DEPARTMENT OF EDUCATION

GRADE 11

CHEMISTRY

MODULE 3

TYPES OF CHEMICAL REACTIONS

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PRIVATE MAIL BAG, P.O. WAIGANI, NCD
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Writer
Evelyn Dimacale

Content Editors
Science Department
Subject Review Committee

Language Editor
Dr. Steven Winduo

Course Format Editor
Joydee Mabbagu

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CHEMISTRY

MODULE 3

TYPES OF CHEMICAL REACTIONS

IN THIS MODULE YOU WILL LEARN ABOUT:

11.3.1: INDICATORS OF CHEMICAL CHANGE

11.3.2: TYPES OF REACTIONS

Acknowledgements

We acknowledge the contribution of all Lower and Upper Secondary Teachers who in one way or another helped to develop this Course.

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MR. DEMAS TONGOGO
Principal-FODE

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SECRETARY'S MESSAGE

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

This course is part and parcel of the new reformed curriculum. The learning outcomes are student-centred with demonstrations and activities that can be assessed.

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

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

The course promotes Papua New Guinea values and beliefs which are found in our Constitution and Government Policies. It is developed in line with the National Education Plans and addresses an increase in the number of school leavers as a result of lack of access to secondary and higher educational institutions.

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

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

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

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

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

DR. UKE KOMBRA, PhD
Acting Secretary for Education.

MODULE 3: TYPES OF CHEMICAL REACTIONS

Introduction

We are able to survive in this world because of the chemical reactions that are going on all around us, all the time. In Grades 9 and 10 you have learnt the use of the Periodic Table, concept of writing chemical formulae and equations. In this Module, you will learn to identify and classify chemical reactions and write balanced chemical equations.

It is also important to revise your Science 9 and 10 to understand the content of this module.

This course module consists of two (2) topics:

1. **Topic 11.3. 1 Indicators of Chemical Change**

   In this topic you, will learn some indication of chemical change that occurs during a chemical reaction. The combining reactants give a visible sign, usually by a colour or temperature change or a gas is given off.

2. **Topic 11.3.2 Types of Reactions**

   In this topic, you will identify and describe different types of chemical reactions as a process of changing chemical substances that take part in a chemical reaction and substances that are produced. A simplified explanation to the types of reactions will be discussed later in the following pages.

Good luck in studying this course module. We hoped you will the module interesting regarding types of chemical reactions.

Learning Outcomes

After going through the Module, you are expected to:

- identify and name the types of chemical reactions.
- write balanced chemical equations for various chemical reactions.
- write gross ionic and net ionic equations and identify spectator ions.
- explain whether a given chemical species is undergoing an oxidation or reduction process.
- define and describe different types of chemical reactions as exothermic or endothermic.

Time Frame

10 weeks

This module should be completed within 10 weeks.

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

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

DO NOT LEAVE ANY QUESTION UN-ANSWERED.

Terminologies

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

A**lkali**	Is just a base that can dissolve in water and a base will not dissolve in water.
Aqueous solution	Means when an ionic solid dissolves in water.
Base	Is an opposite of acid that will not dissolve in water.
Chemical changes	Are changes in the composition and structure of a substance which form new substances.
Decomposition reaction	Is a type of chemical reaction where one reactant yields two or more products. It is just an opposite of a synthesis or combination reaction.
Displacement reaction	Is a type of reaction in which a less reactive element is displaced by a more reactive element in a chemical reaction.
Double displacement reaction	Is a type of reaction when two ionic compounds are mixed together.
Law of Conservation of Mass	States that matter can neither be created nor destroyed in a chemical reactions, mass does not change.
Neutralization	The reaction between an acid and an alkali (base).
Oxidation	Is the loss of electrons, gain of oxygen, loss of hydrogen, and increased in oxidation number.
Periodic Table	Is a table where chemical elements are arranged in order of increasing atomic number.

Precipitation reaction	Is the formation of insoluble solids in a reaction. Precipitates usually form when two metal aqueous salts are combined.
Redox reaction	Is a reaction in which one substance is reduced and the other substance is oxidized.
Reduction	Is the gain of electrons, loss of oxygen, gain of hydrogen, and decreased in oxidation number.
Single displacement reaction	Is a type of reaction where a more re-active element displaces another less re-active element from a compound.
Solubility rules	Describe which compounds are soluble (dissolves) and which are insoluble (does not dissolves).
State symbols	Are added to the reactants and products after completing the symbol equation.
Symbol equation	Shows that the word equations are replaced with the symbols of the elements.
Synthesis reaction	Is a type of reaction where a substance is formed as a result of chemical combination**.**
Chemical equations	Represent the names of the reactants and products that are written on each side of the arrow to make a word equation for the reaction.

11.3.1 Indicators of Chemical Changes

There are many types of chemical reactions. The following are common ways of deciding that a chemical reaction has occurred.

(a) If there is a significant change in the temperature.
(b) If a solid or precipitate is formed.
(c) If there is a change in colour.
(d) If a gas is evolved or given off.
(e) If there is disappearance of solid which is not physical dissolution of the solid in the solvent.

Change of Temperature and Formation of Precipitate
Change of Temperature
Some reactions are **exothermic**, which means they release heat, while others are **endothermic**, they absorb heat. By checking the temperature of the reaction mixture using a thermometer, it is possible to note if there is a reaction going on even if the formation of solid or a colour change is absent.

Have you ever observed your gas stove while in use? What is the colour of its flame? During combustion (burning), liquefied petroleum gas (LPG) reacts with oxygen to form carbon

dioxide and water vapour. In the process, it releases energy in the form of heat and light, for example, the flame. There are chemical reactions wherein no products are seen. A change in temperature or the light can also indicate that a chemical reaction is taking place.

The effect of temperature in a chemical reaction

If you increase the temperature of a reactant molecule, the reaction will become faster, and if you lower the temperature, it will make the reaction slower. Reactants must move fast and hit each other hard enough for a chemical reaction to take place. Increasing the temperature increases the average speed of the reactant molecules. As more molecules move faster, the faster the formation of the products.

Formation of Precipitate (undissolved solid particles)

Some reactions are accompanied by precipitation of solids which are quickly seen or observed. A **precipitate** forms when two aqueous solutions react together. **Aqueous solution** means when an ionic solid dissolves in water.

Most precipitates are white. Those containing transition metal ions, however, often have a characteristic colour such as b lue forcopper(II)hydroxide ($Cu(OH)_2$), and green for iron(II) hydroxide ($Fe(OH)_2$). Some precipitates appear as small, separate particles in solution, while others, especially those containing hydroxide ions (OH^-) may appear gelatinous or "jelly-like".

For example, if you add a drop of lead nitrate solution (a clear solution) to a solution of sodium iodide, solid particles will appear. These solid particles are called precipitates.

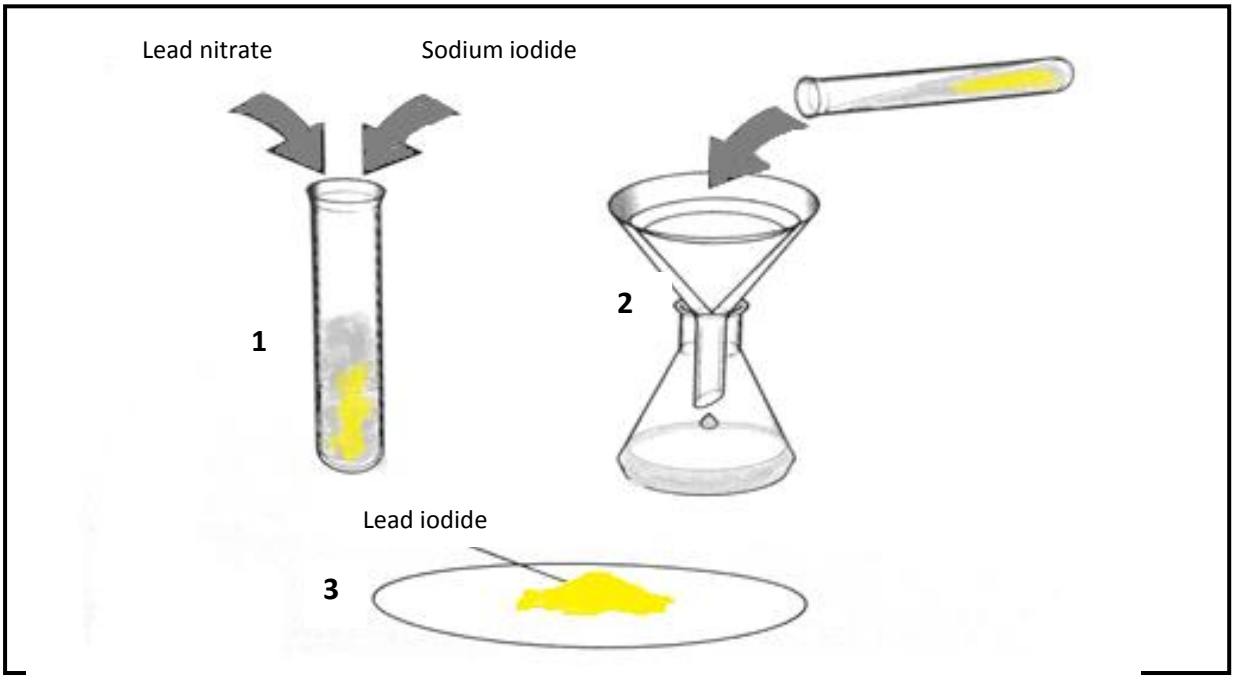
The equation for this reaction is as follows.

Sodium iodide	+	lead (II) nitrate	→	sodium nitrate	+	lead iodide
$2NaI_{(aq)}$	+	$Pb(NO_3)_{2(aq)}$	→	$2NaNO_{3(aq)}$	+	$PbI_{2(s)}\downarrow$

Note: An arrow downwards symbolizes the formation of precipitates.

Precipitates are removed from the remaining solution either by filtration (the separation of solid) or by allowing them to settle before decanting (pouring) off the unwanted solution.

The yellow precipitate (insoluble solid) lead iodide (PbI_2) in the diagram on the next is removed using filtration technique to separate the solid from the solution of sodium nitrate ($NaNO_3$).



Solid lead iodide, PbI_2, is removed using filtration method.

To decide if a precipitate forms, the ions (particles that have charges) involved and the solubility rules must be learnt.

What are solubility rules?

Solubility rules describe which compounds are soluble (dissolves) and which are insoluble (does not dissolve). The solubility rules are used to predict the formation of precipitates when solutions are mixed.

Ion(s) present in compound	Solubility	Exceptions
Sodium (Na^+); Potassium, (K^+);Ammonium(NH_4^+)	All are soluble.	
Nitrate (NO_3^-)	All are soluble.	
Chloride (Cl^-)	All are soluble.	Silver chloride and lead chloride
Sulphate (SO_4^{2-})	All are soluble.	Lead sulphate, calcium sulphate and barium sulphate.
Carbonate (CO_3^{2-})	All are insoluble.	Group I (sodium carbonate, potassium carbonate) and ammonium carbonate
Hydroxide (OH^-)	All are insoluble.	Group I (sodium hydroxide, potassium hydroxide)
Iodide (I^-)	All are soluble.	Silver iodide and lead iodide

Solubility grids are useful for predicting what precipitates (if any) will form when two solutions are mixed.
Only combinations of cations (positive ions) like sodium ions (Na^+) or magnesium ions (Mg^{2+}) with anions (negative ions) like chloride ions (Cl^-) or sulphate ions (SO_4^{2-}) are considered:

√ **represents a combination giving an insoluble compound (precipitate).**
χ **represents a combination giving a soluble compound.**

Combinations of "anion with anion" and "cation with cation" are ignored and marked with a dash (–).

Predicting precipitation
When you mix a solution of calcium chloride ($CaCl_2$) in a solution of sodium carbonate (Na_2CO_3), the calcium ions (Ca^{2+}) and chloride ions (Cl^-) of the first solution are put across the top of the grid. Ions in the second solution, sodium ions (Na^+), and carbonate ions (CO_3^{2-}) are written down the grid.

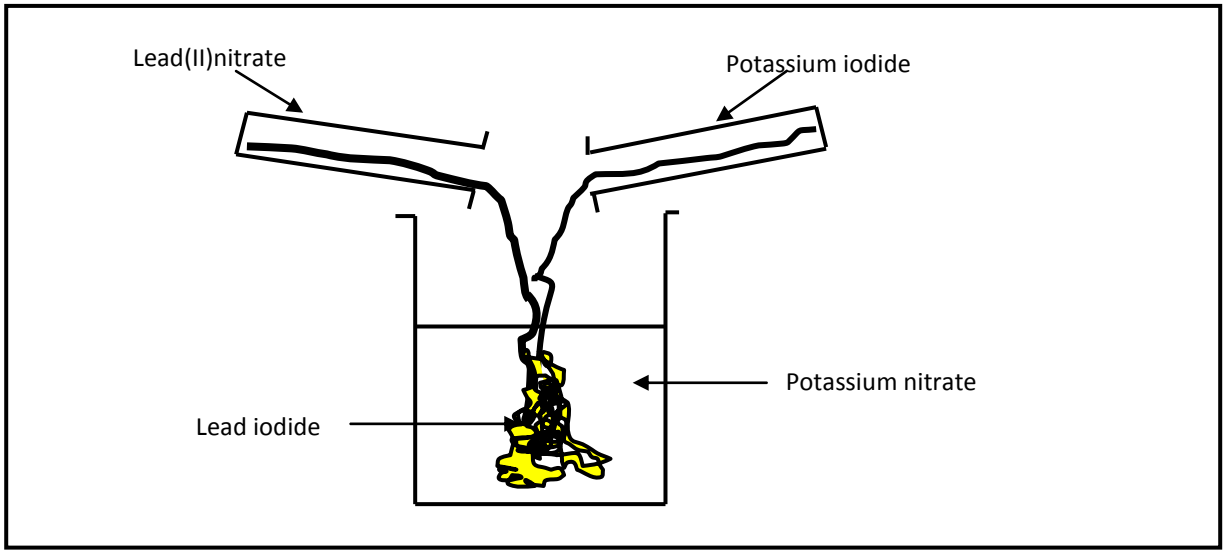
Ions from calcium chloride ($CaCl_2$) solution			
Ions from sodium carbonate (Na_2CO_3) solution		Ca^{2+}	Cl^-
	Na^+	–	X
	CO_3^{2-}	√	–

From the solubility rules, we can predict that a precipitate of calcium carbonate, $CaCO_3$, will form when the two solutions are mixed.
So, a tick (√) is placed on the grid where calcium ions (Ca^{2+}) and carbonate ions (CO_3^{2-}) ions meet. Sodium chloride is soluble, so the sodium ions (Na^+) and chloride ions (Cl^-) will not form a precipitate; hence an **X** is placed on the grid where they meet. Cation to cation and anion to anion combinations are ignored, so given a dash (–).
Precipitation may occur if the concentration of a compound exceeds its solubility (such as when mixing liquids or changing the temperature).

Change of Colour, Gas Given Off, and Disappearance of Solids
Change of colour
In some cases, products formed are of different colour to the reactants and hence, are easily noticed.

For example, if you get a piece of shiny iron nail and leave it for some time, the nail will have a reddish – brown substance. This reddish – brown substance is rust.

Another example, when lead (II) nitrate solution (colourless) and potassium iodide solution (colourless) are mixed, the solution turns yellow.



A chemical reaction showing the change of colour when two aqueous solutions of lead(II)nitrate and potassium iodide are mixed.

This is represented in a chemical equation below:

lead (II) nitrate	+	potassium iodide	→	lead iodide	+	potassium nitrate
$Pb(NO_3)_{2(aq)}$	+	$2KI_{(aq)}$	→	$\mathbf{PbI_{2(aq)}}$	+	$2KNO_{3(aq)}$
colourless		colourless		yellow		colourless

Gas given off

There are many reactions during which gases are evolved. Reactions of some metallic carbonates and acids and also reactions of alkali metals with water lead to the formation of gases. As most of the gases are colourless, observations are made using behaviour of gases. For example, formation of carbon dioxide when calcium carbonate reacts with acids are confirmed by passing the evolved gas through limewater or calcium hydroxide ($Ca(OH)_2$) which turns milky when carbon dioxide (CO_2) is passed through solution.

Formation of Carbon Dioxide

calcium carbonate	+	hydrochloric acid	→	calcium chloride	+	carbon dioxide	+	water
$CaCO_{3(aq)}$	+	$2HCl_{(aq)}$	→	$CaCl_{2(aq)}$	+	$\mathbf{CO_{2(g)}}$	+	$H_2O_{(l)}$

Passing Carbon Dioxide in Limewater, $Ca(OH)_2$

calcium hydroxide	+	carbon dioxide	→	calcium carbonate	+	water
$Ca(OH)_{2(aq)}$	+	$\mathbf{CO_{2(g)}}$	→	$CaCO_{3(s)}$	+	$H_2O_{(l)}$

Also, when you drop a small piece of sodium metal in a test tube of water, you will see bubbles evolved. This is because of the formation of hydrogen gas. Thus, the formation of bubbles is an indication that a reaction has taken place.

sodium iodide	+	water	→	sodium hydroxide	+	hydrogen gas
$2Na_{(s)}$	+	$2H_2O_{(l)}$	→	$2NaOH_{(aq)}$	+	$\mathbf{H_{2(g)}}$

Disappearance of Solids

When some solids react with liquids they change phase from solid to liquid while forming products. This should not be mistaken with physical dissolution of solids in liquids such as dissolving of salt or sugar in water.

Salt or sugar dissolved in water is a physical change where the process is comparatively slow and usually needs constant stirring to facilitate the dissolving process.

For example, when a piece of solid magnesium is dropped in hydrochloric acid solution, magnesium will dissolve and disappear during the reaction. Magnesium chloride solution and hydrogen gas are produced.

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

Learning Activity 1

5 minutes

Answer the following questions:

1. List the indicators of a chemical change.
   (i) _____
   (ii) _____
   (iii) _____
   (iv) _____
   (v) _____

2. When solid zinc is added to dilute sulphuric acid, zinc sulphate solution is produced and hydrogen gas is evolved during the reaction. List down at least two (2) indicators of a chemical change.
   (i) _____
   (ii) _____

3. If a colourless solution of silver nitrate and a colourless solution of sodium chloride are mixed together, a pale yellow precipitate will form. What is the name of this precipitate?
   _____

4. From the list of substances below, predict if they are **soluble** (can dissolved in water) or **insoluble** (cannot dissolve in water and thus forming precipitate) using the solubility grid in the previous page.
   (i) Barium sulphate _____
   (ii) Sodium nitrate _____
   (iii) Calcium carbonate _____

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

11.3.2 Types of Reactions

Several types of chemical reactions can occur based on what happens.
A Periodic Table must be used in order to succeed in studying different types of chemical reactions.

What is a Periodic Table?
A Periodic Table is a table where chemical elements are arranged in order of increasing atomic number. Elements with similar properties are arranged in the same column called **groups**, and the elements with the same number of electron shells are arranged in the same row called **periods**. (Refer to the Periodic Table at the end of the module.)

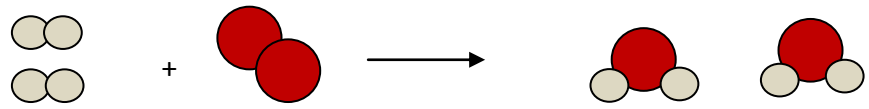
Chemical Change
To form a compound from its elements, a chemical change or reaction must takes place. This is called a **chemical change**. Chemical changes are called **chemical reactions**. In any chemical reaction, the atoms of the different elements combine together to form a new substance called a **compound.**

Diatomic molecules (two atoms chemically combined together) also undergo chemical changes. Some examples of diatomic molecules are oxygen (O_2), hydrogen (H_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

For example, when hydrogen gas is heated with oxygen gas, it will form water as shown in the equation below:

Word Equation: Hydrogen gas + Oxygen gas ⟶ Water

Symbol Equation: $2H_{2(g)}$ + $O_{2(g)}$ ⟶ $2H_2O_{(g)}$



Two hydrogen gas One oxygen gas Two water molecules

2 X 2 = 4 + 2 X 16 = 32 ⟶ 2 X 18 = 36

How did we get the above calculations?
Simply by multiplying the number of atoms to its atomic mass. In the case of Hydrogen gas above;

2 X 2 = 4 — Total number of atoms
Mass number of hydrogen
Number of hydrogen atoms

The same procedure applies to oxygen atom.

Mass can neither be created nor destroyed. In any chemical reaction, mass is always conserved. This is called the Law of Conservation of Mass, which means that the mass of reactants is equal to the mass of the products.

From the above reaction between hydrogen and oxygen to form water, the mass of the reactants is the same as the mass of the products.

Hydrogen gas + Oxygen gas ⟶ Water
$H_{2(g)}$ + $O_{2(g)}$ ⟶ $2H_2O_{(g)}$

2 X 2 = 4 + 2 X 16 = 32 ⟶ 2 X 18 = 36

Total mass of reactants is 36g Total mass of product is 36g

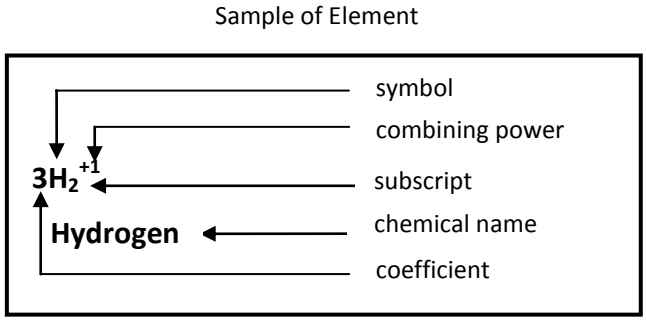
A chemical reaction occurs when a new substance is formed and atoms or ions are rearranged. Mass is conserved in such a reaction.

Chemical Equations

What is a chemical equation?

The rearrangement of atoms during a chemical reaction can be represented by a chemical equation. A **chemical equation** shows the number and types of atoms of the reacting substances, called **reactants**, which are rearranged to form new substances called **products**. A plus (+) sign is used where there is more than one reactant or product. Reactants and products are separated by an arrow (⇻

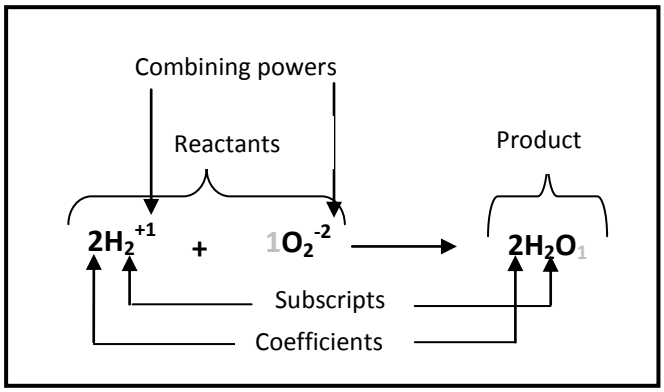
We can write an equation by using **chemical names, symbols** or **chemical formulae**. The most important thing to remember is that an equation must represent a reaction known to take place.



We will use again the reaction when hydrogen gas, H_2 reacts with oxygen gas, O_2 to form water, H_2O to have a better understanding on chemical equations.

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

The chemical equation for the reaction is written as:

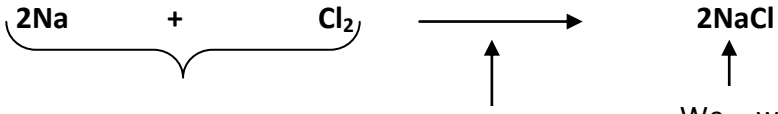


The word equation (equation in words) for the reaction between sodium and chlorine is written below.

Word equation:

sodium + chlorine → sodium chloride

We can also make use of chemical formulas to write the chemical equation.



We write the reacting substances or reactants on the left side of the equation.

The arrow means 'react to form' the reaction proceeds from left to right.

We write the substances or products formed or products on the right side of the equation. Here, the sodium chloride, NaCl, is the only product.

An equation is usually written using chemical formulas, unless you are instructed to write a word equation.

Four Steps in Writing Chemical Equations

1. Write down the symbols or formulae of the reactants and the products.
2. Check the number of atoms of each element on both sides of the equation.
3. Balance the equation by placing coefficient or numbers in front of the formulas of the substances in the equation. The number '1' is not written.
4. Include the state symbols in the equation.

There are many symbols used when writing chemical equations.

- (s) indicates the substance is in the solid state.
- (l) indicates the substance is in liquid state.
- (g) indicates the substance is in gaseous state.
- (aq) indicates the substance is in an aqueous solution (dissolved in water).
- $\xrightarrow{\Delta}$ indicates heat has been added to the reaction.
- MnO_2 indicates a catalyst (in this case, manganese dioxide (MnO_2) has been added to the reaction.

Before we move on in writing chemical equations, let us revise the formula of common ions and **valencies** (charge of the atoms to become ions) of some atoms.

For example:

- Symbol for sodium atom is **Na.**
- Sodium atom is in Group I of the Periodic Table and it has a positive 1 charge, so we write the sodium ions as **Na^+**.
- There are 11 electrons (11 atomic numbers) in sodium atom and its **electron configuration** (electron arrangement) is 2.8.1. Therefore, sodium has a positive 1 charge, because sodium has only **one electron to lose** from its **outer shell or valence shell** (the shell that is furthest from the nucleus) to become a **positive ion**, sodium ions, **Na^+**. All metals form positive ions.
- The same procedure applies to non- metals like chlorine. The symbol for chlorine is **Cl**.
- Chlorine atom is in Group VII of the Periodic Table. It has a negative 1 charge. We write the chlorine ion (or chloride ions) as **Cl^-**.

- There are 17 electrons (17 atomic numbers) in chlorine atom. Its **electron configuration is 2.8.7.**
- Since, chlorine has 7 electrons in its outer shell; it has to **gain 1 electron** to achieve a noble gas structure, in this case, Neon (Ne), having 8 electrons in its outer shell. In doing so, it becomes a **negative ion,** chloride ions, **Cl^-**. All non-metals form negative ions.

Atoms tend to lose or gain electrons to become ions. All metals lose its extra electrons from their outer shell to become a positive ion, called cations. All non-metals gain electron/s to become negatively charged ions called anions.

Example 1:

Making use of the **four steps** given, we can write balanced chemical equation for the reaction between sodium and chlorine to form sodium chloride.

Step 1

Write down the word and symbol equations for the reactants and products.

sodium + chlorine → sodium chloride

Na + Cl_2 → NaCl

Step 2

Check the number of atoms of each element on both sides of the equation. There are **two (2) chlorine atoms** on the left – hand side, but only **one (1) chlorine atom** on the right-hand side. This means that the equation is not balanced.

Na + Cl_2 → NaCl

Step 3

Balance the equation. To balance the number of chlorine atoms, we need to put a number as a **coefficient** (the number added or placed in front of the formula of a substance). The coefficient needed is 2. Therefore, **2 in front of sodium chloride, NaCl** is written. The equation is still not balanced because there are **two sodium atoms on the right** hand **side** and **one sodium atom on the left-hand side**. To balance the equation, we need to put a **'2' in front of 'Na'**.

2Na + Cl_2 → 2NaCl

Step 4

Add the state symbols.

$2Na_{(s)}$ + $Cl_{2(g)}$ → $2NaCl_{(s)}$

The balanced equation is read as **"two atoms of sodium react with one molecule of chlorine to form two molecules of sodium chloride".**

Example 2:

Again, making use of the **four steps** given, we can write the balanced equation for the reaction between methane and oxygen gas.

Step 1

Write down the word equation symbol equation for the reactants and products.

methane	+	oxygen	→	carbon dioxide	+	water vapour
CH_4	+	O_2	→	CO_2	+	H_2O

Step 2

Check the number of atoms of each element on both sides of the equation.

There are **four hydrogen atoms** and **two oxygen atoms** on the left- hand side but only **two hydrogen atoms** and **three oxygen atoms** on the right-hand side. This means that the equation is not balanced.

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

Step 3

Balance the equation. To balance the number of hydrogen, put a number as a **coefficient.** The coefficient needed is **2 in front of water, H_2O,** and **oxygen gas, O_2**. Now, the equation is already balanced.

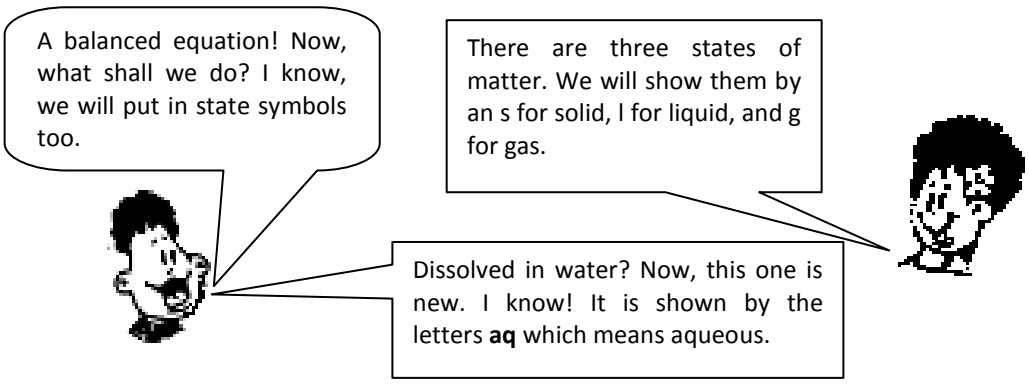
$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Step 4

Add the state symbols. Note the state symbol for H_2O is 'g' because water vapour (a gas) is formed.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

The balanced equation is read as "one molecule of methane reacts with two molecules of oxygen to form one molecule of carbon dioxide gas and two molecules of water vapour ".



Types of chemical reactions

A. Synthesis Reaction

Synthesis reaction can also be called **combination reaction**. It is a type of chemical reaction where two or more simple substances (the reactants) are combined directly or put together to form a single product (the product). This means the formation of compound from the combination of its element.

General form:	**A**	**+**	**B**	**→**	**AB**

An example is the reaction in which sodium (Na) combines with chlorine (Cl_2) to form sodium chloride, or table salt (NaCl).

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

The different chemical reactions involved in the synthesis reactions are the following:

(i) Metal + Oxygen Gas → Metal oxide

When metal reacts with oxygen gas, a **metal oxide** (when metal reacts with oxygen) will form.

Examples:

$2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$

$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$

(ii) Metal + Sulphur → Metal sulphide

When metal reacts with sulphur, a **metal sulphide** (when metal reacts with sulphur) will form.

Examples:

$Zn_{(s)} + S_{(s)} \rightarrow ZnS_{(s)}$

$Mg_{(s)} + S_{(s)} \rightarrow MgS_{(s)}$

$Fe_{(s)} + S_{(s)} \rightarrow FeS_{(s)}$

$Cu_{(s)} + S_{(s)} \rightarrow CuS_{(s)}$

(iii) Metal + Halogen → Metal halide

When metal reacts with halogen, a **metal halide** (when metal reacts with halogen in group VII atoms) will form.

Examples:

$2Na_{(s)} + F_{2(g)} \rightarrow 2NaF_{(s)}$

$2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(s)}$

$2Ag_{(s)} + Br_{2(g)} \rightarrow 2AgBr_{(s)}$

$Mg_{(s)} + I_{2(g)} \rightarrow MgI_{2(s)}$

(iv) Non - metal + Oxygen gas → Non-metal oxide

When non- metal reacts with oxygen gas, a **non- metal oxide** (when a non-metal reacts with oxygen) will form.

Examples:

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

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

$2C_{(s)} + O_{2(g)} \rightarrow 2CO_{(g)}$

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

Learning Activity 2

minutes

Answer the following questions:

1. Balance the following synthesis reactions.

   a. $Na_{(s)} + O_{2(g)} \rightarrow Na_2O_{(s)}$

   b. $Al_{(s)} + S_{(s)} \rightarrow Al_2S_{3(s)}$

   c. $K_{(s)} + F_{2(g)} \rightarrow KF_{(s)}$

   d. $C_{(s)} + CO_{2(g)} \rightarrow CO_{(g)}$

   e. $Cu_{(s)} + O_{2(g)} \rightarrow CuO_{(s)}$

   f. $Ag_{(s)} + Cl_{2(g)} \rightarrow AgCl_{(s)}$

   g. $K_{(s)} + Br_{2(g)} \rightarrow KBr_{(s)}$

   h. $H_{2(g)} + Cl_{2(g)} \rightarrow HCl_{(g)}$

   i. $N_{2(g)} + H_{2(g)} \rightarrow NH_{3(g)}$

2. Name the products of the following synthesis reaction:

   (i) A reaction between iron and sulphur will form:
   _____

   (ii) A lithium is heated with chlorine will form
   _____

   (iii) A calcium oxide is reacted with carbon dioxide will produce:
   _____

   (iv) A sulphur is heated strongly with oxygen in the air is producing
   _____

   (v) When magnesium is heated with oxygen, a white solid is formed called
   _____

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

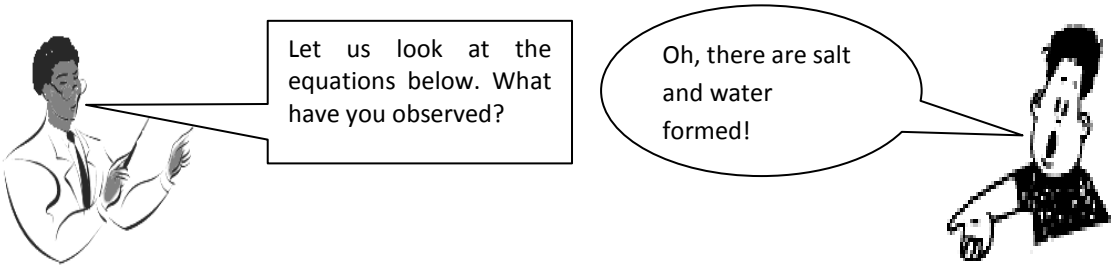
B. Neutralization Reaction

Have you ever had indigestion? The burning feeling comes from too much hydrochloric acid in your stomach. You can cure the pain quickly by taking a tablet. The tablet contains an alkali which gets rid of the acid. Acids and alkali are chemical opposites. They react together and 'cancel each other out '. If we mix just the right amount of acid and alkali together, we get a neutral solution.

The acid-base reactions take place when hydrogen ions (H^+) and hydroxide ions (OH^-) combine to form water. Acids react with bases so that the properties of both are lost to form water and a salt. The reaction between an acid and an alkali (base) is called **neutralization**.

The reaction between an acid and alkali (base) is called neutralization.

General Equation: **Acid + Alkali → Salt + Water**



(i) hydrochloric acid + sodium hydroxide → sodium chloride + water

$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$

(ii) nitric acid + sodium hydroxide → sodium nitrate + water

$HNO_{3(aq)} + NaOH_{(aq)} \rightarrow NaNO_{3(aq)} + H_2O_{(l)}$

(iii) sulphuric acid + sodium hydroxide → sodium sulphate + water

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

What is a salt?

You know that sodium chloride is a salt. Chemists use the word salt to describe any metal compounds that are made from acids.

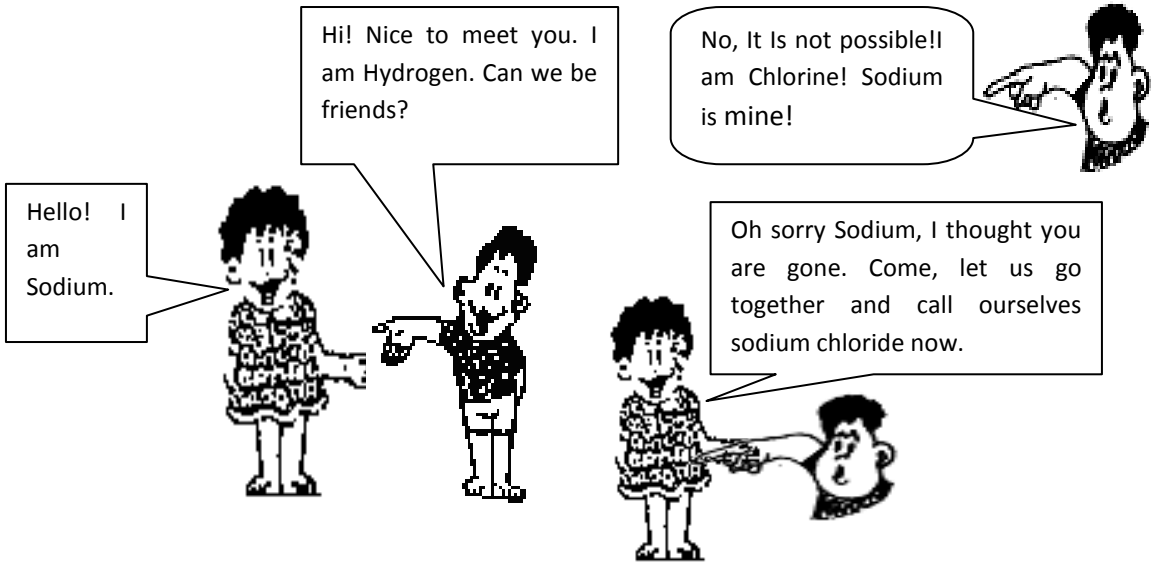
Can you remember the formulas of strong acids in Grade 10?

These strong acids are commonly found in school laboratories such as hydrochloric acid (HCl), sulphuric acid (H_2SO_4) and nitric acid (HNO_3). We find that all acids contain hydrogen.

Naming salts

Each acid has its own salts. Look at this table:

Acid	its salts	Example
Hydro**chlo**ric acid	→ **chlorides**	Sodium chloride, NaCl
Sulphuric acid	→ **sulphates**	Copper sulphate, $CuSO_4$
Nitric acid	→ **nitrates**	Potassium nitrate, KNO_3



The cartoons above show the displacement of hydrogen by sodium. Hydrogen in an acid is replaced by metal sodium and a salt, sodium chloride is formed.

We will discuss some bases like metal oxides and **metal carbonate** (when a metal reacts with carbonate ions (CO_3^{2-}) as you progress.

Can you give the difference between an alkali and a base? An **alkali** is just a base that can dissolve in water and a base will not dissolve in water. Alkalis are part of a larger group of compounds called **bases**.

Do you think bases will react with acids, as alkalis do?

For example,

sulphuric acid + zinc oxide → zinc sulphate + water

$H_2SO_{4(aq)}$ + $ZnO_{(s)}$ → $ZnSO_{4(aq)}$ + $H_2O_{(l)}$

Also, when nitric acid reacts with magnesium oxide, it forms a **salt** and **water**. As in neutralization, what is the name of the salt that will form in the reaction between nitric acid and magnesium oxide?

nitric acid + magnesium oxide → magnesium nitrate + water

$2HNO_{3(aq)}$ + $MgO_{(s)}$ → $Mg(NO_3)_{2(aq)}$ + $H_2O_{(l)}$

The **different chemical reactions** involved in neutralization reactions are the following:

(i) Acid + Aqueous ammonia solution → Salt + Water

When an acid reacts with aqueous ammonia solution, it forms salt and water.

Examples: $HCl_{(aq)} + NH_4OH_{(aq)} \rightarrow NH_4Cl_{(aq)} + H_2O_{(l)}$

From the above balanced chemical equation, the **ionic equation** is;

$$\mathbf{H^+_{(aq)}} + \cancel{Cl^-_{(aq)}} + \cancel{NH_4^+{}_{(aq)}} + \mathbf{OH^-_{(aq)}} \rightarrow \cancel{NH_4^+{}_{(aq)}} + \cancel{Cl^-_{(aq)}} + \mathbf{H_2O_{(l)}}$$

The chloride ions ($Cl^-_{(aq)}$) and ammonium ions ($NH_4^+{}_{(aq)}$) stay in the solution. They are not changed before and after a chemical reaction. They do not appear in the ionic equation. They are called **spectator ions** (the ions that are struck out from the equation).

Spectator ions do not take part in a chemical reaction.

Therefore, the **net ionic equation** is;

$$\mathbf{H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}}$$

The net ionic equation shows only ions, which are affected in the reaction.

(ii) Acid + Metal carbonate → Salt + Water + Carbon dioxide

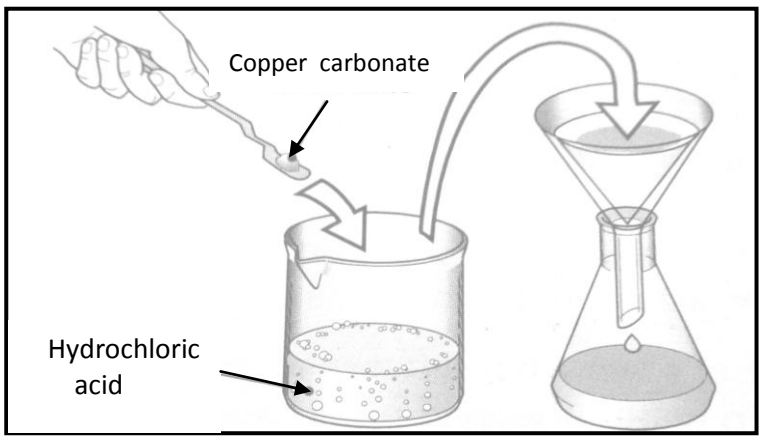
When an acid reacts with metal carbonates it forms salt, carbon dioxide, and water.

For example:

Hydrochloric acid + copper carbonate → copper chloride + water + carbon dioxide

$2HCl_{(aq)} + CuCO_{3(s)} \rightarrow CuCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

How is the equation between an acid and a metal carbonate different from the equation for an acid reacting with a base? How would you test for the gas given off?



Hydrochloric acid reacting with copper ccarbonate

From the balanced chemical equation, the **ionic equation** is;

$$\mathbf{2H^+_{(aq)}} + 2Cl^-_{(aq)} + Cu^{2+}_{(aq)} + \mathbf{CO_3^{2-}_{(aq)}} \rightarrow Cu^{2+}_{(aq)} + 2Cl^-_{(aq)} + \mathbf{H_2O_{(l)}} + \mathbf{CO_{2\,(g)}}$$

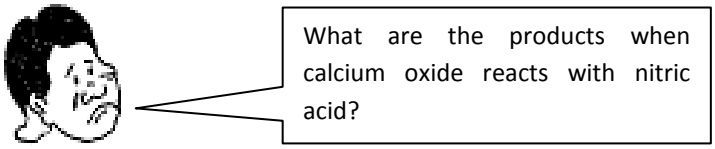
The chloride ions ($Cl^-_{(aq)}$) and copper ions ($Cu^{2+}_{(aq)}$) stay in the solution. They are the spectator ions from the reaction. Remember the meaning of spectator ions?
Therefore, the **net ionic equation** is;

$$\mathbf{2H^+_{(aq)} + CO_3^{2-}_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}}$$

(iii) Acids + Metal oxide → Salt + Water

When an acid reacts with a metal oxide, it forms salt and water. Remember the general equation**:**

Acid + base → salt + water?



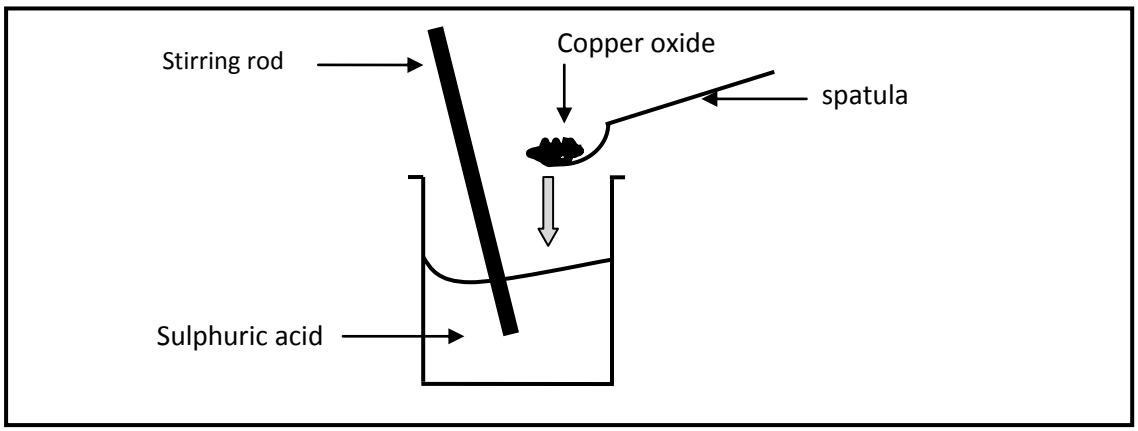
Calcium (CaO), can react with nitric acid (HNO_3), forming calcium nitrate, ($Ca(NO_3)_2$) and water as in equation below:

$$CaO_{(s)} + 2HNO_{3\,(aq)} \rightarrow Ca(NO_3)_{2(aq)} + H_2O_{(l)}$$

Let us take a look at the other example given below:

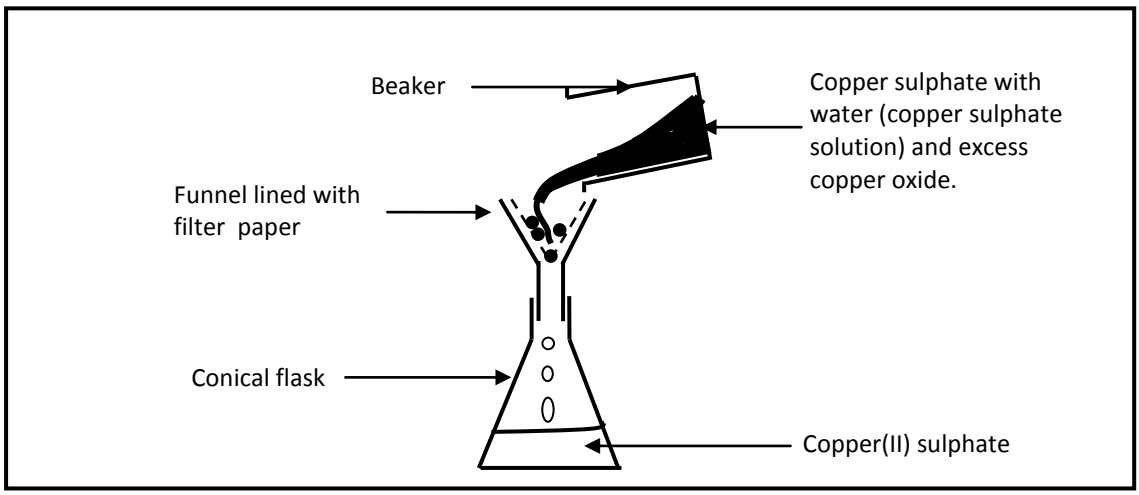
Copper (II) oxide reacts with sulphuric acid to produce copper(II)sulphate and water.

1. The copper oxide is added to 25cm^3 of sulphuric acid solution and stirred with a glass rod. Keep adding copper oxide until no more dissolve (which means that the acid has used up all the copper oxide).



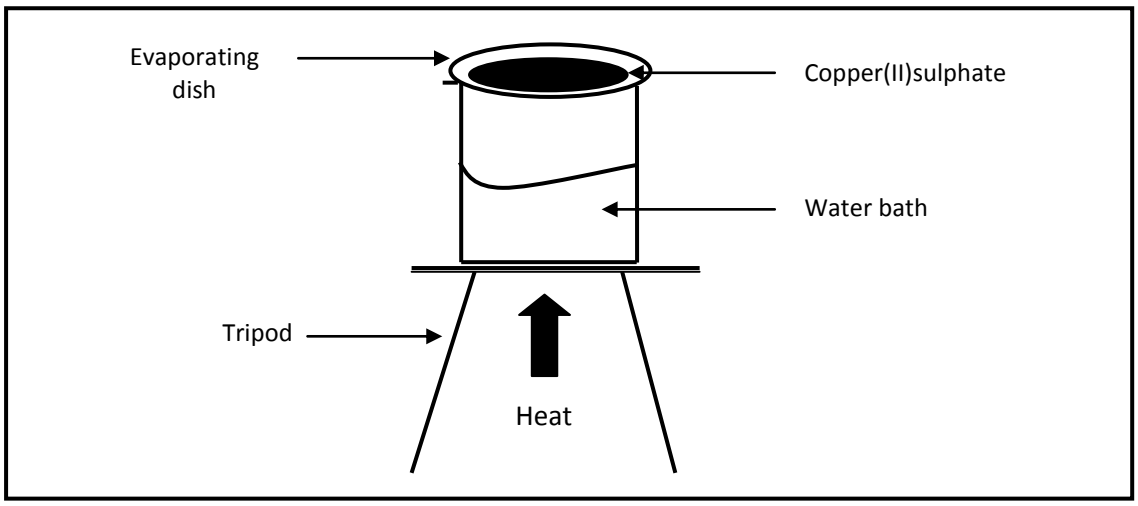
The reaction between copper oxide and sulphuric acid

2. Filter the mixture and you should have a blue solution of copper(II)sulphate.



Filtration of copper(II) sulphate solution with excess copper oxide

3. The solution is then heated on a water bath (to control the temperature). Stop heating when crystals of copper(II)sulphate are seen.



Evaporation of copper(II) sulphate using water bath

The complete balanced equation for the reaction is;

sulphuric acid + copper (II) oxide → copper(II)sulphate + water

$H_2SO_{4(aq)} + CuO_{(s)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

The ionic equation would be;

$\mathbf{2H^+_{(aq)}} + SO_4^{2-}{}_{(aq)} + Cu^{2+}{}_{(aq)} + \mathbf{O^{2-}_{(aq)}} \rightarrow Cu^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} + \mathbf{H_2O_{(l)}}$

Therefore, the net ionic equation is;

$\mathbf{2H^+_{(aq)} + O^{2-}_{(aq)} \rightarrow H_2O_{(l)}}$

Can you point out the spectator ions from the above reactions?

(iv) Acid + Metal hydroxide → Salt + Water

A typical example of this type of chemical reaction is when an **acid** reacts with an **alkali**, forming a **salt** and **water** as shown already.

The balanced chemical equation for the neutralization reaction between hydrochloric acid and potassium hydroxide is given below. It also shows the ionic equation and net ionic equation.

Complete balanced equation:

hydrochloric acid + potassium hydroxide → potassium chloride + water

$HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

Ionic equation:

$\mathbf{H^+_{(aq)}} + Cl^-_{(aq)} + K^+_{(aq)} + \mathbf{OH^-_{(aq)}} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + \mathbf{H_2O_{(l)}}$

Net Ionic equation:

$\mathbf{H^+_{(aq)}} + \mathbf{OH^-_{(aq)}} \rightarrow \mathbf{H_2O_{(l)}}$

Therefore, the spectator ions are chloride ion, $Cl^-_{(aq)}$ and potassium ion, $K^+_{(aq)}$ because they remain unchanged before and after the chemical reaction.

(v) Acid + Metal hydrogen carbonate → Salt + carbon dioxide + Water

The hydrogen carbonate ion has the formula, $HCO_3^-{}_{(aq)}$. Suppose, nitric acid reacts with sodium hydrogen carbonate, sodium nitrate, carbon dioxide and water are the products.

Complete balanced equation:

Nitric acid + sodium hydrogen carbonate → sodium nitrate + carbon dioxide + water

$HNO_{3(aq)} + NaHCO_{3(s)} \rightarrow NaNO_{3(aq)} + \mathbf{CO_{2\,(g)}} + H_2O_{(l)}$

Ionic equation:

$\mathbf{H^+_{(aq)}} + \cancel{NO_3^-{}_{(aq)}} + \cancel{Na^+_{(aq)}} + \mathbf{HCO_3^-{}_{(aq)}} \rightarrow \cancel{Na^+_{(aq)}} + \cancel{NO_3^-{}_{(aq)}} + \mathbf{CO_{2(g)}} + \mathbf{H_2O_{(l)}}$

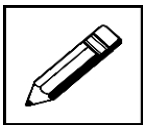
Net Ionic equation:

$\mathbf{H^+_{(aq)}} + \mathbf{HCO_3^-{}_{(aq)}} \rightarrow \mathbf{CO_{2\,(g)}} + \mathbf{H_2O_{(l)}}$

Which are the spectator ions in the above reaction? Can you write their formulas?

Hydrochloric acid and magnesium hydrogen carbonate react almost the same as nitric acid and sodium hydrogen carbonate above to give magnesium chloride, carbon dioxide, and water. Now, try to write your own chemical equation.

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



Learning Activity 3

60 minutes

Answer the following questions:

1. Write the name and the formula of the salt produced by the reaction between the two (2) aqueous solutions given below:
   (i) Hydrochloric acid with sodium sulphate
   Name
   _____
   Formula
   _____
   (ii) Nitric acid with copper chloride
   Name
   _____
   Formula
   _____
   (iii) Sulphuric acid with barium nitrate
   Name
   _____
   Formula
   _____

2. Write a balanced equation for the reaction between the following solutions below:

   (i) $HCl + MgO \rightarrow$ _____ + _____
   (ii) $HNO_3 + Ca(OH)_2 \rightarrow$ _____ + _____
   (iii) $H_2SO_4 + KOH \rightarrow$ _____ + _____

3. From the following neutralization reaction, write
   a. a balanced equation
   b. an ionic equation and
   c. a net ionic equation for the reaction not involving the spectator ions.

   (i) $HNO_3 + NH_4OH \rightarrow$
   a. _____
   b. _____
   c. _____

(ii) H_2SO_4 + Na_2CO_3 $\rightarrow$

a. _____

b. _____

c. _____

(iii) HCl + CaO $\rightarrow$

a. _____

b. _____

c. _____

(iv) HNO_3 + NaOH $\rightarrow$

a. _____

b. _____

c. _____

(v) H_2SO_4 + $KHCO_3$ $\rightarrow$

a. _____

b. _____

c. _____

4. The table below shows the reactions of different acids and bases. Fill in with correct name and formula of the salt as the product of neutralization reaction

	Base	**Acid**	**Name of Salt**	**Formula of Salt**
(i)	Sodium hydroxide	Hydrochloric acid		
(ii)	Ammonium hydroxide	Hydrochloric acid		
(iii)	Iron (II) oxide	Sulphuric acid		
(iv)	Copper (II) oxide	Sulphuric acid		
(v)	Calcium hydroxide	Nitric acid		
(vi)	Zinc oxide	Nitric acid		

5. Define alkalis and a base? Give at least two (2) examples of each written in their formulas.

Alkalis _____

Examples _____ and _____

Base _____

Examples _____ and _____

6. What is a neutralization reaction?

_____

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

C. Displacement Reaction

There are two general types of displacement reaction:

- Single displacement reaction
- Double displacement reaction

1. Single Displacement Reaction

Most reactive metals can displace (remove) less reactive metals from their salt solutions. If a reactive element comes into contact with the compound of a less reactive element, a chemical reaction may take place. The less reactive element is **removed** from the compound and **replaced** by the more reactive element.

For example, if chlorine is added to a solution of potassium bromide, the bromine is replaced by the chlorine forming potassium chloride. Bromine is formed at the same time and can be detected by its colour. Reactions of this type are called **displacement reactions**.

This can be represented by means of the chemical equation below:

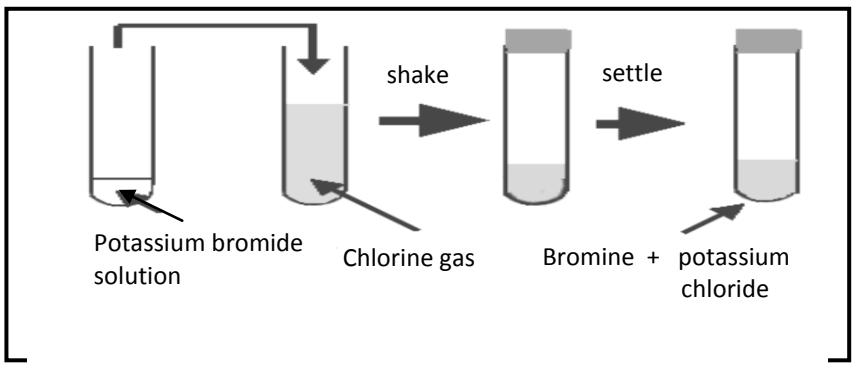
potassium bromide + chlorine gas → potassium chloride + bromine gas

$2KBr_{(aq)} + Cl_{2(g)} \rightarrow 2KCl_{(aq)} + Br_{2(g)}$

↓ ↓

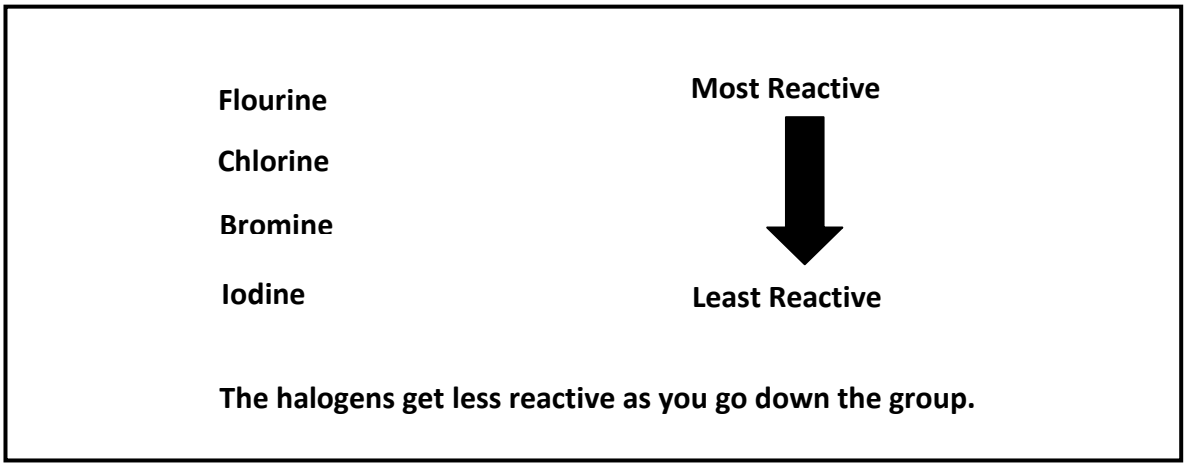
less reactive halogen **more reactive halogen**

Chlorine is more reactive than iodine or bromine. It can 'push' (displace) bromine out of solution.



Bromine is more reactive than iodine. Can it displace iodine from solution? Can you write a word and symbol equation for the reaction between potassium bromide and chlorine gas?
To understand fully the concept of a single displacement reaction, we must understand the reactivity series of metals and halogens.

The **Reactivity Series** is like a 'league table' for metals. The more reactive metals are at the top of the league. The least reactive ones are at the bottom. We will use the Reactivity Series later in our study.



The reactivity series of halogens

The **Reactivity Series of Metals** on the next page shows a list of substances ranked in order of relative reactivity. For example, magnesium metal can knock hydrogen ions out of solution. It is considered more reactive than element hydrogen.

Look at this equation:

magnesium + nitric acid → magnesium nitrate + hydrogen gas

$Mg_{(s)} + 2HNO_{3(aq)} \rightarrow Mg(NO_3)_{2(aq)} + H_{2(g)}$

The ionic equation is,

$Mg_{(s)} + 2H^{+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_{2(g)}$

Zinc can also displace hydrogen ions from solution:

$Zn_{(s)} + 2H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$

Zinc is also **more reactive** than **hydrogen**. Magnesium metal, however, can remove zinc ions from solution:

$Mg_{(s)} + Zn^{2+}_{(aq)} \rightarrow Zn_{(s)} + Mg^{2+}_{(aq)}$

Magnesium is **more reactive** than **zinc**, and the reactivity series including these elements would be Mg > Zn > H. The most active metals are at the top of the table; the least active are at the bottom. Each metal is able to displace the elements below it from solution.

This reactivity series is a useful guide for guessing the products of metal displacement reactions. For example, placing a strip of zinc metal in copper sulphate solution produces metallic copper and zinc sulphate, since zinc is above copper on the series.

A strip of copper placed into a zinc sulphate solution will not produce an appreciable reaction, because **copper** is **below zinc** on the series and cannot displace zinc ions from solution.

METALS	METAL IONS	REACTIVITY
K	K^+	Reacts with water
Na	Na^+	
Li	Li^+	
Ba	Ba^{2+}	
Sr	Sr^{2+}	
Ca	Ca^{2+}	
Mg	Mg^{2+}	Reacts with acids
Al	Al^{3+}	
Mn	Mn^{2+}	
Zn	Zn^{2+}	
Cr	Cr^{2+}	
Fe	Fe^{2+}	
Cd	Cd^{2+}	
Co	Co^{2+}	
Ni	Ni^{2+}	
Sn	Sn^{2+}	
Pb	Pb^{2+}	
H_2	H^+	Included for comparison
Sb	Sb^{2+}	Highly unreactive
Bi	Bi^{2+}	
Cu	Cu^{2+}	
Hg	Hg^{2+}	
Ag	Ag^+	
Au	Au	
Pt	Pt^+	

The reactivity series of some common metals

Hydrogen in Reactivity Series

We can use displacement reactions to give hydrogen a place in the Reactivity Series. Metals above hydrogen in the Reactivity Series will displace hydrogen from its solution.

In your Grade 10 Unit on Chemistry, you learnt that Group I metals are very reactive. That is the reason they are at the top of reactivity series. Potassium, sodium, lithium and calcium can react with cold water vigorously. What do you think will be their reaction with acids?

Single displacement reactions are a type of **redox reaction (reduction –oxidation reaction)**. In reactions, the element being **displaced** is always **reduced**.

Definition of oxidation and reduction

Oxidation can be :	Reduction can be:	Examples
loss of electrons	gains of electrons	$Mg^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow MgCl_{2(s)}$ When magnesium loses 2 electrons to chlorine, an ionic solid magnesium chloride ($MgCl_2$) is formed. Magnesium is oxidized and chlorine is reduced because it has gained electrons from magnesium.
gain of oxygen	loss of oxygen	$2Fe_2O_{3\,(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + 3CO_{2(g)}$ Iron (III) oxide (Fe_2O_3) is reduced, since it loses oxygen and carbon (C) is oxidized, since it gains oxygen to form carbon dioxide (CO_2).
loss of hydrogen	gain of hydrogen	$2H_2S_{(g)} + O_{2(g)} \rightarrow 2S_{(s)} + 2H_2O_{(l)}$ Hydrogen sulphide (H_2S) is oxidized, because hydrogen is lost as hydrogen forms sulphur(S). Oxygen (O_2) is reduced, because hydrogen is gained as oxygen gas (O_2) forms water, (H_2O).
increased in oxidation number	decreased in oxidation	$\overset{0}{K_{(s)}} + 2\overset{+1\ -1}{HCl_{(aq)}} \rightarrow \overset{+1\ -1}{KCl_{(aq)}} + \overset{0}{H_{2(g)}}$ Potassium (K) is oxidized, since its oxidation number (O.N.) increases from zero to one. Hydrogen ion (H^+) is reduced, since its oxidation number decreases from one to zero.

Oxidation numbers are assigned by a set of rules:

RULE	EXAMPLES
1. When an atom exists as elements, they have an oxidation number of zero.	Sodium (Na), chlorine molecule (Cl_2), neon (Ne), carbon (C) and hydrogen molecule (H_2) all have an oxidation number of zero.
2. The oxidation of monoatomic (one atom) ion is the same as the charge on the ion.	Copper ions (Cu^{2+}) has oxidation number of +2, chloride (Cl^-) has an oxidation number of -1.
3. Hydrogen in compounds has an oxidation number of +1, except in metal hydrides, where it is -1.	Hydrogen in water (H_2O), methane (CH_4) and ammonia (NH_3) have an oxidation number of +1. In sodium hydride (NaH), a metal hydride, the oxidation number of hydrogen is -1.
4. Oxygen in compounds has an oxidation number of -2, except in hydrogen peroxide, H_2O_2, where its oxidation number is -1.	Oxygen in magnesium oxide (MgO), sulphuric acid (H_2SO_4), water (H_2O) and potassium permanganate ($KMnO_4$) have an oxidation number of -2. In hydrogen peroxide (H_2O_2), the oxidation number for oxygen is -1.
5. Polyatomic ions (ions containing more than one atom), the sum of the oxidation numbers equals the charge of the ion.	For ammonium ion (NH_4^+), the sum of the oxidation numbers is +1. For sulphate ions (SO_4^{2-}), the sum of the oxidation numbers is -2.
6. The sum of the oxidation numbers of atoms in a molecule is zero.	The sum of the oxidation numbers for the atoms in each of sulphuric acid (H_2SO_4), butane (C_4H_{10}), and water (H_2O) is zero.

Chemical Reactions Involved in Single Displacement Reaction

A. Single Displacement reaction with metal and hydrogen from acid solution

General Equation Form: A + BC → AC + B

Metal A has **displaced B** in the **compound BC** to form a **new compound AC** and the **molecule or gas B**. Remember that B and A are both cations (positively – charged ions) in the example. Element A forms a cation and is a **metal**.

In order for single displacement reaction to happen, it must be more reactive than the metal it is set to displace (replace). In order to determine whether this is the case, we must consult the reactivity series, which ranks metals in order of relative reactivity.

For example, zinc reacts with nitric acid to produce zinc nitrate and hydrogen gas as shown in a chemical equation below:

Zinc + nitric acid → zinc nitrate + hydrogen gas

$Zn_{(s)} + 2HNO_{3\,(aq)} \rightarrow Zn(NO_3)_{2(aq)} + H_{2\,(g)}$

A B C A C B

Hydrogen ion (H^+) from nitric acid (HNO_3) is reduced to hydrogen gas (H_2)

$Zn_{(s)} + 2HNO_{3(aq)} \rightarrow Zn\,(NO_3)_{2(aq)} + H_{2(g)}$

Zinc (Zn) metal is oxidized to zinc ions (Zn^{2+}) in zinc nitrate ($Zn(NO_3)_2$)

The ionic equation for the reaction is:

$Zn_{(s)} + 2H^+_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + H_{2(g)}$

A spectator ion, $NO_3^-{}_{(aq)}$, is not included in the equation.

$Zn_{(s)} + 2H^+_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + H_{2(g)}$

Therefore, the net ionic equation is,

$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$

B. Single Displacement reactions with halogens and metal halides

General Equation Form:	**A**	**+**	**BC**	**→**	**BA**	**+**	**C**

Metal A has **displaced C** in the **compound BC** to form a **new compound BA** and the **molecule or gas C**. Remember that A and C are both anions (negatively – charged ions) in the example. Element A forms an anion and is a non-metal, most frequently with halogens, like fluorine, chlorine, bromine, and iodine.

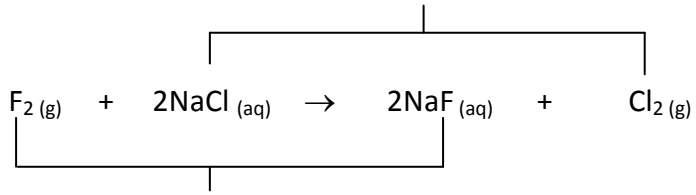
For example, fluorine is heated with sodium chloride solution to produce sodium flouride solution and chlorine gas as shown in a chemical equation below:

fluorine molecule + sodium chloride → sodium fluoride + chlorine molecule

$F_{2(g)} + 2NaCl_{(aq)} \rightarrow 2NaF_{(aq)} + Cl_{2(g)}$

A B C B A C

Chlorine ion (Cl^-) in sodium chloride (NaCl) is reduced to chlorine gas (Cl_2)



$$F_{2\,(g)} + 2NaCl_{\,(aq)} \rightarrow 2NaF_{\,(aq)} + Cl_{2\,(g)}$$

Flourine gas (F_2) is oxidized to fluorine ion (F^-) in sodium fluoride (NaF_2)

The ionic equation for the reaction is not involving the spectator ions.

$$F_{2(g)} + 2Na^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow 2Na^+_{(aq)} + 2F^-_{(aq)} + Cl_{2(g)}$$

Therefore, the net ionic equation is,

$$F_{2(g)} + 2Cl^-_{(aq)} \rightarrow 2F^-_{(aq)} + Cl_{2(g)}$$

Sodium ions ($Na^+_{(aq)}$) are the spectator ions in the reactions.

C. Single Displacement reactions with metals and aqueous salts

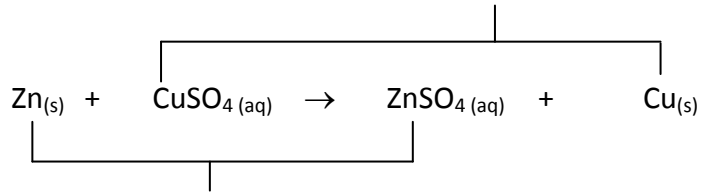
General Equation Form:	**A**	**+**	**BC**	**→**	**AC**	**+**	**B**

Metal A has **displaced B** in the **compound BC** to form a **new compound AC** and the **free metal B**. Remember that **metals A** and **B** are both cations (positively – charged ions) .

For example, zinc reacts with nitric acid to produce zinc nitrate and hydrogen gas as shown in a chemical equation below:

Zinc	+	copper(II) sulphate	→	zinc sulphate	+	copper
$Zn_{(s)}$	+	$CuSO_{4\,(aq)}$	→	$ZnSO_{4\,(aq)}$	+	$Cu_{(s)}$
↓		↓↓		↓↓		↓
A		**B C**		**A C**		**B**

Copper ion (Cu^{2+}) in copper sulphate ($CuSO_4$) is reduced to solid copper (Cu)



$$Zn_{(s)} + CuSO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + Cu_{(s)}$$

Zinc (Zn) metal is oxidized to zinc ions (Zn^{2+}) in zinc sulphate ($ZnSO_4$)

The ionic equation for the reaction not involving the spectator ions is

$$Zn_{(s)} + Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Cu_{(s)}$$

Therefore, the net ionic equation is:

$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

Sulphate ions ($SO_4^{2-}{}_{(aq)}$) are the spectator ions in the reactions.

After learning single displacement reaction, now we will analyze two experiments about single displacement:

- **Experiment 1: Reaction of Acid with Metal**
- **Experiment 2: Reaction of Acid with Metal Carbonate Solution**

Experiment 1: Reaction of Acid with Metal

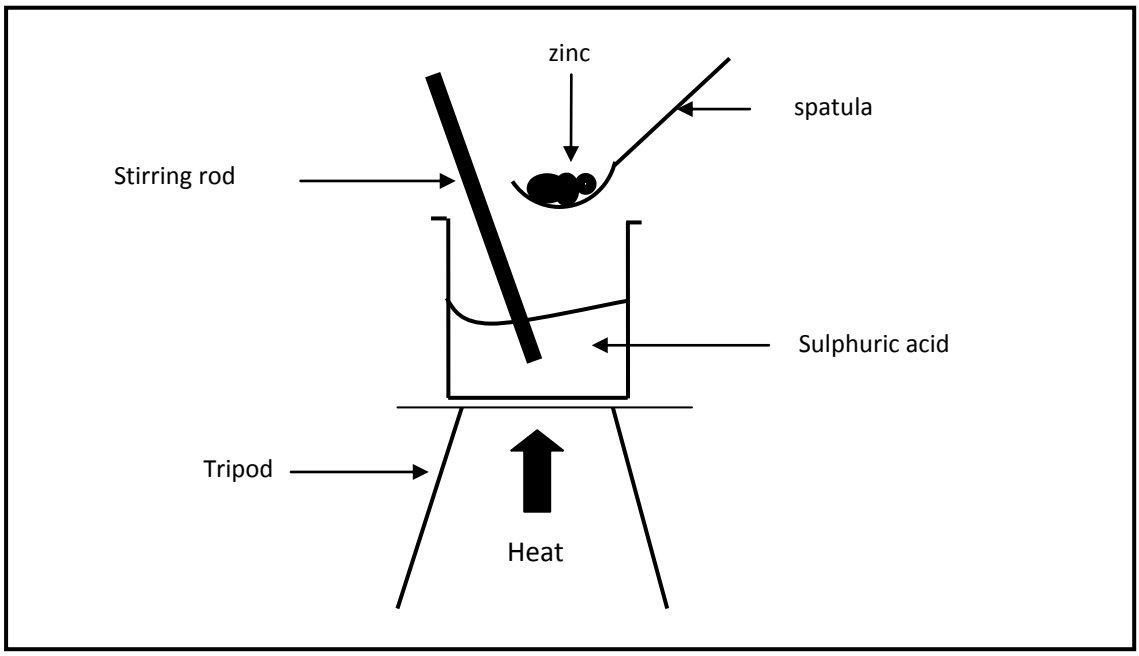
Aim:

To prepare zinc crystals from zinc and sulphuric acid.

Procedure:

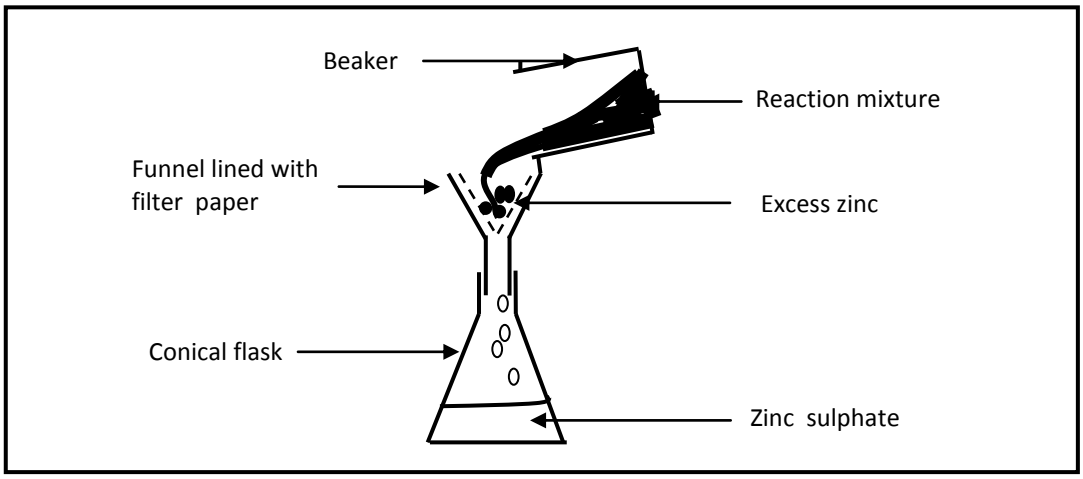
1. Warm 50cm^3 of dilute sulphuric acid in a beaker. Avoid boiling the acid or it will dry too quickly.

2. Add a few pieces of solid zinc to the acid, stirring constantly until no more zinc will dissolve in the acid.



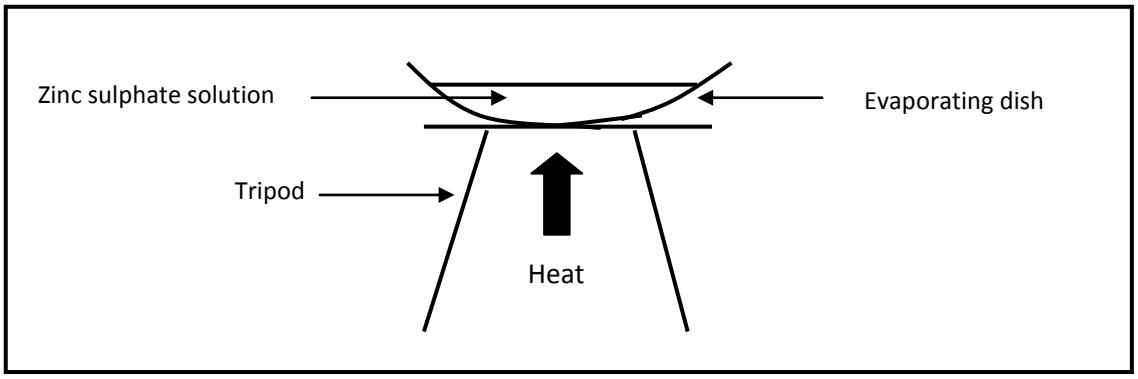
Heating of sulphuric acid with solid zinc.

3. Filter to remove the excess (unreacted) zinc.



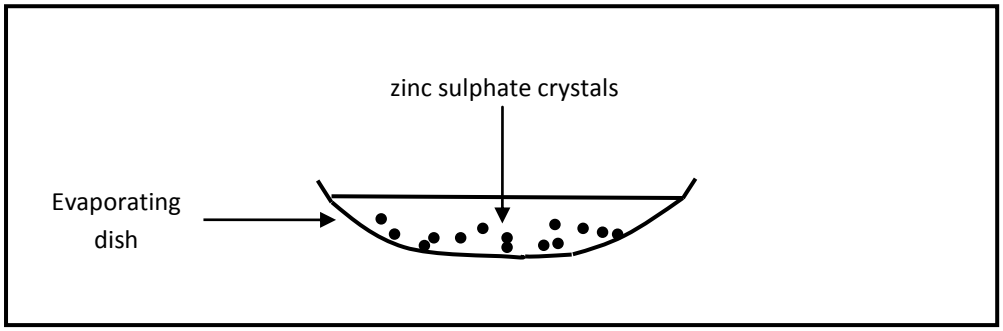
Filtration process to remove excess zinc

4. Collect the filtrate in an evaporating dish and heat it gently to evaporate the excess water.



Heating of zinc sulphate solution

5. Stop heating when crystals start to appear. Allow the solution (filtrate) to cool.



Crystals of zinc sulphate produce from solution

The equation for the above experiment is

Zinc + sulphuric acid $\rightarrow$ zinc sulphate + hydrogen gas

$$Zn_{(s)} + H_2SO_{4\,(aq)} \rightarrow ZnSO_{4\,(aq)} + H_{2\,(g)}$$

Ionic equation:

$$Zn_{(s)} + 2H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + Zn^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + H_{2(g)}$$

Net ionic equation:

$$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

Observation from Experiment 1:

(i) When you add the zinc to the acid, bubbles of hydrogen gas is evolved. The acid has used up the zinc until no more zinc will dissolve at that temperature.

(ii) The filtrate, zinc sulphate has evaporated to dryness to get solid crystals of zinc sulphate.

Experiment 2: Reaction of Acid and a Metal Carbonate Solution

Aim:

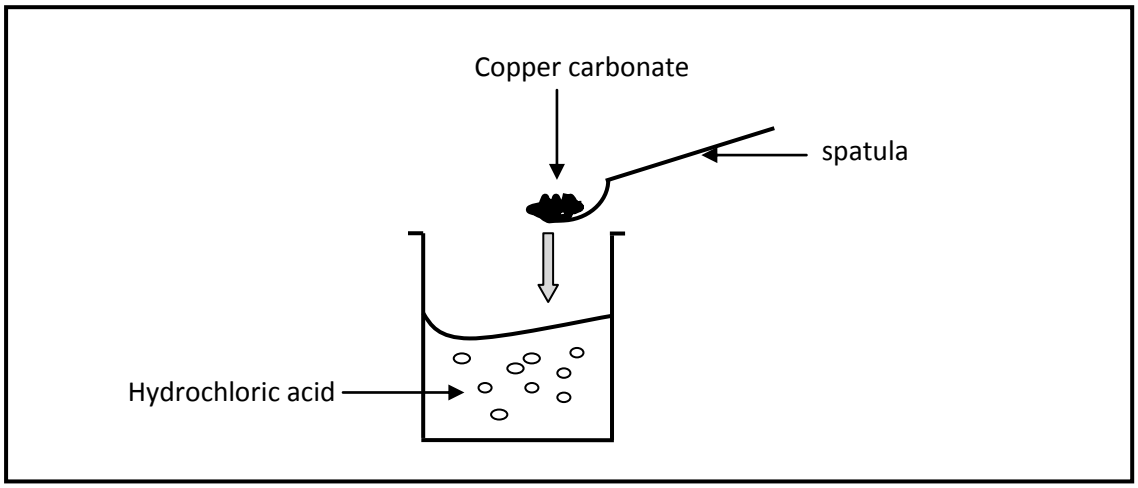
To prepare a sample of copper (II) carbonate by the precipitation method.

Procedures:

1. Collect 25cm^3 of hydrochloric acid in a small beaker.

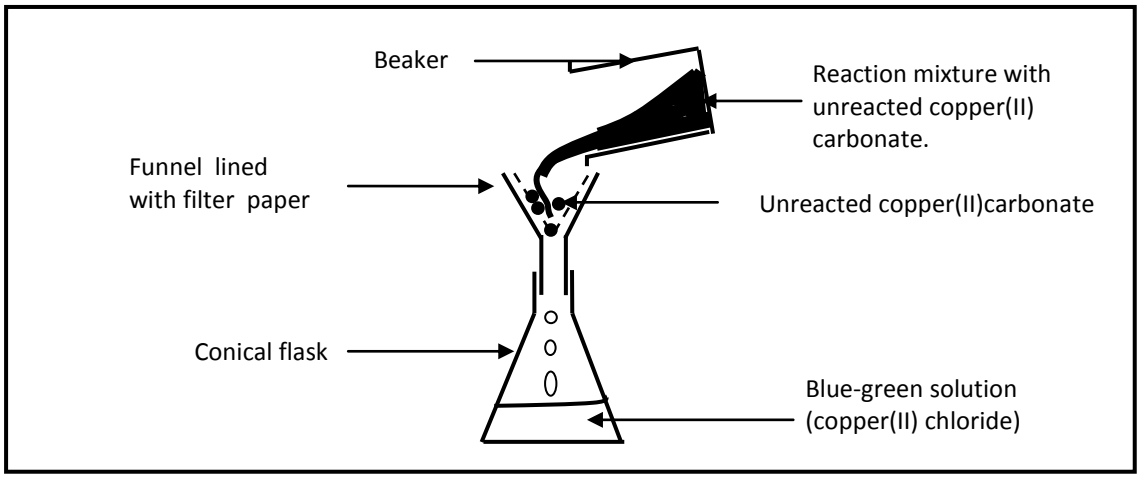
2. Add a spatula of copper(II)carbonate.
   - What happens?
   - What is the gas given off?

3. Add some more copper(II)carbonate until it stops fizzing.



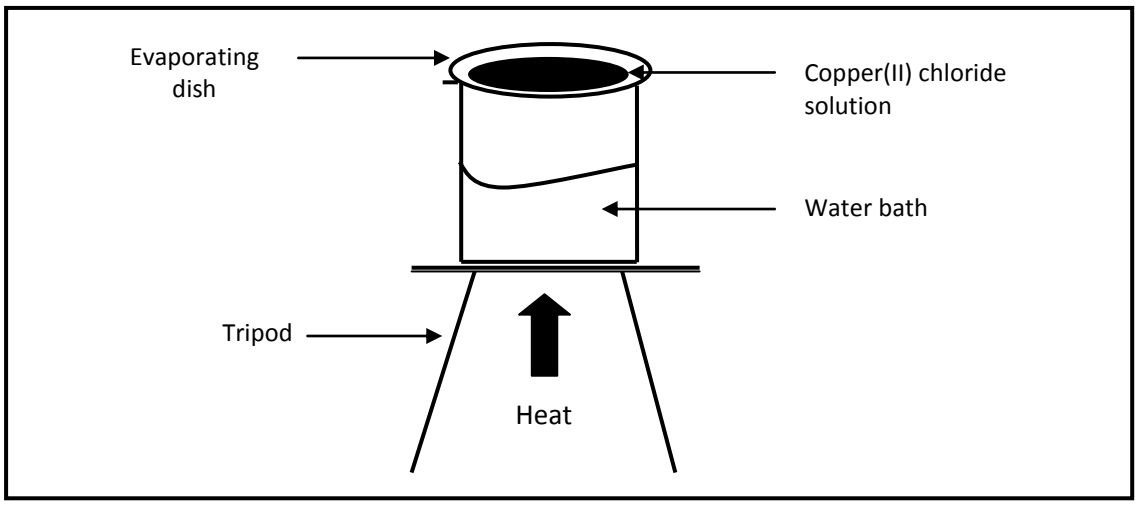
The reaction between copper(II)carbonate and hydrochloric acid.

4. Filter to remove unreacted copper(II)carbonate. You should get a blue-green solution.



Filtration process to remove unreacted copper(II)carbonate crystals.

5. Pour your solution into an evaporating dish. Then, heat it in water bath as shown in the diagram. Stop heating when you see some crystals around the edge of the solution.
   - What is your salt called?



Evaporation of copper(II) chloride using water bath

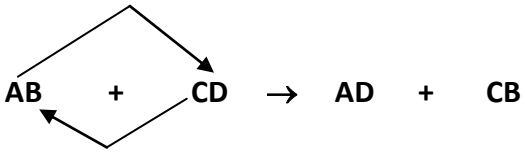
Observation from Experiment 2:

(i) When you add hydrochloric acid with copper carbonate, bubbles of hydrogen gas is evolved. The acid has used up the copper carbonate until no more copper carbonate will dissolve at that temperature. A solution of copper chloride is also produced.

(ii) The filtrate, copper chloride solution has evaporated to dryness to get green crystals of copper chloride.

2. Double Displacement Reaction

In a double displacement reaction, two (2) ionic compounds swap cations (positive ions). The general pattern of a double displacement reaction is:



A and C are cations (positively-charged ions) with B and D being the anions (negatively – charged ions). In a double displacement reaction, the cations swap places and two new compounds are formed. These two new substances are a gas, liquid or solid.

For example:

silver nitrate + sodium chloride → silver chloride + sodium nitrate

$AgNO_{3(aq)}$ + $NaCl_{(aq)}$ → $AgCl_{(s)}$ + $NaNO_{3(aq)}$

Double displacement reactions are further classified as **neutralization**, **precipitation,** and **gas formation** reactions.

- **Neutralization Reactions**

You have learnt neutralization reaction from the start of our module on Chemical Reactions. Neutralization reactions are specific kinds of double displacement reactions. An acid– base reaction occurs, when an acid reacts with equal quantity of base. The acid – base reaction results in the formation of salt (neutral in nature) and water.

Let us see another example to revise a neutralization reaction:

On mixing an aqueous solution of sulphuric acid (H_2SO_4) with an aqueous solution of potassium hydroxide (KOH), potassium chloride (KCl) and water (H_2O) are formed.

sulphuric acid + potassium hydroxide → potassium sulphate + water

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

- **Precipitation Reactions**

Precipitation is the formation of solid in the reaction of two solutions. We will discuss more on precipitation reaction in the next lesson.

- **Gas Formation Reaction**

A double displacement reaction should also occur if an **insoluble** (does not dissolve) gas is formed.

Gases such as hydrochloric acid (HCl) and ammonia (NH_3) are **soluble** (does dissolve) in water, but some other gases, such as hydrogen sulphide (H_2S), are insoluble in water.
Insoluble gases are often formed by the breakdown of unstable (uneven) double displacement reaction products. For example, marble chips or calcium carbonate ($CaCO_3$), react with dilute hydrochloric acid (HCl) to form calcium chloride ($CaCl_2$), and carbonic acid (H_2CO_3). Carbonic acid is unstable and readily decomposes to form carbon dioxide and water.

The equation for the reaction of calcium carbonate and the hydrochloric acid is shown below:

calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water

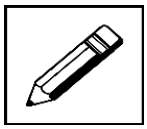
$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

The breaking down of carbonic acid, H_2CO_3 is as follows:

Carbonic acid → carbon dioxide + water

$H_2CO_{3(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

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



Learning Activity 4

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Answer the following questions:

1. State what happens when the following solutions are mixed. Also state if there is no reaction.
   a. flourine + sodium chloride _____
   b. chlorine + sodium bromide _____
   c. bromine + sodium iodide _____
   d. iodine + sodium chlorine _____

---

2. The table below shows the mixing of two reactants. State what would be observed and name the products formed.

Reactant	Observation	Names of Products Formed
a. $KI_{(aq)} + Cl_{2(g)}$		
b. $KBr_{(aq)} + Cl_{2(g)}$		
c. $KCl_{(aq)} + F_2$		

3. From the above reactants in 2, write a complete balanced equation for the reaction. States are required.

a. $KI_{(aq)} + Cl_{2(g)} \rightarrow$ _____

b. $KBr_{(aq)} + Cl_{2(g)} \rightarrow$ _____

c. $KCl_{(aq)} + F_{2(g)} \rightarrow$ _____

---

4. From the reaction given below, write an ionic equation to show which substance is oxidized and which is reduced.

a. $Na_{(s)} + CaCl_{2(aq)} \rightarrow$ _____

(i) Substance oxidized _____

(ii) Substance reduced _____

b. $K_{(s)} + AgNO_{3(aq)} \rightarrow$ _____

(i) Substance oxidized _____

(ii) Substance reduced _____

c. $Mg_{(s)} + CuSO_{4(aq)} \rightarrow$ _____

(i) Substance oxidized _____

(ii) Substance reduced _____

---

5. Predict the products of the following double displacement reactions and balance the equation if necessary.

a. $Na_2CO_{3(aq)} + CuSO_{4(aq)} \rightarrow$ _____

b. $(NH_4)_2SO_{4(aq)} + CaCl_{2(aq)} \rightarrow$ _____

c. $Na_2CO_{3(aq)} + Cu(NO_3)_{2(aq} \rightarrow$ _____

d. $Mg(OH)_{2(aq)} + HNO_{3(aq)} \rightarrow$ _____

---

6. An iron nail is placed separately in an aqueous solution containing both copper (II) sulphate and magnesium sulphate.

(i) Which metal or metals will be displaced?

_____

(ii) One of the reactions of the iron nail will not displace the metal in the above solution, which solution is this?

_____

(iii) Write a complete balanced equation for the reaction which you think an iron nail can react with. States are required.

_____

7. From the following ionic equations below, put a tick (√) if there is a reaction taking place and cross (X) if there is no reaction taking place.

a. $Zn^{2+}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + Zn_{(s)}$ _____

b. $Mg^{2+}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + Mg_{(s)}$ _____

c. $Fe^{2+}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$ _____

d. $Pb^{2+}_{(aq)} + Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + Pb_{(s)}$ _____

e. $Cu^{2+}_{(aq)} + Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + Cu_{(s)}$ _____

---

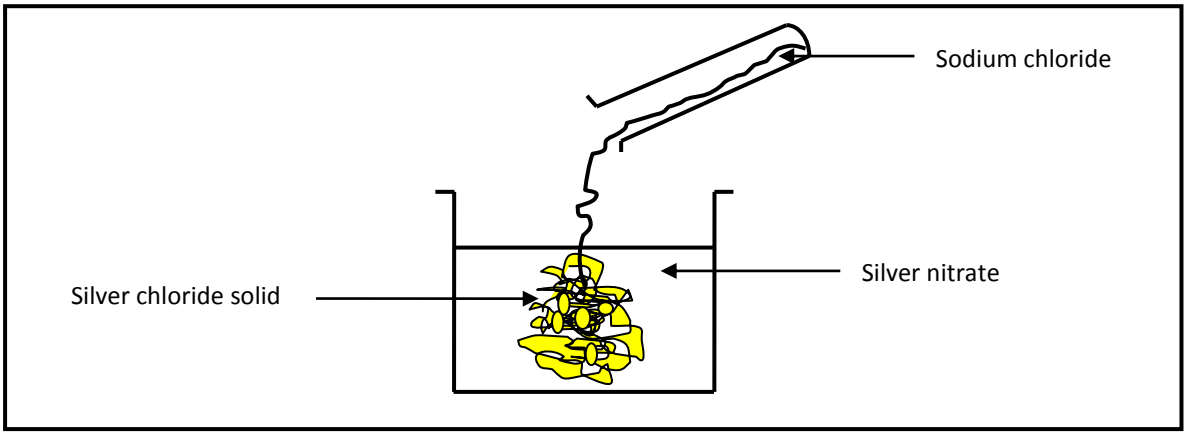
8. Define the word 'spectator ions'.

_____

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

D. Precipitation Reaction

You have learnt already the definition of precipitation reaction from our previous discussion, that a solid is formed when two (2) aqueous solutions are reacted together.



A precipitate of silver chloride that forms in a solution during a chemical reaction.

Again, the general equation form in a double displacement reaction is used:

AB + CD → AD + CB

The different chemical reactions involved in precipitation reaction are given below**:**

Reactions between two aqueous metal salts

1. Precipitates of halides (bromides, iodides and chlorides) of silver
   Think of two aqueous solutions that contain bromide ions and silver ions. Then proceed to the general equation form of double displacement reaction as shown above.

Example 1:
On mixing aqueous solution of potassium bromide and silver nitrate, a precipitate of silver bromide is formed in a solution of potassium nitrate.

The equation for the reaction is:

silver nitrate + potassium bromide → potassium nitrate + silver bromide

$AgNO_{3(aq)} + KBr_{(aq)} \rightarrow KNO_{3(aq)} + \mathbf{AgBr_{(s)}}$

Ionic equation with spectator ions is shown below in which they are struck out from the equation.

$Ag^{+}_{(aq)} + \cancel{NO_3^{-}{}_{(aq)}} + \cancel{K^{+}_{(aq)}} + Br^{-}_{(aq)} \rightarrow \cancel{K^{+}_{(aq)}} + \cancel{NO_3^{-}{}_{(aq)}} + \mathbf{AgBr_{(s)}}$

Therefore, the net ionic equation is, $\mathbf{Ag^{+}_{(aq)} + Br^{-}_{(aq)} \rightarrow AgBr_{(s)}}$

Example 2:
If silver nitrate solution is added to a solution of sodium chloride, sodium nitrate solution will form a white precipitate of silver chloride.

The equation for the reaction is:

silver nitrate + sodium chloride → sodium nitrate + silver chloride

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow NaNO_{3(aq)} + \mathbf{AgCl_{(s)}}$

Ionic equation:

$Ag^{+}_{(aq)} + \cancel{NO_3^{-}{}_{(aq)}} + \cancel{Na^{+}_{(aq)}} + Cl^{-}_{(aq)} \rightarrow \cancel{Na^{+}_{(aq)}} + \cancel{NO_3^{-}{}_{(aq)}} + \mathbf{AgCl_{(s)}}$

Therefore, the net ionic equation not involving the spectator ions is,

$\mathbf{Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)}}$

Can you write your own equation to produce a precipitate of silver iodide?

2. Precipitates of halides (bromides, iodides and chlorides) of lead
   Soluble lead ions can form precipitates with soluble ions of bromides, iodides and chlorides like silver in the above reactions.

For example, when you mix lead(II)nitrate and magnesium bromide solution together, a pale yellow precipitate of lead (II) bromide will form and a solution of magnesium nitrate.

The equation for the reaction is:

lead(II) nitrate + magnesium bromide → magnesium nitrate + lead bromide

$Pb(NO_3)_{2(aq)} + MgBr_{2\ (aq)} \rightarrow Mg(NO_3)_{2(aq)} + \mathbf{PbBr_{2(s)}}$

Ionic equation:

$Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + Mg^{2+}_{(aq)} + 2Br^-_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + \mathbf{PbBr_{2(s)}}$

Therefore, the net ionic equation not involving the spectator ions is:

$\mathbf{Pb^{2+}_{(aq)} + 2Br^-_{(aq)} \rightarrow PbBr_{2(s)}}$

Can you do your own equation to produce a precipitate of lead(II)iodide and lead(II) chloride?

3. Precipitates of sulphates of barium, calcium and lead
   Lead sulphate, barium sulphate, and calcium sulphate are all insoluble solids. They remain solid in water.

To make an insoluble lead(II) sulphate, take a look on the chemical reactions below by mixing two soluble salts of lead (II) nitrate, and copper(II) sulphate.

The equation for the reaction is shown below:

Lead(II) nitrate + copper(II) sulphate → copper nitrate + lead sulphate

$Pb(NO_3)_{2(aq)} + CuSO_{4(aq)} \rightarrow Cu(NO_3)_{2(aq)} + \mathbf{PbSO_{4(s)}}$

Ionic equation:

$Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} + \mathbf{PbSO_{4(s)}}$

Therefore, the net ionic equation not involving the spectator ions is,

$\mathbf{Pb^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)}}$

Can you think of two soluble salts to combine together to make separate precipitates of barium sulphate and calcium sulphate?

4. Precipitates of metal hydroxides (except Group I hydroxides and ammonium compounds)
   Most metal hydroxides form precipitates like hydroxides of transition metal such as copper hydroxide, zinc hydroxide, iron(II) hydroxide, and many more. Group II hydroxides are slightly insoluble in water, such as magnesium hydroxide, and calcium hydroxide.

However, Group I hydroxides such as sodium hydroxide and potassium hydroxide, including ammonium compounds like ammonium hydroxide are soluble in water. They do not form precipitates in any chemical reactions.

Example 1:
When a solution of copper(II) sulphate is reacted with potassium hydroxide solution, a blue precipitates of copper(II) hydroxide and potassium sulphate are formed.

The equation for the reaction is shown below:
Copper(II) sulphate + potassium hydroxide → copper(II) hydroxide + potassium sulphate

$$CuSO_{4(aq)} + 2KOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + K_2SO_{4(aq)}$$

Ionic equation:

$$Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2K^{+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow \mathbf{Cu(OH)_{2(s)}} + 2K^{+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Therefore, the net ionic equation not involving the spectator ions is,

$$\mathbf{Cu^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Cu(OH)_{2(s)}}$$

Example 2:
Iron(II)chloride solution and ammonium hydroxide solution are combined together and the reaction has formed a brown precipitate of iron hydroxide and ammonium chloride.

The equation for the reaction is shown below:

Iron(II)chloride + ammonium hydroxide → iron(II)hydroxide + ammonium chloride

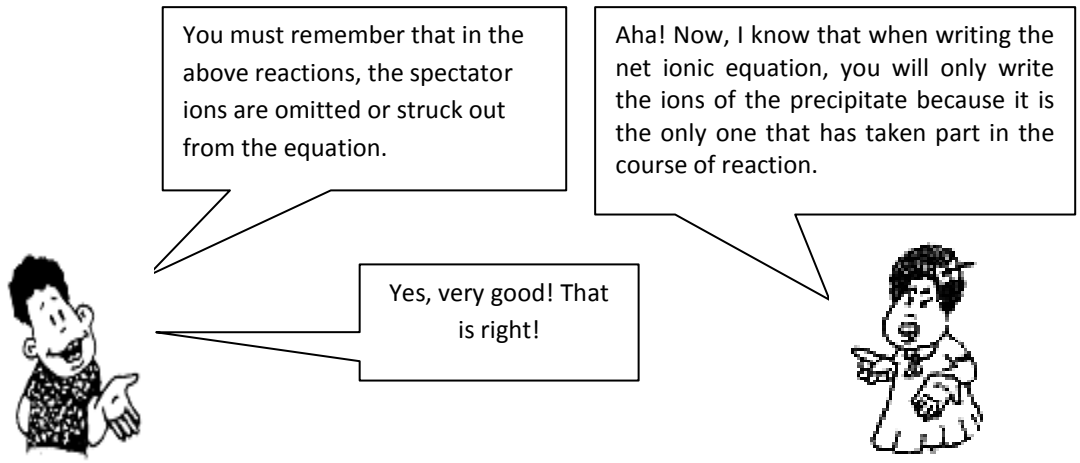
$$FeCl_{2(aq)} + 2NH_4OH_{(aq)} \rightarrow \mathbf{Fe(OH)_{2(s)}} + 2NH_4Cl_{(aq)}$$

Ionic equation;

$$Fe^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2NH_4^{+}{}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Fe(OH)_{2(s)} + 2NH_4^{+}{}_{(aq)} + 2Cl^{-}_{(aq)}$$

Therefore, the net ionic equation not involving the spectator ions is,

$$\mathbf{Fe^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Fe(OH)_{2(s)}}$$



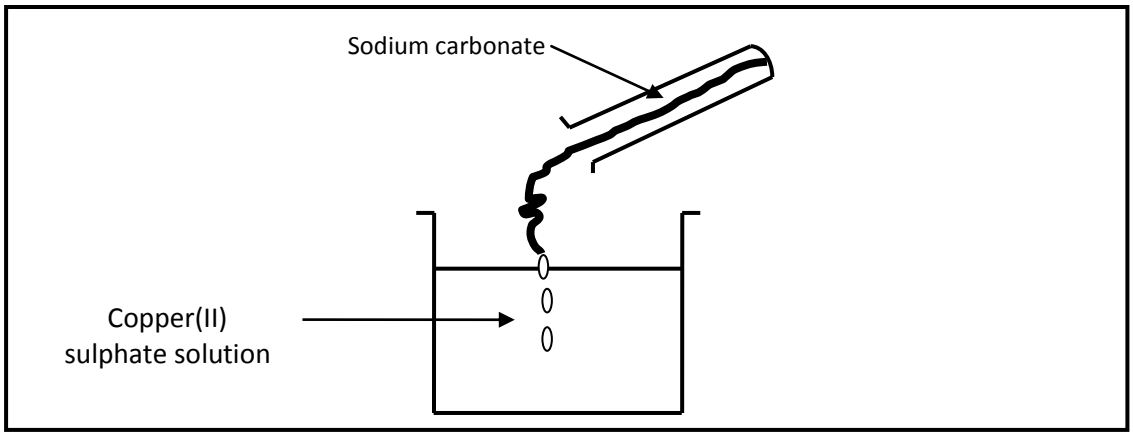
Experiment 3: Formation of precipitate between two aqueous salt solutions

Aim:

To prepare a sample of copper(II) carbonate by the precipitation method

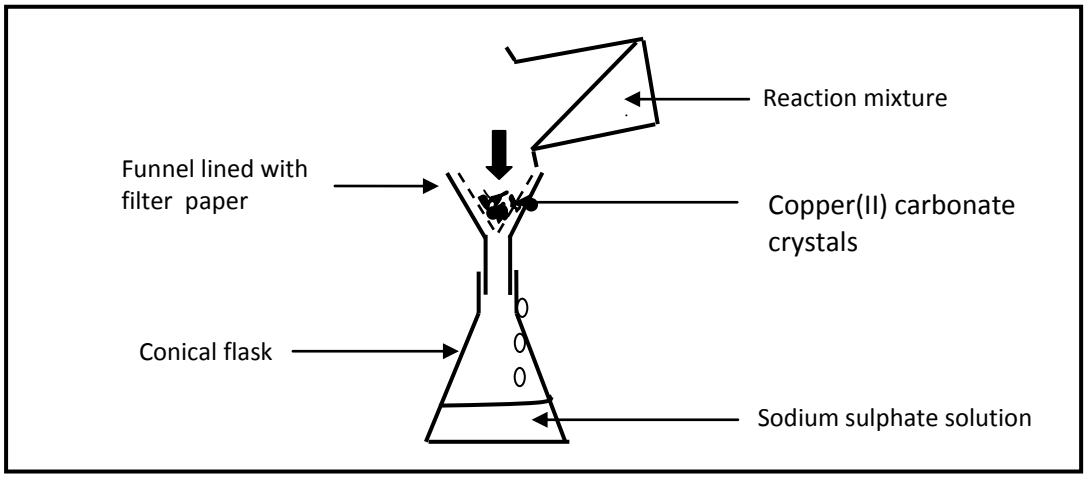
Procedures:

1. Pour about 50cm^3 of copper(II) sulphate solution into a small beaker. Add sodium carbonate solution (in excess) until no more precipitate forms.



Reaction of sodium carbonate and copper(II) sulphate.

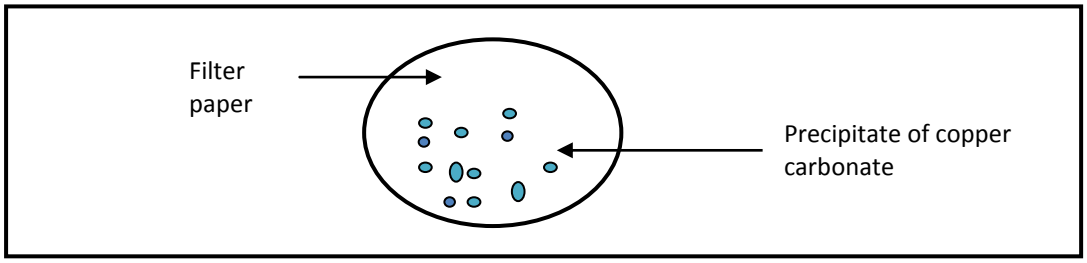
2. Filter to remove the precipitate.



Filtration of copper(II) carbonate

3. Wash the precipitate with a small amount of distilled water to remove any impurities.

4. Allow the precipitate to dry on a piece of filter paper.



Precipitate of copper(II) carbonate after drying on a piece of filter paper.

Observation from Experiment 3:

From the reaction between copper(II) sulphate and sodium carbonate, a precipitate of copper(II) carbonate has formed in a solution of sodium sulphate.

(i) The colour of the salt is blue – green precipitate.

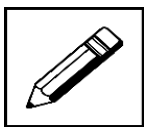
(ii) The crystals of copper carbonate were washed with distilled water to remove impurities.

(iii) The collected precipitate is dried with the use of filter paper.

(iv) The equation for the reaction is therefore;

$CuSO_{4(aq)} + Na_2CO_{3(aq)} \rightarrow CuCO_{3(s)} + Na_2SO_{4(aq)}$

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



Learning Activity 5

0 minutes

Answer the following questions:

1. Name and write the formula of the insoluble solid or precipitate that will form when the two aqueous solutions given below are mixed together.

   (i) Lead nitrate with sodium sulphate
   Name of precipitate
   _____
   Formula
   _____

   (ii) Copper sulphate with sodium hydroxide
   Name of precipitate
   _____
   Formula
   _____

(iii) Lithium chloride with silver nitrate
Name of precipitate

_____

Formula

_____

2. Iron (II) chloride is reacted with sodium hydroxide to form a precipitate iron hydroxide and sodium chloride solution.

(i) Write a complete balanced equation for the reaction above. States are required.

_____

(ii) Write the ionic equation involving the spectator ions in the above reaction. States are required.

_____

(iii) Write the net ionic equation for the reaction. States are required.

_____

3. Barium sulphate is insoluble.

(i) Name two solutions which you could mix to give a precipitate of barium sulphate from the mixture.

_____ and _____

(ii) Describe what you would do to obtain barium sulphate from the mixture?

_____

(iii) Write a word equation for the reaction.

_____

(iv) Write a balanced equation for the reaction. States are required.

_____

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

E. Decomposition Reaction

Decomposition reaction is defined as the breaking down of a single compound into two or more compounds or elements. This reaction is the reverse of synthesis reaction so they work both ways. Most decomposition reactions require heat.

General Equation Form:	**AB** $\rightarrow$	**A**	**+**	**B**

Types of Decomposition Reaction

1. Decomposition of metallic carbonate

Metallic carbonates decompose into metal oxides and carbon dioxide when heated.

Example 1

When calcium carbonate (marble chips) is heated strongly, calcium oxide (lime) and carbon dioxide gas are formed.

The complete balanced equation for the reaction is,

calcium carbonate	→	calcium oxide	+	carbon dioxide
$CaCO_{3(s)}$	→	$CaO_{(s)}$	+	$CO_{2(g)}$

Example 2

Copper (II) carbonate is decomposed to produce copper (II) oxide (black solid) and carbon dioxide.

The complete balanced equation for the reaction is,

copper carbonate	→	copper oxide	+	carbon dioxide
$CuCO_{3(s)}$	→	$CuO_{(s)}$	+	$CO_{2(g)}$

2. Decomposition of metallic hydroxide

Metallic hydroxides (except sodium hydroxide and potassium hydroxide) decompose into metallic oxides and water when heated.

Example 1

Calcium hydroxide will form calcium oxide and water during decomposition reaction.

The complete balanced equation for the reaction is,

calcium hydroxide	→	calcium oxide	+	water
$Ca(OH)_{2(s)}$	→	$CaO_{(s)}$	+	$H_2O_{(g)}$

Example 2

When zinc hydroxide is heated, zinc oxide and water are produced.

The complete balanced equation for the reaction is,

zinc hydroxide	→	zinc oxide	+	water
$Zn(OH)_{2(s)}$	→	$ZnO_{(s)}$	+	$H_2O_{(g)}$

3. Decomposition of metallic hydrogen carbonate

Metallic hydrogen carbonates decompose into metal oxide, water and carbon dioxide gas.

Example 1

Sodium hydrogen carbonate will form sodium oxide, water and carbon dioxide during decomposition reaction.

The complete balanced equation for the reaction is,

sodium hydrogen carbonate → sodium oxide + water + carbon dioxide

$2NaHCO_{3(s)} \rightarrow Na_2O_{(s)} + H_2O_{(g)} + 2CO_{2\,(g)}$

Example 2

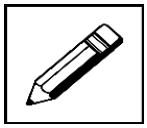
When magnesium hydrogen carbonate is heated, the products of the decomposition are magnesium oxide (white solid), water and carbon dioxide.

The complete balanced equation for the reaction is,

magnesium hydrogen carbonate → magnesium oxide + water + carbon dioxide

$Mg(HCO_3)_{2\,(s)} \rightarrow MgO_{(s)} + H_2O_{(g)} + 2CO_{2\,(g)}$

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



Learning Activity 6

60 minutes

Answer the following questions.

1. Write a chemical equation in words and then write beneath it a balanced equation using formulae for each reaction given.

   a. The heating of magnesium carbonate produces white magnesium oxide with carbon dioxide gas being evolved.
   Word equation:
   _____
   Symbol equation:
   _____

   b. Copper hydroxide is decomposed forming copper oxide and water.
   Word equation:
   _____
   Symbol equation:
   _____

   c. Potassium oxide, carbon dioxide, and water are produced when potassium hydrogen carbonate is heated.
   Word equation:
   _____
   Symbol equation:
   _____

2. Balance the following decomposition reactions below. If the equation is balanced, write the word balanced next to the equation.

(i) $Al_2(CO_3)_{3(s)} \rightarrow Al_2O_{3(s)} + CO_{2(g)}$

(ii) $CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$

(iii) $Fe_2(CO_3)_{3(s)} \rightarrow Fe_2O_{3(s)} + CO_{2(g)}$

(iv) $Mg(HCO_3)_{2(s)} \rightarrow MgO_{(s)} + CO_{2(g)} + H_2O_{(l)}$

(v) $Li_2CO_{3(s)} \rightarrow Li_2O_{(s)} + CO_{2(g)}$

(vi) $Al(OH)_{3(s)} \rightarrow Al_2O_{3(s)} + H_2O_{(l)}$

(vii) $Fe(OH)_{2(s)} \rightarrow FeO_{(s)} + H_2O_{(g)}$

3. What is a decomposition reaction?

_____

_____

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

F. Combustion Reaction

Every time you strike a match, burn a candle, build a fire, or light a grill, you see the combustion reaction. Combustion reaction always involves molecular oxygen, O_2.
This process is commonly called the process of **burning.** Anytime, anything burns, it is combustion reaction. Three things are needed to make fire: **fuel, heat** and **oxygen**. This is called the **fire triangle.** If one part is removed, the fire goes out.
The reaction of a fuel with oxygen is called **combustion**. Combustion reactions are almost always **exothermic** (meaning they give off heat).

For example, when wood burns, it must do so in the presence of oxygen and a lot of heat is produced.

The combustion reaction showing the release of carbon dioxide

Wood as well as many common items that burn is organic which means they are made up of carbon, hydrogen and oxygen. When organic molecule burns, the reaction products are **carbon dioxide** and **water** (as well as heat).

Types of Combustion Reactions

1. Combustion of hydrocarbons (methane, ethane, propane, and butane)
   Hydrocarbons are compounds containing only carbon and hydrogen. They are the fuels we get from crude oil which contains mixture of hydrocarbons such as methane (CH_4), ethane (C_2H_6), propane (C_3H_8) and butane (C_4H_{10}).

Hydrocarbon	**+**	**Oxygen**	**→**	**Carbon dioxide**	**+**	**water**

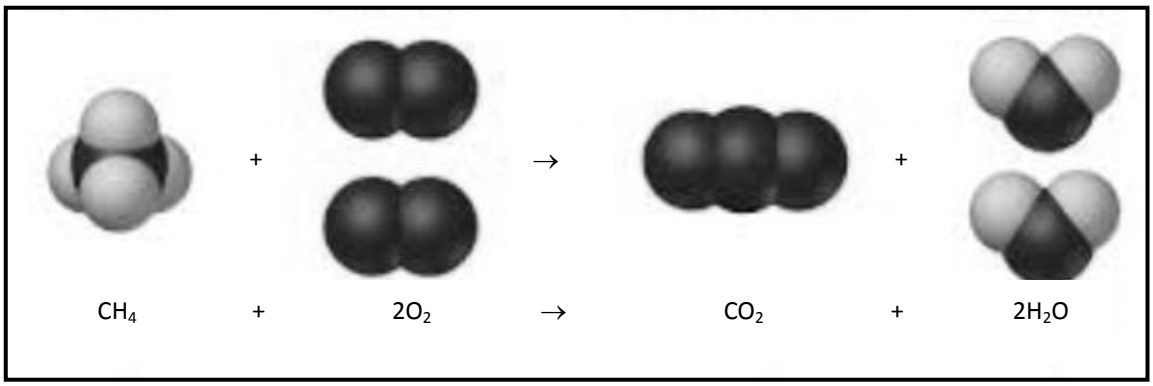
If any flame is seen in the combustion, it is called **burning**. **Carbon dioxide** and **water** are produced if the 'burning 'is complete. This is called **complete combustion** because there is enough oxygen supply in the reaction vessel. However, during an **incomplete combustion**, **carbon monoxide** and **water** are produced because there is less oxygen to support the burning process.

We can consider respiration as combustion without flames, whereas the combustion of fossil fuels such as coal and petroleum are examples of burning. Due to rapid industrialization and clearing of forest by burning large quantities, carbon dioxide is released into the atmosphere.

As a result, too much carbon dioxide in the atmosphere will cause global warming. **Global warming** has caused much serious harm on earth such as drastic weather changes like floods, droughts, and heat waves. The melting of the polar ice caps and glaciers will eventually lead to flooding of many coastal countries in the world.

Study the following examples of combustion reaction of hydrocarbon below:

a. Combustion of methane



Combustion of methane producing carbon dioxide and water

b. Combustion of butane (C_4H_{10})

butane + oxygen → carbon dioxide + water

$2C_4H_{10(g)} + 13O_{2(g)} \rightarrow 8CO_{2(g)} + 10H_2O_{(g)}$

c. Combustion of magnesium

When a solid magnesium ribbon is burned with oxygen, a white solid ash is formed called magnesium oxide.

The balanced equation for the reaction is:

magnesium + oxygen → magnesium oxide

$2Mg_{(s)} + O_{2\,(g)} \rightarrow 2MgO_{(s)}$

d. Combustion of sulphur

The burning of sulphur (yellow solid) in oxygen will form sulfur dioxide, an acidic substance which is present in acid rain.

The balanced equation is given as:

sulphur + oxygen → sulphur dioxide

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

e. Combustion of carbon

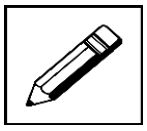
The burning of carbon with oxygen will form carbon dioxide, a slightly acidic gas which is found in the atmosphere and is important in the process of photosynthesis (process of food making).

The reaction is shown in the equation,

carbon + oxygen → carbon dioxide

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

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



Learning Activity 7

0 minutes

Answer the following questions:

1. Define the word combustion.

_____

---

2. List the two (2) types of combustion reaction.

a.

_____

b.

_____

3. Show a balanced equation for the complete and incomplete combustion of methane (CH_4). States are required.
   (i) Complete combustion of methane
   _____
   (ii) Incomplete combustion of methane
   _____

4. List down the three things needed to make fire.
   (i) _____
   (ii) _____
   (iii) _____

5. Write a balanced equation for the combustion of carbon.
   _____

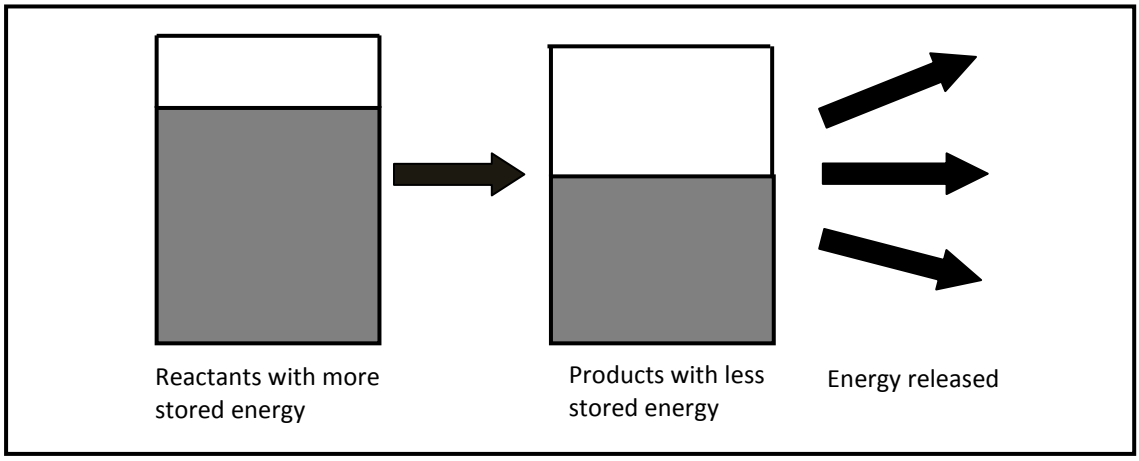
6. What is the effect on Earth if there is too much carbon dioxide in the atmosphere as a result of combustion process?
   _____

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

G. Exothermic and Endothermic Reactions

Exothermic reactions

In the previous lesson, you looked at the combustion of fuel. Think again about what happens when you burn a fuel. Is energy given out or taken in as the fuel burns?



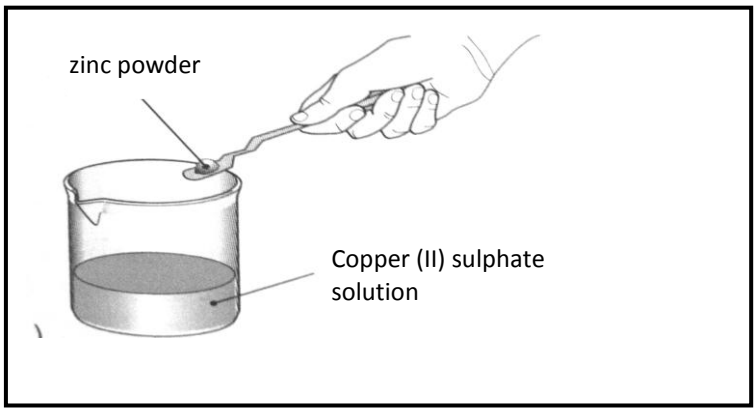
An exothermic reaction showing that heat is given out to the surrounding.

For example, when you collect 25mL of copper sulphate solution in a small beaker and add a spatula (spoonful) of zinc powder. Tthen stir it with a glass rod. There is an indication that a chemical reaction will occur as follows:

The zinc will displace copper from copper sulphate solution as in the equation below:

Zinc + copper(II) sulphate → zinc sulphate + copper

$Zn_{(s)}$ + $CuSO_{4\,(aq)}$ → $ZnSO_{4\,(aq)}$ + $Cu_{(s)}$



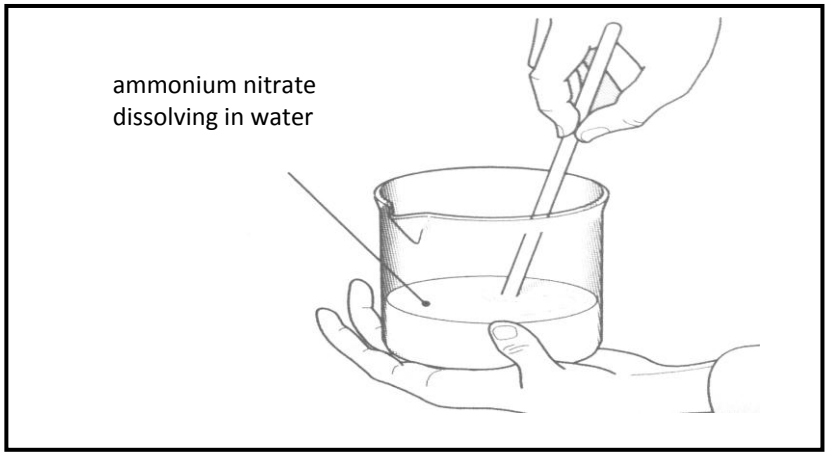
An exothermic reaction between zinc and copper sulphate.

Therefore, the displacement of reaction between zinc and copper gives out heat. This reaction is an example of **exothermic reaction**.

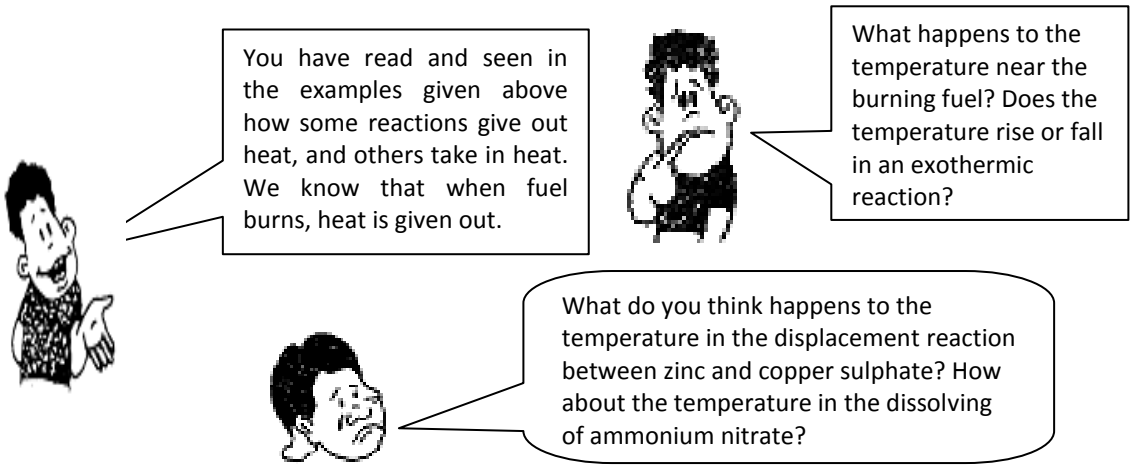
Endothermic reactions

These are reactions that take in energy from the surroundings. This energy is usually transferred as heat energy, causing the reaction mixture and its surroundings to get cooler. The temperature decrease can also be detected using a thermometer.

For example, when 3 spatulas of ammonium nitrate is added to 25cm^3 of water in a small beaker and stir with a glass rod, then hold the beaker in the palm of your hand, the outside of the beaker will feel cold . This is because when ammonium nitrate dissolves, it takes in heat energy from its surroundings. Its surroundings include the beaker, glass rod, the around it, and your hand.



An example of endothermic reaction.

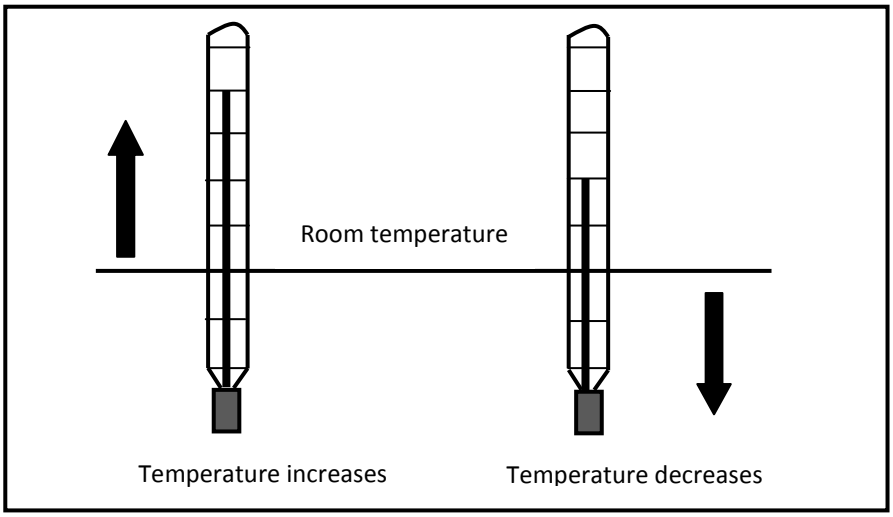


In an exothermic reaction, heat energy is given out during the reaction. This raises the temperature of the surroundings. The **products** formed are at **lower energy** level than that of the **reactants.** Hence, the excess energy is given out to the surroundings.

On the other hand, in an endothermic reaction heat energy is absorbed from the surroundings during the reaction. This makes the temperature of the surroundings fall. The **products** formed are at **high energy** level than that of the **reactants.** Hence, energy is absorbed by the reactants from the surroundings.

Exothermic reaction releases heat to the surroundings, while endothermic reaction absorbs heat from the surroundings.

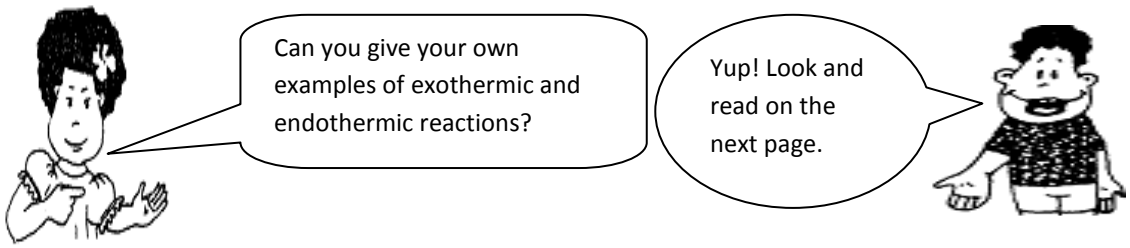
What happens to the temperature?
Look at the diagram to check your ideas.



In exothermic reaction, temperature increases while in endothermic reaction temperature decreases.

In exothermic reactions, the products formed are at lower energy level than that of the reactants .In endothermic reaction, the products formed are at high energy level than that of the reactants.

When trying to classify a process as exothermic or endothermic, watch how the temperature of the surrounding changes. Remember, exothermic process releases heat and causes the temperature of the immediate surroundings to rise. An endothermic process absorbs heat and cools the surrounding.



Exothermic processes	Endothermic processes

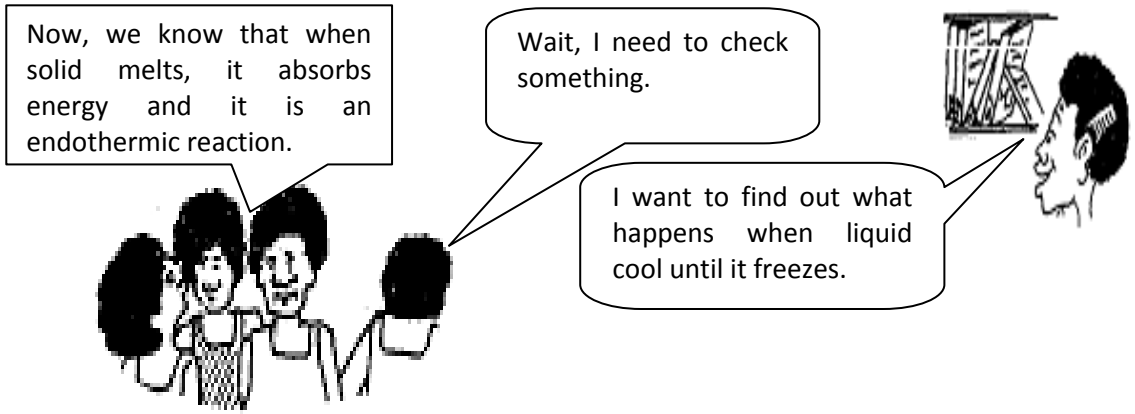
Freezing like making ice cubes (solidification – change of liquid state to solid).	Melting of ice cubes (change of state from solid to liquid).
Formation of snow in clouds.	Conversion of frost to water vapour (change of solid state to gas).
Condensation of rain from water vapour (change of gas state to liquid).	Evaporation of water(change of liquid state to gas).
Combustion like a candle flame.	Baking bread.
Oxidation like rusting of iron.	Cooking an egg.
Combining of atoms to make molecule like H_2, O_2 and others.	Producing sugar by photosynthesis
Mixing water and strong acids.	Thermal (heat) decomposition like splitting a gas molecule apart.
Mixing water with anhydrous (no water) salts.	Mixing water and ammonium nitrate.
Neutralization like mixing of acid and base.	Dissolving sodium carbonate in water.
Producing carbon dioxide and water from sugar by respiration	Taking photograph with a film – light energy is absorbed to decompose the silver bromide.

From the examples given above, it is easy to understand the difference between exothermic and endothermic reactions using the **change of states**.

For example, when **solid ice melts to liquid**, the following points will happen as a sign that **endothermic reaction** is taking place:

(i) Heat energy is absorbed by the particles in the solid. The heat energy is converted into kinetic energy. The particles start to vibrate faster about their fixed positions.
(ii) When the temperature is high enough, the vibrations of the particles become sufficient to overcome the attractive forces between them. The particles begin to break away from their fixed positions.
(iii) The particles are no longer in the fixed positions. The particles slide over one another and the substance is now a liquid.

The same thing will happen to frost changing into a gas (sublimation), evaporation of water, photosynthesis, decomposition and to all the above examples given under endothermic reaction. They all **absorb energy** from the surrounding environment.

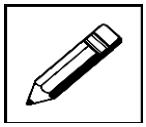


When a **liquid becomes a solid**, an **exothermic reaction** occurs as described below:

- Energy is given out by the particles of a liquid. The particles lose kinetic energy and begin to move more slowly.
- When the temperature is low enough, the particles no longer have enough energy to move about freely. Some particles start to settle into fixed positions.
- All the particles settle into fixed positions. Particles can only vibrate about their fixed positions and the substance is now solid.

The same thing will happen to a neutralization reaction, respiration, combustion and condensation of gas state into liquid state and to all the above examples given under exothermic reaction. They all **give out heat energy** to the surrounding environment.

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



Learning Activity 8

40 minutes

Answer the following questions.

1. Identify the following reactions as exothermic and endothermic reactions.

   a. Burning of petrol _____
   b. Reacting acids and base _____
   c. Dissolving ammonium nitrate _____
   d. Green leaves absorb light energy to make starch _____
   e. Dissolving solid sodium hydroxide in water _____
   f. Decomposition of calcium carbonate _____
   g. Oxidation of food in the body to make body warm _____

2. Define:
   a. exothermic reaction
   _____
   b. endothermic reaction
   _____

---

3. Complete the sentences below:
   (i) In an _____ reaction, heat energy is released to the surroundings.
   (ii) An _____reaction, heat energy absorbed from the surroundings.
   (iii) The products formed in an exothermic reaction are at _____ energy level than the reactants.
   (iv) The products formed in an endothermic reaction are at _____ energy level than that of the reactants.
   (v) Temperature falls in an _____ reaction.
   (vi) Temperature rise in an _____ reaction.

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

REVISE WELL USING THE MAIN POINTS ON THE NEXT PAGE.

SUMMARY

You will now revise this module before doing Assessment 3. Here are the main points to help you revise. Refer back to the module topic if you need more information.

- **Chemical changes** are changes in the composition and structure of a substance. They are always accompanied by energy changes. Chemical changes are called chemical reactions.
  - **Law of Conservation of Mass** states that matter can neither be created nor destroyed. In chemical reactions, mass does not change.
- **Periodic Table** is a table in which the chemical elements are arranged in order of increasing atomic number.
  - Elements with similar properties are arranged in the same column (group).
  - Elements with same number of electron shells are arranged in the same row (period).
- **Chemical equations**
  - **Word equations** represent the names of the reactants and products that are written on each side of the arrow to make a word equation for the reaction.

  Example: carbon + oxygen $\rightarrow$ carbon dioxide
  - **Symbol equation** shows that the word equations are replaced with the symbols of the elements.

  Example: $C + O_2 \rightarrow CO_2$
  - **State symbols** are added to the reactants and products after completing the symbol equation. The state symbols are (s) for solid, (l) for liquid, (aq) for aqueous and (g) for gas.

  Example: $C_{(g)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$
- **Indicators of chemical change**
  - **If there is a significant change in temperature** – some reactions are exothermic because they release heat. Others are endothermic because they absorb heat. By monitoring the temperature of the reaction mixture using a thermometer, it is possible to note if there is a reaction going on even if the formation of solid or a colour change is absent.
  - **If a solid is formed** – a solid that is formed in a chemical reaction is called a precipitate. Precipitate forms when two aqueous solutions (when an ionic solid dissolves in water) are reacted together.
  - **If there is a change in colour** - in chemical reactions, some coloured products are formed and are easily noticed. For example, if silver nitrate solution is added to a solution of sodium chloride, sodium nitrate solution will form and a white precipitate of silver chloride.

The equation for the reaction is:

silver nitrate + sodium chloride → sodium nitrate + silver chloride

$AgNO_{3(aq)}$ + $NaCl_{(aq)}$ → $NaNO_{3\,(aq)}$ + **$AgCl_{(s)}$**

colourless colourless colourless white

- **If a gas is given off** – there are reactions which gases are evolved. Reactions of metals with acids lead to the formation of hydrogen gas. Metallic carbonates forms carbon dioxide gas.
  For example, when hydrochloric acid reacts with copper carbonates, it forms copper chloride, water and carbon dioxide.

The equation is;

hydrochloric acid + copper carbonates → copper chloride + water + carbon dioxide

$HCl_{(aq)}$ + $CuCO_{3(s)}$ → $CuCl_{2(aq)}$ + $H_2O_{(l)}$ + $CO_{2(g)}$

- **If there is a disappearance of solid** – when some solids react with acid and solid will disappear forming new products. This happens because the acid is using up the solid like metal in particular.

- **Types of chemical reactions**
  - **Synthesis reaction** is a type of reaction where a substance is formed as a result of chemical combination. In this reaction, two or more reactants form one product.
    The general form for a synthesis reaction is: **A + B → AB**.
  - **Neutralization reactions** are a specific kind of a double displacement reaction.
    An acid- base reaction occurs, when an acid reacts with equal quantity of base. The acid – base reaction results in the formation of salt (neutral in nature) and water.
    The general equation is: **Acid + Base → Salt + Water**
  - **Displacement Reaction** is a type of reaction in which a less reactive element is displaced by a more reactive element in a chemical reaction.
  - **Precipitation reaction** is the formation of insoluble solids in a reaction. Precipitates usually form when two metal aqueous salts are combined.
  - **Decomposition** is a type of chemical reaction where one reactant yields two or more products. It is the opposite of a synthesis or combination reaction. Most of decomposition reactions require heat.

    The general equation form is: **AB → A + B**

- **Types of Synthesis Reaction**
  - **Metal + Oxygen Gas → Metal oxide**

Example:

calcium	+	oxygen	→	calcium oxide
$2Ca_{(s)}$	+	$O_{2(g)}$	→	$2CaO_{(s)}$

  - **Metal + Sulphur → Metal sulphide**

Example:

zinc	+	sulphur	→	zinc sulphide
$Zn_{(s)}$	+	$S_{(s)}$	→	$ZnS_{(s)}$

  - **Metal + Halogen → Metal halide**

Example:

sodium	+	fluorine	→	sodium flouride
$2Na_{(s)}$	+	$F_{2(g)}$	→	$2NaF_{(s)}$

  - **Non - metal + Oxygen gas → Non-metal oxide**

Examples:

Carbon	+	oxygen	→	carbon dioxide
$C_{(s)}$	+	$O_{2(g)}$	→	$CO_{2(g)}$

- **Salts** are compounds made when we replace the hydrogen in an acid by a metal. Examples of salts are sodium chloride (NaCl), copper sulphate ($CuSO_4$) and silver nitrate ($AgNO_3$).
  - Hydrochloric acid (HCl) makes salts called **chlorides.**
  - Sulphuric acid (H_2SO_4) gives **sulphates.**
  - Nitric acid (HNO_3) gives **nitrates.**

- **Types of Neutralization Reaction**
  - **Acid + Aqueous ammonia solution → Salt + Water**

Examples:

hydrochloric acid	+	ammonium hydroxide	→	ammonium chloride	+	water
$HCl_{(aq)}$	+	$NH_4OH_{(aq)}$	→	$NH_4Cl_{(aq)}$	+	$H_2O_{(l)}$

Ionic equation:

$$H^+_{(aq)} + Cl^-_{(aq)} + NH_4^+{}_{(aq)} + OH^-_{(aq)} \rightarrow NH_4^+{}_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

Net ionic equation:

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

Spectator ions are the ions which do not take part in the course of reaction.

- **Acid + Metal carbonate → Salt + Water + Carbon dioxide**

Example:

hydrochloric acid + copper carbonate → copper chloride + water + carbon dioxide

$2HCl_{(aq)} + CuCO_{3(s)} \rightarrow CuCl_{2\,(aq)} + H_2O_{(l)} + CO_{2(g)}$

Ionic equation:

$2H^+_{(aq)} + 2Cl^-_{(aq)} + Cu^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Cl^-_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Net ionic equation:

$2H^+_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow H_2O_{(l)} + CO_{2\,(g)}$

Acids + Metal oxide → Salt + Water

Example:

sulphuric acid + copper oxide → copper sulphate + water

$H_2SO_{4\,(aq)} + CuO_{\,(s)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$

Ionic equation:

$2H^+_{(aq)} + SO_4^{2-}{}_{(aq)} + Cu^{2+}_{(aq)} + O^{2-}_{(aq} \rightarrow Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + H_2O_{(l)}$

Net ionic equation:

$2H^+_{(aq)} + O^{2-}_{(aq)} \rightarrow H_2O_{(l)}$

- **Acid + Metal hydroxide → Salt + Water**

Example:

hydrochloric acid + potassium hydroxide → potassium chloride + water

$HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

Ionic equation:

$H^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$

Net Ionic equation:

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

- **Acid + Metal hydrogen carbonate → Salt + carbon dioxide + Water**

Example:

nitric acid + sodium hydrogen carbonate → sodium nitrate + carbon dioxide + water

$HNO_{3(aq)} + NaHCO_{3(aq)} \rightarrow NaNO_{3(aq} + CO_{2(g)} + H_2O_{(l)}$

Ionic equation:

$H^+_{(aq)} + NO_3^-{}_{(aq)} + Na^+_{(aq)} + HCO_3^-{}_{(aq} \rightarrow Na^+_{(aq)} + NO_3^-{}_{(aq)} + CO_{2(g)} + H_2O_{\,(l)}$

The net Ionic equation:

$H^+_{(aq)} + HCO_3^-{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

- **Types of Displacement Reaction**
  - **Single displacement** is a type of reaction where a more active element displaces another less active element from a compound.

For example, if you put a piece of iron nail into a copper sulphate solution, the iron (more reactive) displaces the copper (less reactive), as shown in this equation:

Iron + copper(II) sulphate → iron(II)sulphate + copper

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

Iron (Fe) is oxidized to iron ion (Fe^{2+}) in iron sulfate solution and copper ion(Cu^{2+}) is reduced to solid copper (Cu).

- **Double displacement** is a type of reaction when two ionic compounds are mixed together. The two ionic compounds switch cations (positive ions).

The general pattern of a double displacement is:

AB + CD → AD + CB

Example 1:

ammonium sulphide + iron(II) sulphate → iron sulphide + ammonium sulphate

$(NH_4)_2S_{(aq)} + FeSO_{4(aq)} \rightarrow FeS_{(s)} + (NH_4)_2SO_{4(aq)}$

Example 2:

lead(II) nitrate + sodium sulphate → lead(II) sulphate + sodium nitrate

$Pb(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow PbSO_{4(s)} + 2NaNO_{3(aq)}$

- **Reactivity Series** is a list of metals in order of increasing reactivity.
  You can use the Reactivity Series to make predictions about reactions. A more reactive metal can **displace** a less reactive metal from its compounds.

  For example: copper oxide + zinc → zinc oxide + copper

- **Redox reaction** is a reaction in which one substance is reduced and the other substance is oxidized.
  - **Oxidation** is the loss of electrons, gain of oxygen, loss of hydrogen, and increased in oxidation number.
  - **Reduction** is the gain of electrons, loss of oxygen, gain of hydrogen, and decreased in oxidation number.

- **Types of Double Displacement Reactions**
  - **Neutralization reaction** is the formation of neutral salt and water.
  - **Precipitation reaction** is the formation of insoluble solids called precipitate.
  - **Gas formation reaction** is the formation of gas in a reaction.

- **Precipitation reaction** is the formation of insoluble solids in a reaction. Precipitates usually form when two metal aqueous salts are combined.

- **Solubility rules** – describes which compounds are soluble and which are insoluble. This rule is used to decide if a precipitate forms in a reaction.

- **Types of Precipitation Reactions**
  - **Precipitates of halides of silver**

Example:

silver nitrate + potassium bromide → potassium nitrate + silver bromide

$AgNO_{3(aq)} + KBr_{(aq)} \rightarrow KNO_{3(aq)} + \mathbf{AgBr_{(s)}}$

Ionic equation:

$Ag^{+}_{(aq)} + NO_3^{-}{}_{(aq)} + K^{+}_{(aq)} + Br^{-}_{(aq)} \rightarrow K^{+}_{(aq)} + NO_3^{-}{}_{(aq)} + \mathbf{AgBr_{(s)}}$

Net ionic equation:

$\mathbf{Ag^{+}_{(aq)} + Br^{-}_{(aq)} \rightarrow AgBr_{(s)}}$

  - **Precipitates of halides (bromides, iodides and chlorides) of lead**

Example:

lead(II)nitrate + magnesium bromide → magnesium nitrate + lead bromide

$Pb(NO_3)_{2(aq)} + MgBr_{2\ (aq)} \rightarrow Mg(NO_3)_{2(aq)} + \mathbf{PbBr_{2(s)}}$

The ionic equation is,

$Pb^{2+}_{(aq)} + 2NO_3^{-}{}_{(aq)} + Mg^{2+}_{(aq)} + 2Br^{-}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2NO_3^{-}{}_{(aq)} + \mathbf{PbBr_{2(s)}}$

Net ionic equation:

$\mathbf{Pb^{2+}_{(aq)} + 2Br^{-}_{(aq)} \rightarrow PbBr_{2(s)}}$

  - **Precipitates of sulphates of barium, calcium and lead**

Example:

lead (II) nitrate + copper (II) sulphate → copper nitrate + lead (II) sulphate

$Pb(NO_3)_{2(aq)} + CuSO_{4(aq)} \rightarrow Cu(NO_3)_{2(aq)} + \mathbf{PbSO_{4(s)}}$

Ionic equation:

$Pb^{2+}_{(aq)} + 2NO_3^{-}{}_{(aq)} + Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow Cu^{2+}_{(aq} + 2NO_3^{-}{}_{(aq)} + \mathbf{PbSO_{4(s)}}$

Net ionic equation:

$\mathbf{Pb^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4(s)}}$

  - **Precipitates of metal hydroxides (except group I hydroxides and ammonium compounds)**

Example:

copper(II) sulphate +potassium hydroxide → copper hydroxide + potas sium sulphate

$CuSO_{4(aq)} + 2KOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + K_2SO_{4(aq)}$

Ionic equation:

$Cu^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2K^{+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow \mathbf{Cu(OH)_{2(s)}} + 2K^{+}_{(aq} + SO_4^{2-}{}_{(aq)}$

Net ionic equation:

$$Cu^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Cu(OH)_{2(s)}$$

- **Decomposition** is a type of chemical reaction where one reactant yields two or more products. It is just an opposite of a synthesis or combination reaction. Most of decomposition reactions require heat.

  The general equation form is: **AB → A + B**

  Types of Decomposition Reaction

  - **Decomposition of metallic carbonate**

  Example:

  calcium carbonate → calcium oxide + carbon dioxide

  $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$

  - **Decomposition of metallic hydroxide**

  Example:

  calcium hydroxide → calcium oxide + water

  $Ca(OH)_{2(s)} \rightarrow CaO_{(s)} + H_2O_{(g)}$

  - **Decomposition of metallic hydrogen carbonate**

  Example:

  Sodium hydrogen carbonate → sodium oxide + water + carbon dioxide

  $2NaHCO_{3(s)} \rightarrow Na_2O_{(s)} + H_2O_{(g)} + 2CO_{2(g)}$

- **Combustion** is when a compound containing carbon, combines with the oxygen gas in the air, the reaction is called combustion. Heat is the most useful product of most combustion reactions.
- Three things are needed for a fire: fuel, heat, and oxygen. This is called the **fire triangle**. If one part is removed the fire goes out.
- Methane (CH_4), belongs to a class of compounds called **hydrocarbons**. The compounds composed only of carbon and hydrogen. The product of this reaction is heat.
- When there is enough oxygen supply during combustion, carbon dioxide, and water are produced. This is called **complete combustion.**
- If there is a limited supply of oxygen during combustion, carbon monoxide and water are produced. This is called **incomplete combustion**.

For example:
Complete Combustion of Methane:

methane + oxygen → carbon dioxide + water

$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

Incomplete combustion of methane:

methane + oxygen → carbon monoxide + water

$2CH_{4\,(g)} + 3O_{2\,(g)} \rightarrow 2CO_{\,(g)} + 4H_2O_{\,(g)}$

- **Exothermic and endothermic**
  - **Exothermic** is reaction which gives out heat. The temperature rises. The **product** formed is at **lower energy** level than that of the **reactants.** Hence, the excess energy is given out to the surroundings.

    Examples of exothermic reaction are adding concentrated acid to water, combustion reaction, neutralization, respiration, and dissolving solid sodium hydroxide and anhydrous copper sulfate in water.

  - **Endothermic** is reaction which takes in heat. The temperature falls. The **products** formed are at **high energy** level than that of the **reactants**; hence energy is absorbed by the reactants from the surroundings.

    Examples of endothermic reaction are decomposition of compounds such as calcium carbonate, taking a photograph with a film, photosynthesis and dissolving certain salts like ammonium nitrate and sodium carbonate crystals in water.

NOW YOU MUST COMPLETE ASSESSMENT 3 AND RETURN IT TO THE PROVINCIAL CENTRE CO-ORDINATOR.

ANSWERS TO LEARNING ACTIVITIES 1-8

Learning Activity 1

1. (i) If there is a significant change in the temperature
   (ii) If a solid or precipitate is formed
   (iv) If there is a change in colour
   (iv) If a gas is evolved or given off
   (v) If there is disappearance of solid

2. Any two from the list are correct answers
   (i) Hydrogen gas is given off.
   (ii) The zinc will disappear.
   (iii) The temperature is increased.

3. Silver chloride

4. (i) Insoluble
   (ii) Soluble
   (iii) Insoluble

Learning Activity 2

1. a. $\mathbf{2}Na_{(s)} + O_{2(g)} \rightarrow Na_2O_{(s)}$
   b. $\mathbf{2}Al_{(s)} + \mathbf{3}S_{(s)} \rightarrow Al_2S_{3(s)}$
   c. $\mathbf{2}K_{(s)} + F_{2(g)} \rightarrow \mathbf{2}KF_{(s)}$
   d. $C_{(s)} + CO_{2(g)} \rightarrow \mathbf{2}CO_{(g)}$
   e. $\mathbf{2}Cu_{(s)} + O_{2(g)} \rightarrow \mathbf{2}CuO_{(s)}$
   f. $\mathbf{2}Ag_{(s)} + Cl_{2(g)} \rightarrow \mathbf{2}AgCl_{(s)}$
   g. $\mathbf{2}K_{(s)} + Br_{2(g)} \rightarrow \mathbf{2}KBr_{(s)}$
   h. $H_{2(g)} + Cl_{2(g)} \rightarrow \mathbf{2}HCl_{(g)}$
   i. $N_{2\,(g)} + \mathbf{3}H_{2(g)} \rightarrow \mathbf{2}NH_{3(g)}$

2. (i) Iron sulphide
   (ii) Lithium chloride
   (iii) Calcium carbonate
   (iv) Sulphur dioxide
   (v) Magnesium oxide

Learning Activity 3

1. (i) Sodium chloride (NaCl)
   (ii) Copper nitrate ($Cu(NO_3)_2$)
   (iii) Barium sulphate ($BaSO_4$)

2. (i) $\mathbf{2}HCl + MgO \rightarrow \mathbf{MgCl_2} + \mathbf{H_2O}$
   (ii) $\mathbf{2}HNO_3 + Ca(OH)_2 \rightarrow \mathbf{Ca(NO_3)_2} + \mathbf{2H_2O}$
   (iii) $H_2SO_4 + \mathbf{2}KOH \rightarrow \mathbf{K_2SO_4} + \mathbf{2H_2O}$

3. (i) a. $HNO_3 + NH_4OH \rightarrow \mathbf{NH_4NO_3} + \mathbf{H_2O}$
   b. $H^+ + NO_3^- + NH_4^+ + OH^- \rightarrow NH_4^+ + NO_3^- + H_2O$
   c. $H^+ + OH^- \rightarrow H_2O$

   (ii) a. $H_2SO_4 + Na_2CO_3 \rightarrow \mathbf{Na_2SO_4} + \mathbf{H_2O} + \mathbf{CO_2}$
   b. $2H^+ + SO_4^{2-} + 2Na^+ + CO_3^{2-} \rightarrow 2Na^+ + SO_4^{2-} + H_2O + CO_2$
   c. $2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$

   (iii) a. $\mathbf{2}HCl + CaO \rightarrow \mathbf{CaCl_2} + \mathbf{H_2O}$
   b. $2H^+ + 2Cl^- + Ca^{2+} + O^{2-} \rightarrow Ca^{2+} + 2Cl^- + H_2O$
   c. $2H^+ + O^{2-} \rightarrow H_2O$

   (iv) a. $HNO_3 + NaOH \rightarrow \mathbf{NaNO_3} + \mathbf{H_2O}$
   b. $H^+ + NO_3^- + Na^+ + OH^- \rightarrow Na^+ + NO_3^- + H_2O$
   c. $H^+ + OH^- \rightarrow H_2O$

   (v) a. $H_2SO_4 + \mathbf{2}KHCO_3 \rightarrow \mathbf{K_2SO_4} + \mathbf{2H_2O} + \mathbf{2CO_2}$
   b. $2H^+ + SO_4^{2-} + 2K^+ + 2HCO_3^- \rightarrow 2K^+ + SO_4^{2-} + 2H_2O + 2CO_2$
   c. $2H^+ + 2HCO_3^- \rightarrow 2H_2O + 2CO_2$

4.

	Base	Acid	Name of Salt	Formula of Salt
(i)	Sodium hydroxide	Hydrochloric acid	**Sodium chloride**	**NaCl**
(ii)	Ammonium hydroxide	Hydrochloric acid	**Ammonium chloride**	**NH_4Cl**
(iii)	Iron(II)oxide	Sulphuric acid	**Iron(II) sulphate**	**$FeSO_4$**
(iv)	Copper(II)oxide	Sulphuric acid	**Copper(II) sulphate**	**$CuSO_4$**
(v)	Calcium hydroxide	Nitric acid	**Calcium nitrate**	**$Ca(NO_3)_2$**
(vi)	Zinc oxide	Nitric acid	**Zinc nitrate**	**$Zn(NO_3)_2$**

5. Alkalis are bases that can dissolve in water. Examples are sodium hydroxide and potassium hydroxide. Other examples are Group I hydroxides and ammonium hydroxide.

Bases are insoluble in water. They can react with acids to form a neutral salt and water. Examples are magnesium oxide and calcium oxide. Other examples are any metal oxides except for the oxides of Group I.

6. Neutralization reaction is a reaction between an acid and a base forming a neutral salt and water only.

Learning Activity 4

1 a. A sodium flouride and chlorine gas will form because fluorine is more reactive than chlorine and it will displace chlorine from sodium chloride.
b. A sodium chloride and bromine gas will form because chlorine is more reactive than bromine and it will displace bromine from sodium bromide.
c. A sodium bromide and iodine gas will form because bromine is more reactive than iodine and it will displace iodine from sodium iodide.
d. There is no reaction because iodine is less reactive than chlorine in sodium chloride.

2.

Reactant	Observation	Names of Products Formed
a. $KI_{(aq)} + Cl_{2(g)}$	**Chlorine will displace iodine from potassium iodide because chlorine is more reactive than iodine.**	**Potassium chloride and iodine molecule.**
b. $KBr_{(aq)} + Cl_2$	**Chlorine will displace bromine from potassium bromide because chlorine is more reactive than bromine.**	**Potassium chloride and bromine molecule.**
c. $KCl_{(aq)} + F_2$	**Fluorine will displace chlorine from potassium chloride because fluorine is more reactive than chlorine.**	**Potassium fluoride and chlorine molecule.**

3. a. $\mathbf{2}KI_{(aq)} + Cl_{2(g)} \rightarrow \mathbf{2}KCl_{(aq)} + I_{2(g)}$
b. $\mathbf{2}KBr_{(aq)} + Cl_{2(g)} \rightarrow \mathbf{2}KCl_{(aq)} + Br_{2\,(g)}$
c. $\mathbf{2}KCl_{(aq)} + F_{2(g)} \rightarrow \mathbf{2}KF_{(aq)} + Cl_{2(g)}$

4. a. $2Na_{(s)} + CaCl_{2(aq)} \rightarrow 2NaCl_{(aq)} + Ca_{(s)}$
(i) $Na_{(s)}$ is oxidized to $Na^{+}_{(aq)}$
(ii) $Ca^{2+}_{(aq)}$ is reduced to $Ca_{(s)}$

b. $K_{(s)} + AgNO_{3(aq)} \rightarrow KNO_{3(aq)} + Ag_{(s)}$
(i) $Na_{(s)}$ is oxidized to $Na^{+}_{(aq)}$
(ii) $Ag^{+}_{(aq)}$ is reduced to $Ag_{(s)}$

c. $Mg_{(s)} + CuSO_{4(aq)} \rightarrow MgSO_{4(aq)} + Cu_{(s)}$

(i) $Mg_{(s)}$ is oxidized to $Mg^{2+}_{(aq)}$

(ii) $Cu^{2+}_{(aq)}$ is reduced to $Cu_{(s)}$

5. a. $Na_2CO_{3(aq)} + CuSO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + CuCO_{3(s)}$

b. $(NH_4)_2SO_{4(aq)} + CaCl_{2\,(aq)} \rightarrow 2NH_4Cl_{\,(aq)} + CaSO_{4\,(s)}$

c. $K_2CO_{3(aq)} + Cu(NO_3)_{2\,(aq)} \rightarrow 2KNO_{3\,(aq)} + CuCO_{3(s)}$

d. $Mg(OH)_{2(aq)} + 2HNO_{3(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2H_2O_{(l)}$

6. (i) copper

(ii) Magnesium sulphate

(ii) $Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{\,(s)}$

7. A. √

B. χ

C. √

D. χ

E. √

8. Spectator ions are ions which do not take part in a chemical reaction.

Learning Activity 5

1. (i) Lead sulphate ($PbSO_4$)

(ii) Copper hydroxide,$Cu(OH)_2$

(iii) Silver chloride (AgCl)

2. (i) $FeCl_{2(aq)} + 2NaOH_{(aq)} \rightarrow Fe(OH)_{2(s)} + 2NaCl_{(aq)}$

(ii) $Fe^{2+}_{(q)} + 2\cancel{Cl}^{-}_{(aq)} + 2\cancel{Na}^{+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Fe(OH)_{2(s)} + 2\cancel{Na}^{+}_{(aq)} + 2\cancel{Cl}^{-}_{(aq)}$

(iii) $Fe^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Fe(OH)_{2(s)}$

3. (i) Barium nitrate ($Ba(NO_3)_2$) and copper sulphate ($CuSO_4$).
Any solution containing barium ions (Ba^{2+}), sulphate ions (SO_4^{2-}) that are soluble in water is possible to make a precipitate of barium sulphate, $BaSO_4$.

(ii) When you mix barium nitrate ($Ba(NO_3)_2$) and copper sulphate ($CuSO_4$) together, a precipitate of barium sulphate ($BaSO_4$) is produced and a solution of copper nitrate ($Cu(NO_3)_2$). Filtration method is used to collect the precipitate.

(iii) Barium nitrate + copper sulphate → barium sulphate + copper nitrate

(iv) $Ba(NO_3)_{2(aq)} + CuSO_{4(aq)} \rightarrow BaSO_{4(s)} + Cu(NO_3)_{2(aq)}$

Learning Activity 6

1. a. Word equation:
   Magnesium carbonate → magnesium oxide + carbon dioxide
   Symbol equation:
   $MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$

   b. Word equation:
   Copper hydroxide → copper oxide + water
   Symbol equation:
   $Cu(OH)_{2(s)} \rightarrow CuO_{(s)} + H_2O_{(l)}$

   c. Word equation:
   Potassium hydrogen carbonate → potassium oxide + water + Carbon dioxide
   Symbol equation:
   $2KHCO_{3(s)} \rightarrow K_2O_{(s)} + H_2O_{(l)} + 2CO_{2(g)}$

2. Balance the following decomposition reactions below. If the equation is balanced, write the word balanced next to the equation.

   (i) $Al_2(CO_3)_{3(s)} \rightarrow Al_2O_{3(s)} + \mathbf{3}CO_{2(g)}$
   (ii) $CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$ **balanced**
   (iii) $Fe_2(CO_3)_{3(s)} \rightarrow Fe_2O_{3(s)} + \mathbf{3}CO_{2(g)}$
   (iv) $Mg(HCO_3)_{2(s)} \rightarrow MgO_{(s)} + \mathbf{2}CO_{2(g)} + H_2O_{(l)}$
   (v) $Li_2CO_{3(s)} \rightarrow Li_2O_{(s)} + CO_{2(g)}$ **balanced**
   (vi) $\mathbf{2}Al(OH)_{3(s)} \rightarrow Al_2O_{3(s)} + \mathbf{3}H_2O_{(l)}$
   (vii) $Fe(OH)_{2(s)} \rightarrow FeO_{(s)} + H_2O_{(g)}$ **balanced**

3. Decomposition reaction is defined as the breaking down of single compound to two or more compounds or elements.

Learning Activity 7

1. Combustion is a reaction of fuel with oxygen to produce carbon dioxide and water. Sometimes, carbon monoxide is also produce and water if the combustion is incomplete. It is commonly called the process of burning.
2. The two types of combustion reaction are;
   (i) Complete combustion and
   (ii) Incomplete combustion
3. (i) $CH_{4\,(g)} + \mathbf{2}O_{2\,(g)} \rightarrow CO_{2\,(g)} + \mathbf{2}H_2O_{\,(g)}$
   (ii) $\mathbf{2}CH_{4\,(g)} + \mathbf{3}O_{2\,(g)} \rightarrow \mathbf{2}CO_{(g)} + \mathbf{4}H_2O_{\,(g)}$
4. (i) Fuel
   (ii) Heat
   (iii) Oxygen
5. $C_{(s)} + O_{2\,(g)} \rightarrow CO_{2\,(g)}$

6. If there is too much carbon dioxide in the atmosphere as a result of combustion process, the Earth will heat up resulting to a global warming.

Learning Activity 8

1. a. exothermic
   b. exothermic
   c. endothermic
   d. endothermic
   e. exothermic
   f. endothermic
   g. exothermic

2. a. Exothermic reaction is a reaction which releases heat to the surroundings.
   b. Endothermic reaction is a reaction which absorbs heat from the surroundings.

3. (i) exothermic
   (ii) endothermic
   (iii) lower
   (iv) higher
   (v) endothermic
   (vi) exothermic

REFERENCES

Elvins, C., D. and Jones, N. Lukins.1999. Heinmann Chemistry One. 3rd edition. Pp 162-164

Gallagher, R. anad P. Ingram. 1997. Chemistry Made Clear. GCSE edition. pp 64 – 67, 70 – 73, 132 – 136.

Leong, O. H., and C. L. Sian Dr. 2007. Chemistry Expression An inquiry Approach ' O ' Level Special. Express. Pp. 67 – 70, 247 – 251, 317 – 320, 362 – 364.

Marasinghe, B. Dr. 2010. Upper Secondary Chemistry. A Textbook of Chemistry for Grades 11 and 12. (New Syllabus in Papua New Guinea). pp.40 – 50, 53.

Ramsden, E. N. 1986. Chemistry for GCSE. 2nd edition. pp. 89 – 90, 99 – 102, 293 – 295.

Regan, S., L. Bramley, T. Gemellano, D. Wilson and S. Wiseman. 2004. Chemistry A Contextual Approach. Heinemann Queensland Science project.pp 339 – 342, 349 – 350, 404-405.

Ryan, L. 1996. Chemistry For You. National Curriculum Edition for GCSE. Stanley Thornes (Publishers) Ltd. Pp. 89 – 90, 99 – 102.

books.google.co.uk
www.ric.edu/faculty

www.britannica.com

chemistry.about.com

antoine.frostburg.edu

www.wikipedia.org/wiki

chemistry.tutorvista.com

www.Chemistryexplained

www. horton. ednet.ns .ca

www.educationscotland.gov.uk