.

Review Exercises:

1. What is meant by the term reactivity series?
2. Explain why the reactivity series is important in Chemistry.
3. What happens when magnesium is burnt in carbon dioxide?
   a) Identify the products formed.
   b) Write a word equation for the reaction that takes place.
4. During the cleaning of a laboratory, the laboratory assistant accidentally dropped a piece of magnesium into iron(II) sulphate solution.
   a) What happened to the magnesium in the solution?
   b) Write a word equation for the reaction that took place.
5. Discuss the physical properties of alloys which make them of a better use than pure metals.
6. Explain the following:
   a) Sodium is not found as a free element in nature.
   b) Gold has been known and used for thousands of years longer than aluminium. Explain why.
7. Elements A, B, C, D and E were dipped separately in each of the nitrate solutions as shown in the table below.
   Tick (✓) means a reaction occurred. Cross (×) means there was no observable change.

		Nitrates of				
		A	B	C	D	E
Metals	A	–	✓	✓	✓	×
	B	×	–	×	×	×
	C	×	✓	–	✓	×
	D	×	✓	×	–	×
	E	✓	✓	✓	✓	–

Arrange the five elements in order of their reactivity starting with the most reactive element.

While investigating how fast metals can react with water, learners of Light Secondary School placed 6 metals (M, T, N, U, B, C) in each conical flask containing 200 cm^3 of water and measured the volume of hydrogen gas produced after 2 minutes for every metal as shown in *Figure 5.29*.

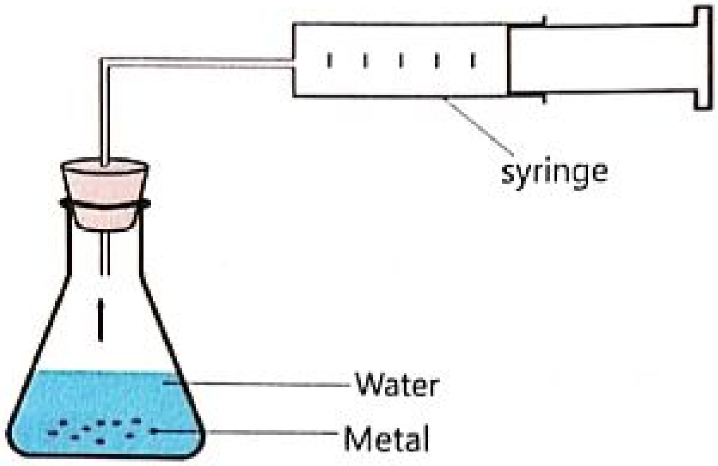


Figure 5.29: Placing metals in water

Metal	Volume of hydrogen produced After 2 minutes (cm^3)
M	0
T	5
N	10
U	8
B	20
C	29

a) Arrange the elements in order of decreasing reactivity.

b) Given that the elements were selected from group I and group II of the Periodic Table, identify any two elements that are likely to be members of group I.

c) For the 2 elements identified in (b) above, state the element that is likely to be from;

i) Period 3

ii) Period 4 of the Periodic Table

d) Give a reason for your choice in c) (ii).

GLOSSARY

acid: a substance which dissolves in water to produce hydrogen ions as the only positive ions in the solution

acidic: having a pH less than 7

alkali: a soluble base

alkaline: having a pH greater than 7

allotropes: two or more different physical forms in which an element can exist

alloy: a uniform mixture of two or more metals

ammonia: a compound of nitrogen and hydrogen with the formula NH_3

anion: a negatively charged ion

antacid: a substance that neutralises stomach acidity

ascorbic acid: another name for vitamin C

atmosphere: space around us

atom: the smallest electrically neutral indivisible particle of an element that can take part in a chemical reaction

atomic number: the number of protons in the nucleus of an atom

base: a substance that dissolves in water to release hydroxide ions as the only negative ions in a solution

bond: an attraction between atoms, ions or molecules that enables formation of chemical compounds

carbon: a chemical element with an atomic number 6 and a symbol C

carbonated drinks: drinks that contain dissolved carbon dioxide

cations: positively charged ions

chemical equation: symbolic representation of a chemical reaction inform of symbols and formulae

chemical formula: chemical symbols showing the elements present in a compound and their relative proportions

conductors: materials that contain movable electric charges

covalent bond: a chemical bond that involves sharing of electron pairs between atoms

corrosive: a substance that damages or destroys other substances which come into contact with it

displacement reaction: a reaction involving an element displacing another element from its compounds in an aqueous solution or its oxide

double decomposition: a reaction involving an acid reacting with a soluble salt in a solution to form a soluble salt and an insoluble salt

electrolysis: technique of using direct current to drive a non-spontaneous chemical reaction

electronic configuration: distribution of electrons of an atom in an atomic orbital

environment: everything around us

flammable: easily set on fire

fractional crystallisation: method of separating substances based on differences in their solubility

fuel: any substance we use to provide energy

groups: columns of elements in the Periodic Table of chemical elements

greenhouse gases: a gas that absorbs and emits radiant energy with in the thermo infra-red range

hard water: water that does not easily form lather with soap

helium: a chemical element with the symbol he and atomic number 2

hydroxide: an anion that consists of oxygen and hydrogen atoms held together by the covalent bond and carries a negative electric charge

indicator: a special substance used to test whether a given substance is an acid, an alkali or neutral

ion: an atom or molecule that has a net electrical charge

lime water: common name for calcium hydroxide

mass number: total number of protons and neutrons in the nucleus of an atom

metal: an element that forms positive ions

metal carbonates: bases that undergo reaction with an acid to give a salt, water and carbon dioxide

mineral acid: any inorganic acid obtained from minerals

neutral atom: an atom where the charges of the electrons and the protons are equal

neutralisation reaction: a reaction where an acid reacts with a base to from a salt and water only

neutralise: make an acidic or alkaline substance chemically neutral

neutron: a particle with no net electric charge

noble gases: nonreactive gaseous elements like helium, neon, argon, krypton and radon

non-metal: elements that tend to gain electrons to form anions during chemical reactions

nucleus: a positively charged centre of the atoms consisting of protons and neutrons

overall charge: the number of protons minus the number of electrons

Periodic Table: tabular display of the chemical elements which are arranged by atomic number, electron configuration and chemical properties

period: a horizontal row of the Periodic Table

pH scale: shows how acidic or basic a substance is

precipitate: an insoluble salt that emerges from a liquid

proton: particle with a positive electric charge of +1

pure water: water without any dissolved salts and gases

reactant: substances that are initially present in a chemical reaction and are consumed during the reaction to form products

reaction: a process through which reactants are converted to products

reactivity series: arrangement of elements from the highest to the lowest basing on their reactivity

renewable: can be used up and can be replaced after their use

salt: a substance formed when an acid reacts with a base

solubility: the mass of salt in grams that is required to saturate 100g of water at a given temperature

solubility curve: a graph based on data comparing the amount of solute that will dissolve in a given amount of solvent at various temperatures

solution: a homogenous mixture composed of two or more substances

solvent: a substance that dissolves in a solid, liquid or gas to form a solution

sustainable: ability to exist constantly

zinc granules: small solid crystals of zinc

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Baroque Senior Two Chemistry Learner's Book has been developed basing on the ***New Lower Secondary School Curriculum***, whose aim is to enable learners to make informed decisions as citizens and family members and to give them the learning skills they need to think critically and study efficiently.

The Learner's Book has been designed to allow learners to work individually and with others in practical and interactive activities that are related to theoretical concepts in this course. It is expected that learners will apply investigative and problem-solving skills, communicate scientific information effectively, and appreciate the contribution that the study of Chemistry makes to their understanding of the world.

The Learner's Book caters for the different teaching and learning styles and needs in order to enable learners to develop skills and knowledge that will be of long-term value in a world where Chemistry is always increasingly important, rather than just focusing on large quantities of factual information.

The following aspects have been given consideration while developing the content of this book; respect for human life, awareness of the importance of living in harmony with the environment, independent and critical thinking, and the innovative application of science and technology in solving problems.

On completion of this course of study, the learner should be able to;

- tell the significance of Chemistry in a modern scientific world,
- explain the relevance of Chemistry in everyday life,
- explain the role of Chemistry in enabling materials to be used in the service of mankind,
- develop the ability to work independently and collaboratively with others,
- describe the importances and limitations of science in relation to social and economic development.

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