



# Chemistry for Year 11



Curriculum Development Unit  
Ministry of Education  
2014



# **CHEMISTRY FOR YEAR 11**

**Curriculum Development Unit**

**Ministry of Education**

**Fiji**

**2014**

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**CDU**

# PREFACE

Welcome to **Chemistry for Year 11**.

This book was prepared as an introduction to the broad field of Chemistry and related activities. It covers the syllabus of Year 11 Chemistry which was formulated in 2014.

The syllabus is the framework for the learning and teaching of Chemistry while this textbook is a resource for students and teachers, it is hoped that it will be useful in implementing the syllabus and it is to be used to **complement lessons prepared by teachers**.

Students and teachers are encouraged to use other resource materials as well.

Suggestions for amendments are welcome.

MINISTRY OF EDUCATION

SUVA.

01<sup>st</sup> September, 2014

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# STRAND 1                      GENERAL CHEMISTRY

**Strand Outcome:** Appreciate the science of chemistry and demonstrate knowledge and skills on the use of scientific apparatus to carry out scientific investigations.

## Sub-strands

- 1.1 Introduction to Chemistry
- 1.2 Fundamental Tools of Chemistry

## 1.1 Introduction to Chemistry

### Achievement Indicators

Upon completion of this sub-strand, students will be able:

- ✓ Find and describe the study of Chemistry as a science subject.
- ✓ Research and present some chemical based industries in Fiji and their contribution to the development of the country.
- ✓ Determine and explain on the career opportunities that require the knowledge of Chemistry.

### Study of Chemistry

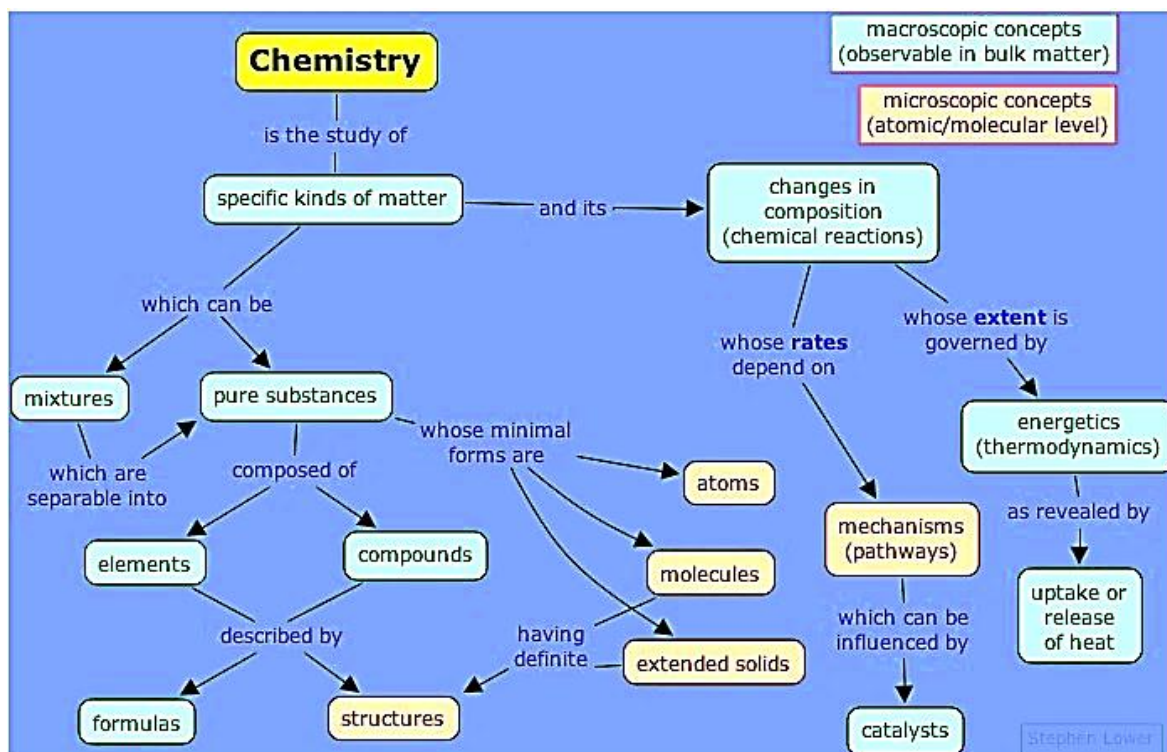
Chemistry is the study of matter and its interactions with other matter and energy. The basics of how things work is the study of chemistry. In a nutshell, the importance of chemistry is that it explains the world around us.

Chemistry explains how food changes during cooking, how it rots, how to preserve food, how the body uses the food that is eaten and how ingredients interact to make food.

Chemistry is used to help decide what cleaner is best for dishes, laundry and ourselves. Chemistry is used to select bleaches, disinfectants and even ordinary soap to use.

Basic chemistry is needed to understand how vitamins, supplements and drugs can help or harm the human body.

## Overview of Chemistry



Source: kbscmstsmkktj.blogspot.com

## Chemistry Relates to Human Life

Chemistry is important because everything one does is chemistry. Even the human body is made of chemicals. Chemical reactions occur when one breathes, eat, or just sit there reading.

It is important to understand chemistry if a student is studying any of the sciences because all of the sciences involve matter and the interactions between types of matter.

## Common chemicals used daily

1. Food
2. Air
3. Soap
4. Emotions
5. Any object that can be seen or touched.

## Examples of chemical-based industries in Fiji

1. South Pacific Fertilizers
2. Standard Concrete
3. Food Industries
4. Sugar Industry

## Some importance of chemical-based industries:

1. Responsible for the manufacture of a vast range of useful products which helps in improving our health and standard of living.
2. Provide job opportunities.
3. Support other industries example, automobile.

## Careers in Chemistry

**Agricultural Scientist**



**Biochemist**



**Geochemist**



**Industrial Chemist**



**Science Teacher**



**Nurse**



**Medical Doctor**



**Pharmacist**



### Exercise 1.1.1

Complete the table below by pasting pictures of the 8 Chemistry related careers given. More can be added to the table.

<b>Chemist</b>	<b>Marine Scientist</b>
<b>Chemical Analyst</b>	<b>Chemical Engineer</b>
<b>Ecologist</b>	<b>Electrochemist</b>
<b>Environmental Chemist</b>	<b>Food Technologist</b>

## 1.2 Fundamental Tools of Chemistry




### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ List and describe the names, correct uses and care in handling some basic apparatus.
- ✓ Outline and talk about the importance of safety rules in the laboratory for handling and storing chemicals as well as the cleaning of equipment after use.
- ✓ Determine and write the correct units for measurements from the apparatus and use them in calculations with correct significant figures.
- ✓ Study and design hypothesis, aim and inference for experiments.
- ✓ Conduct experiments using variables and controls and communicate by compiling scientific reports correctly.

### Basic Laboratory Equipment

Given below are the pictures, names and uses of some common laboratory equipment used in Chemistry.

<b>Beakers</b>	<b>Measuring cylinder</b>	<b>Conical flask</b>
		
Used for stirring, mixing or heating large volumes.	Used for measuring large volumes.	Used for collecting and mixing liquid.

<p><b>Volumetric (standard) flask</b></p>  <p>Used for measuring large volumes accurately.</p>	<p><b>Burette</b></p>  <p>Used for measuring liquid during titration.</p>	<p><b>Pipette</b></p>  <p>Used for transferring and measuring small volumes accurately.</p>
<p><b>Digital Balance</b></p>  <p>Used for measuring mass of chemicals.</p>	<p><b>Dial o gram</b></p>  <p>Used for measuring mass of chemicals.</p>	<p><b>Triple beam balance</b></p>  <p>Used for measuring mass of chemicals.</p>
<p><b>Spring balance</b></p>  <p>Used for measuring mass of chemicals.</p>	<p><b>Gas syringe</b></p>  <p>Used for measuring the volume of gas.</p>	<p><b>Thermometer</b></p>  <p>Used for measuring temperature of chemicals.</p>
<p><b>Stop watch</b></p>  <p>Used for measuring time.</p>	<p><b>Barometer</b></p>  <p>Used for measuring gas pressure.</p>	<p><b>Meter ruler</b></p>  <p>Used for measuring dimensions.</p>
<p><b>Ammeter</b></p>  <p>Used for measuring current.</p>	<p><b>Bunsen burner</b></p>  <p>Used for heating chemicals.</p>	<p><b>Funnels</b></p>  <p>Used for filtering chemicals.</p>

**Test tube holder**

Used for holding test tubes during an experiment.

**Goggles**

Used to protect eyes when carrying out experiments.

**Spatula**

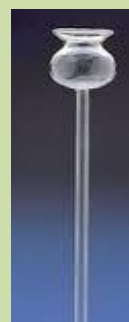
Used for transferring chemicals from a stock sample.

**Crucible**

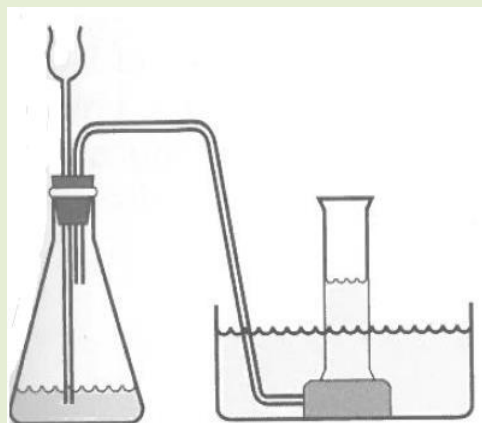
Used for heating solids strongly.

**Separating funnel**

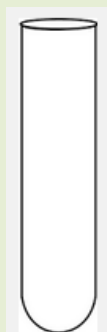
Used for separating two immiscible solvents.

**Thistle funnel**

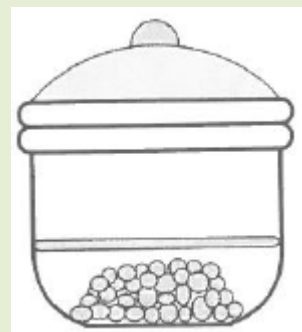
Used for pouring liquids into flasks.

**Bee hive shelf and gas jar**

Used for collecting gases over water.



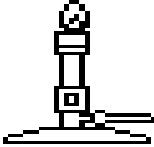





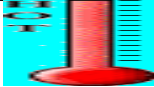
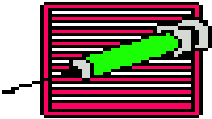
**Boiling tube**

Used for heating small volumes.

**Desiccator**

Used for drying chemicals.

## Safety Rules in Laboratory

1	<p><b>No eating or drinking in the lab.</b></p> <p>This means no chewing gum, sweets, chewing hair ends, holding a pencil in your mouth, nail-biting, etc.</p>	
2	<p><b>Handle everything as if it's pathogenic (disease causing).</b></p>	
3	<p><b>Keep flame and flammable solutions far apart.</b></p> <p>Set up the laboratory so that if a flame is to be used, it is located far from the exit, so that most students are closer to the exit. Have any open alcohol beakers far from the flame - for instance on another workbench.</p>	
4	<p><b>Keep electrical equipment far from water.</b></p>	
5	<p><b>Clean any spills immediately.</b></p>	
6	<p><b>Always wear proper safety protection gears such as goggles, gloves, close shoes, lab coat, and nose mask.</b></p>	
7	<p><b>Always clean glassware before it is used.</b></p> <p>This ensures that impurities are cleaned away. Clean using some distilled water first, before adding any liquid or solid to the glassware.</p>	
8	<p>Be careful when weighing out chemicals and reagents.</p> <p><b>Do NOT return excess materials to the stock container.</b></p>	
9	<p><b>Check all water baths with a thermometer before</b> putting your hand into the water.</p>	
10	<p><b>All sharp objects (such as needles, razors, pins and toothpicks) should be discarded in a closed container.</b></p>	

## Measurements and Calculations

In the world of chemistry, measurement is used during weighing out dry chemicals, noting the amount of a liquid used during titration or measuring the temperature of a liquid during a chemical reaction. Hence, measurements can be taken on characteristics ranging from temperature to length to weight. In the laboratory classes, the most frequently made measurements include mass and volume.

## Quantities and units of measurement

Given in the table below are some quantities, their symbols and the units of measurement.

Quantity	Symbol	Name of Unit	Abbreviated form of unit
Length	<i>l</i>	metres	m
Mass	m	grams, kilograms	g, kg
Time	t	seconds	sec
Temperature	T	degrees Celsius, Kelvin	°C, K
Volume	V	millilitres, litres, cubic metres	mL, L, m <sup>3</sup>
Pressure	P	Pascals	Pa
Area	A	square metres	m <sup>2</sup>
Concentration	c	grams per litre, mol per litre	g L <sup>-1</sup> , mol L <sup>-1</sup>
Current	I	Amperes	A
Energy	E	Joules	J
Weight	w	Newtons	N
Electric charge	q or Q	Coulombs	C

**Note:** All scientific measurements are carried out in the International System of units (SI), which is the modern form of the metric system and is the most widely used system of measurement.

### Significant Figures (S.F.)

In the laboratory, most numbers are obtained due to a result of measurements. Any number resulting from measurement is not an absolute or exact whole number. For example, when you use a thermometer to measure the air temperature, you cannot know what the absolute temperature is. Instead, you can measure the air temperature to the nearest degree for e.g. 28.5 °C.

**Significant figures** are the digits in a number which are known precisely, plus one estimated digit.

The number of significant figures indicates the precision or reliability of a measurement.

#### Rules to identify the significant figures:

1. ALL non-zero numbers (1, 2, 3, 4, 5, 6, 7, 8 and 9) are ALWAYS significant.
2. ALL zeroes between non-zero numbers are ALWAYS significant.
3. ALL zeroes which are SIMULTANEOUSLY to the right of the decimal point AND at the end of the number are ALWAYS significant.
4. ALL zeroes which are to the left of a written decimal point and are in a number  $>$  or  $=$  10 are ALWAYS significant.

**NB: A helpful way to check rules 3 and 4 is to write the number in scientific notation. If you can get rid of the zeroes, then they are NOT significant.**

**Example:** How many significant figures are present in the following numbers?

Number	Number of significant figures	Rules Applied
37,869	5	1
4.978	4	1
800.07	5	1,2,4
0.0003 ( $3 \times 10^{-4}$ )	1	1,4
6.1000	5	1,3
401.060	6	1,2,3,4
2,000,000 ( $2 \times 10^6$ )	1	1
20.0 ( $2.00 \times 10^1$ )	3	1,3,4

### Exercise 1.2.1

1. Indicate the number of significant figures in the following measured numbers:

- a. 246.32                      b. 107.854                      c. 100.3  
d. 0.678                        e. 1.008                        f. 0.00340

2. Round off the following numbers to the number of significant figures given in brackets.

- a. 3.580 (2)                      b. 2.2400 (4)                      c. 0.04792 (2)  
d. 333.0 (2)                      e.  $2.463 \times 10^{23}$  (1)                      f. 6.643 (3)  
g. 3.450 (2)                      h. 3529 (1)                      i. 342.0 (2)  
j. 0.04791 (2)                      k.  $6.9342 \times 10^{-1}$  (3)                      l. 37.852 (3)

### Calculating with Significant Figures

Refer to the examples on rounding off numbers before you learn how to apply the rules on significant figures when performing calculations.

**Example 1** - Suppose you wish to round 52.5347 to *four significant figures*. Look at the fifth figure. It is a 4, a number less than 5. Therefore, you will simply drop every figure after the fourth, and the original number rounds off to 52.53.

**Example 2** - Round 2.68721 to *three significant figures*. Look at the fourth figure. It is 7, a number greater than 5, so you round the original number up to 2.69.

**Example 3** - Round 639.835 to *five significant figures*. Look at the sixth figure. It is a 5, so now you must look at the fifth figure also. That is a 3, which is an odd number, so you round the original number up to 639.84.

When performing calculations, you will also have to be aware of significant figures. The following rules are used:

- **Addition or subtraction**

The number of digits to the right of the decimal point in the answer is equal to the number of digits to the right of the decimal point in the number with the **least such digits**.

**Example**

**Addition**

$$\begin{array}{r} 47.35 \\ + 1.587 \\ \hline 48.937 \end{array}$$

48.937 rounds off to 48.94

**Subtraction**

$$\begin{array}{r} 1.6 \\ - 0.1423 \\ \hline 1.4577 \end{array}$$

1.4577 rounds off to 1.5

- **Multiplication or division**

The answer should have the same number of significant figures as the number with the fewest significant figures.

**Example:** Multiplication

$$34 \times 43.21123 = 1469.18182 \text{ this rounds off to } 1500 \text{ or } 1.5 \times 10^3.$$

**Example:** Division

$$8197.4 \div 4 = 2,049.35 \text{ which rounds off to } 2,000 \text{ or } 2 \times 10^3.$$

**Exercise 1.2.2**

1. Calculate the answers to the appropriate number of significant figures.  
a.  $32.567 + 135.0 + 1.4567$       b.  $378 - 67$       c.  $23.7 \times 3.8$       d.  $28.367 \div 3.74$
2. Convert the measurement 101000 g to scientific notation.
3. Circle the letter of the best answer to the problem:  $123000 \text{ m} \times 3234 \text{ m}$   
A.  $3.97 \times 10^{-7} \text{ m}^2$   
B.  $398 \text{ m}^2$   
C.  $39800000 \text{ m}^2$   
D.  $3.98 \times 10^8 \text{ m}^2$
4. When performing the calculation  $44.540 \text{ g} + 11.1 \text{ g} + 1722.84 \text{ g}$ , the final answer must have  
A. one decimal place.  
B. the units of  $\text{g}^3$ .  
C. three significant figures.  
D. two decimal places.
5. 21.4 mL of a solution is best measured with a  
A. standard flask.  
B. graduated cylinder.  
C. pipette.  
D. burette.

6. 923 g divided by 20312 cm<sup>3</sup> to the correct number of significant figure(s) is

- A. 0.0454 g/cm<sup>3</sup>.
- B.  $4.00 \times 10^{-2}$  g/cm<sup>3</sup>.
- C. 0.04 g/cm<sup>3</sup>.
- D. 0.045 g/cm<sup>3</sup>.

7. What is the value of  $9.42 \div (5.24 \times 9.5)$ ?

- A. 0.18
- B. 0.2
- C. 0.1816
- D. 0.181

8. Round 0.060185 to three significant figures.

- A. 0.06
- B. 0.0601
- C. 0.0602
- D. 0.06018

9. What is the answer of the following calculation to the correct number of significant figures?

$$(8.41 \times 2.873) + 4.77$$

- A. 29.0
- B. 28.97
- C. 28.9
- D. 28.93

10. Do the following calculation and give your answer to the correct number of significant figure (s).

$$6.0 \times 6.0$$

10. How many significant figure(s) should the answer to these problems have?

i.  $2.5 \times 3.42$

ii.  $2.33 \times 6.085 \times 2.1$

iii.  $121.4 + 1001.23$

iv.  $879.53 - 24$

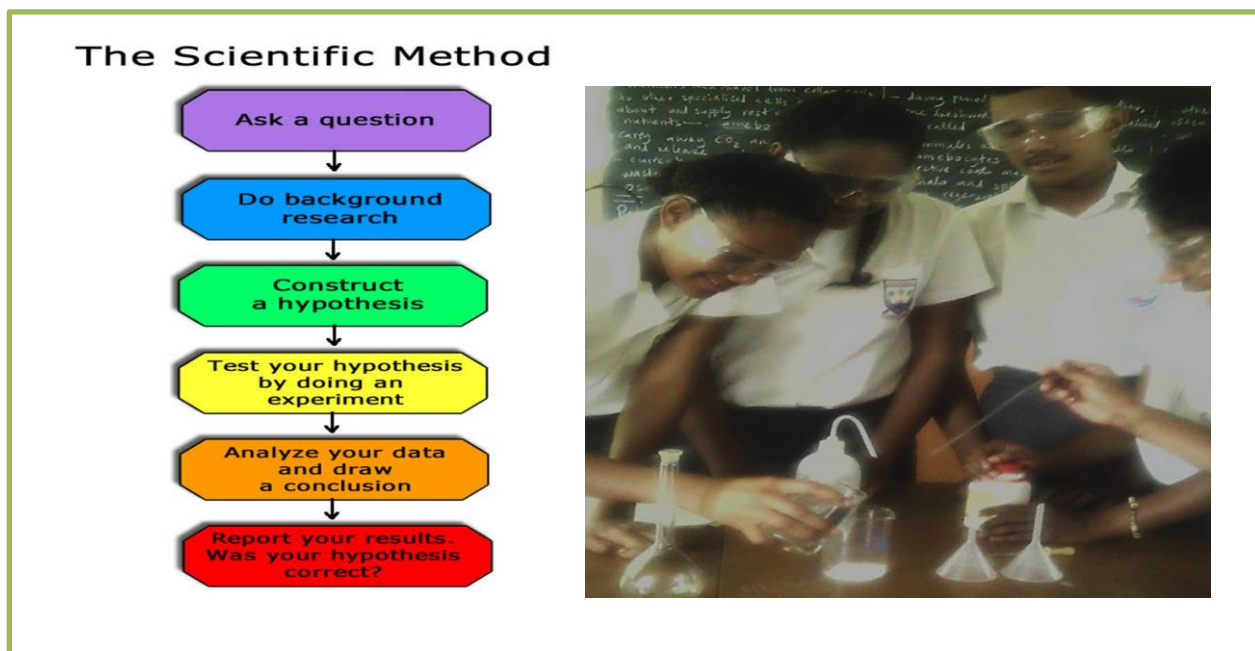
v.  $(3.4517 \times 10^7) \times (5.51 \times 10^{-4})$

vi.  $(3.42 \times 10^{-4}) \div (4.780 \times 10^{-6})$

11. Density can be calculated by mass (g) divided by volume (cm<sup>3</sup>). In an experiment, a student was required to measure the mass of a piece of iron. He was not able to find a weighing balance. However, he knew that the volume of the piece of iron is 27 cm<sup>3</sup>. If the density of iron is 7.87 g/cm<sup>3</sup>, calculate the mass of the iron to the correct number of significant figures.

## The Scientific Method

The Scientific Method involves a series of steps that are used to investigate a natural occurrence.



The following are important terms used during an investigation:

- ❖ **Hypothesis**- is an idea or explanation that you test through study and experimentation. Outside science, a theory or guess can also be called a *hypothesis*.
- ❖ **Experiment** – is a way of testing a hypothesis.
- ❖ **Control experiment** – the experiment in which the variable is not changed. It serves as the standard of comparison.
- ❖ **Variable**- is the thing that is changed in an experiment.
- ❖ **Observation** – noting and recording something with or without instruments.
- ❖ **Inference** – a conclusion that attempts to explain or make sense of an observation. In other words it is a conclusion reached on the basis of evidence and reasoning.

### A Sample Practical Report

The following gives a general outline on how to report a chemistry investigation to the teacher:

1. **Date, Experiment Number, Experiment Title**
2. **Aim:** gives the purpose of the experiment to the reader.
3. **Introduction:** gives an explanation for the reason(s) of conducting the experiment and what is the plan for conducting the experiment.
4. **Materials:** list of chemicals and glassware used during the experiment.
5. **Procedure/Method:** gives information on what was done in the experiment. It is written in past tense and in third person narration.
6. **Results/Observation:** includes both qualitative and quantitative observation e.g, data obtained, a graph, a diagram and the description of the result.

**Qualitative:** these are descriptions of the result obtained.

**Quantitative:** these are exact measurements e.g. mass, time.

**Graph drawing** is an important skill. The following questions should be a reminder of how to draw a graph.

a. **What type of graph do I need?**

To choose which type of graph to use, consider the two sets of data to be represented:

- i) If both sets of data consist of numerical values, a *line graph* is usually the appropriate choice.
- ii) If one set of data is categorical and therefore consists of words (for example *name of metal*) then a *bar chart* is usually the best choice.
- iii) *Pie charts* are used to represent proportions, for example the amount of each metal found in an alloy.

b. **Which variable goes on the x axis?**

It is important to remind pupils which axis is which, for example by stating “x is a cross (across) so the axis across is the x-axis (across)”. In an experiment, the **independent variable** (the thing you changed) always goes on the x axis and the dependent variable (the thing you measured) goes on the y axis.

If pupils find it hard to remember which is the independent variable, one way to work it out which works for most practicals is to ask pupils which column of the results table they can fill in before the experiment (the independent variable) and for which they have to wait until the experiment is complete (the dependent variable).

c. **What scale should I use?**

The scale should look like the markings on a ruler, with evenly ascending numbers. Identify the minimum and maximum measurements from their data, and work out how best to represent the range on the graph paper. Include labels on both the x axis and the y axis, and where appropriate also include the units used for measurements.

d. **How do I plot the graph?**

Draw an example of a graph. Demonstrate how to find the correct places on both axes and combine that information to locate the exact position of each point. Use a pencil to plot points, and to use an x rather than a dot for maximum accuracy.

7. **Discussion:** these are answers to the questions in the laboratory manual and may also include explanation to the results obtained.
8. **Conclusion:** shows what has been learnt from the experiment. This is a summary of the Aim/Objectives in 2-3 sentences.
9. **Reference:** the various sources used to get information for the experiment and to do the lab write-up are cited using the American Chemical Society (ACS) method.

**Exercise 1.2.3**

1. Select and draw a suitable type of graph for the three sets of data given below. Use a graph paper.
  - a. Solubility of  $\text{NH}_4\text{Br}$  in  $\text{H}_2\text{O}$

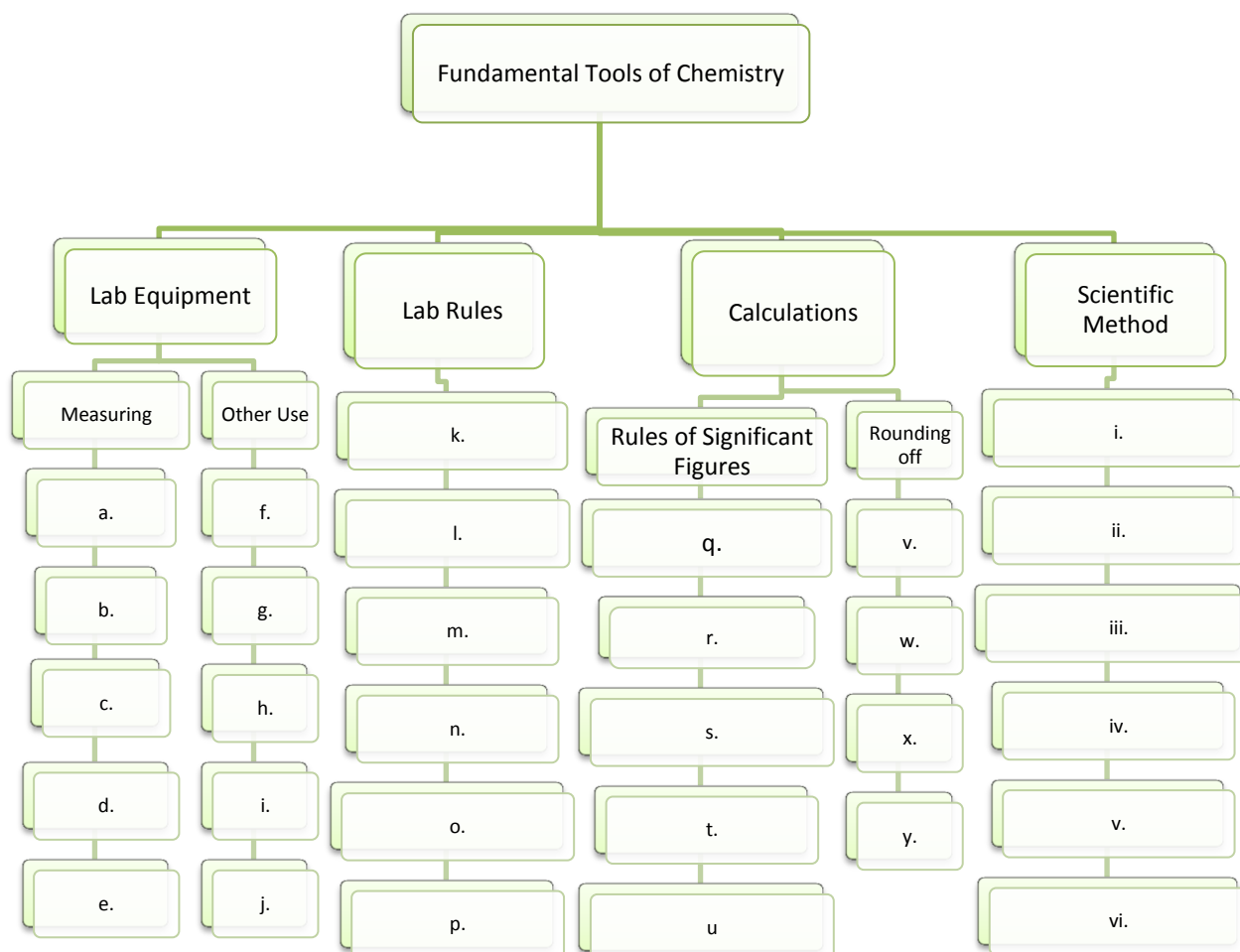
Temperature ( °C)	Mass of $\text{NH}_4\text{Br}$ per 100 g of $\text{H}_2\text{O}$ (g)
0	60
20	75
40	90
60	105
80	120
100	135

- b. Percentage of elements in the earth's crust.      c. The density of some metals:

Element	%
Oxygen	49.9
Silicon	26.0
Aluminium	7.0
Iron	4.0
Calcium	3.0
Potassium	2.5
Sodium	2.5
Magnesium	2.0
Hydrogen	1.0
Others	2.0

Element	Density (g/cm <sup>3</sup> )
Sodium	0.97
Calcium	1.53
Magnesium	1.74
Aluminium	2.70
Zinc	7.14
Iron	7.87
Lead	11.3
Copper	8.93

2. Complete the Concept Map below and present to the class. You can add more items to the concepts given.



# STRAND 2 INVESTIGATING MATTER

**Strand Outcome:** Describe the changes and explain the differences in structures and properties of the types of matter.

## Sub-strands

- 2.1 States of Matter
- 2.2 Atomic Structure and Bonding
- 2.3 Elements, Compounds and Mixtures

## 2.1 States of Matter

### Achievement Indicators

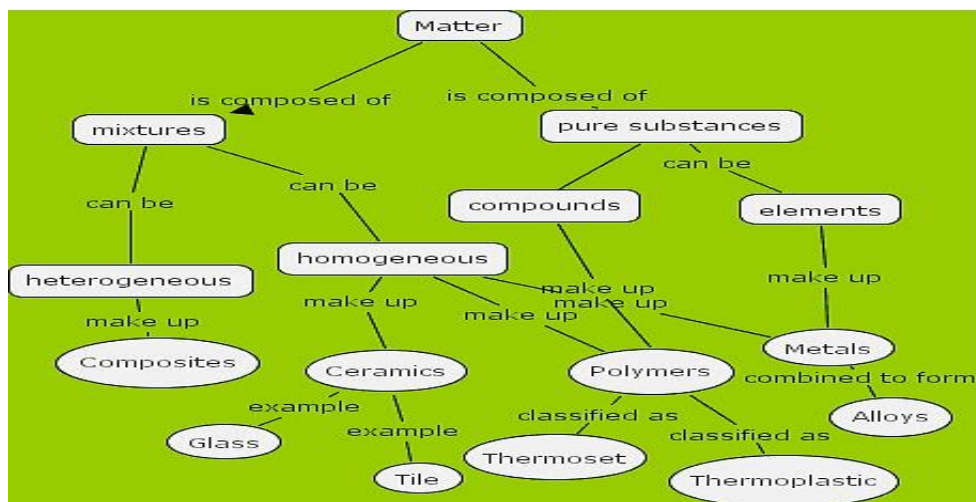
Upon completion of this sub-strand, students will be able to:

- ✓ Describe and illustrate the change in state of matter when adding and removing energy using the particle model.
- ✓ Formulate and explain the heating and cooling curves of ice (water vapour) and naphthalene.
- ✓ Measure and compare the densities of solids, liquids and solutions.
- ✓ Compare and explain the relationship between temperature and density of liquids and solutions.
- ✓ Evaluate and discuss the relationship between density and concentration of solutions.
- ✓ State and describe the diffusion of gas using gas particles behaviour.
- ✓ Show and explain gas pressure using the particle model when temperature or pressure changes.

Matter is anything that takes up space and has mass. Everything around us is made up of matter. All matter exists in three states: solid, liquid and gas. These three states can be clearly understood when considering the substance, water.

The solid form of water is ice, the liquid form of water is water that we drink and the gaseous form of water is water vapour. This unit looks at what happens to substances when they are heated or cooled causing them to move or change from one state to another.

## Concept map on Matter:



## The Particle Model of Matter

Matter is made up of tiny particles called atoms. These particles are so small that they are invisible. In order to study the behaviour of matter, scientists thought up a way to explain it using a model. This model, called the **Particle Model of Matter** helps us to understand what matter is and the way matter behaves.

The Particle Model of matter simply states three things about matter:

1. Matter is made up of very small particles.
2. Particles in a matter are constantly moving.
3. The higher the temperature, the faster the particles will move.

## States of Matter

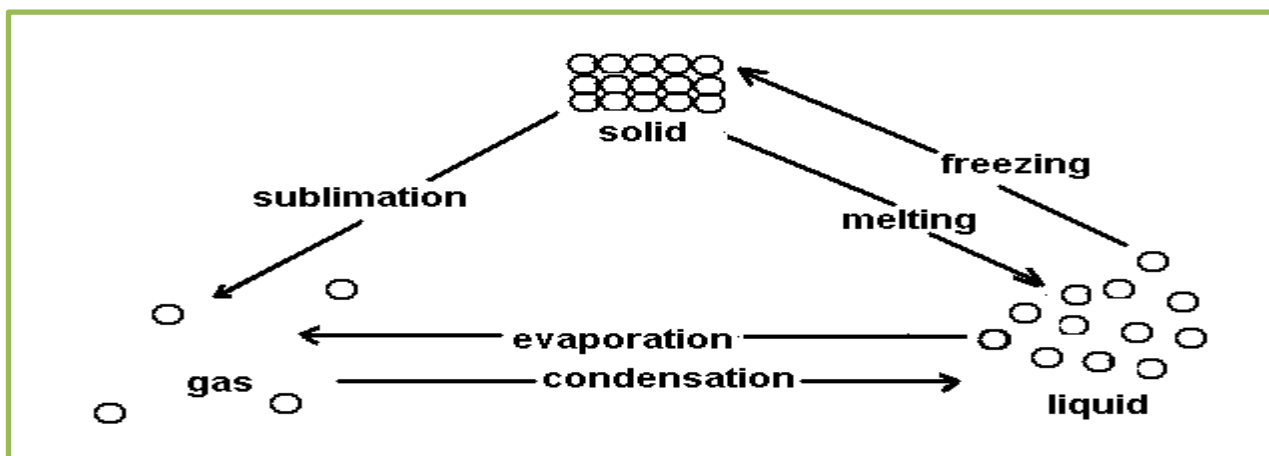
The particle model helps us to understand what matter looks like in the three states. It also helps us to understand the properties of the substance that are in a certain state.

This is discussed in greater detail in the table given below:

State of Matter	Property of State	Diagram
<b>Solid</b>	<p>Particles are held very closely together.                      Particles are fixed in a certain position and vibrate in that position.                      The amount of vibration in a fixed position depends on the temperature: the higher the temperature, the faster the vibration.</p> <p><b>Result of properties</b>                      Solids are hard</p>	
<b>Liquid</b>	<p>Particles are close together but are not set in one place.                      Particles in a liquid may freely move around and bump into each other.                      The higher the temperature, the faster the movement.</p> <p><b>Result of Properties</b>                      Liquids flow out of vessels and take up the shape of the container they are in.</p>	
<b>Gas</b>	<p>The particles are far apart and free to move;                      Particles in gases travel a longer distance before bumping into each other because they are far apart;                      The hotter the gas, the faster the motion.</p>	

## Changes of State

When water is placed in a freezer, the liquid water becomes ice. It is said that the liquid water changes from liquid state to solid state. This means that water has changed from one state of matter to another. When changes of state are physical changes, there is no new substance formed. The figure below shows some changes of state:



Thus, when water changes from liquid to solid (ice), how can this be explained using the particle model of matter?

*In liquid water, the particles are close together and able to glide over each other. When the particles are cooled, they lose energy and move more closer together until they form solid water or ice.*

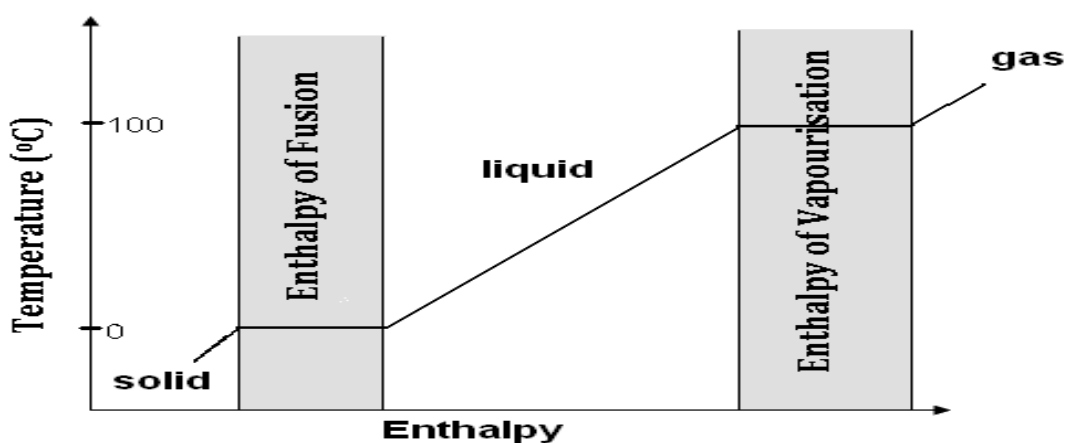
## The Heating Curve of Ice

The property of a substance depends on what state of matter the substance is in. For instance, when a substance is solid, it is hard, when it is liquid it is soft and when it is in gas form, it can push a piston to power turbines when compressed.

In this section, we will look at what happens to the particles of matter when it is changing from one state to another. In order to understand this, we must look at energy. For instance, a piece of wood, when burning can produce enough energy to warm a test tube of water. This heat energy is given a special name called **enthalpy** which is the energy contained in a substance whether kinetic or potential. In the case of water, ice has the lowest enthalpy and water vapour has the highest enthalpy.

When ice is heated, or given more energy, a graph can be drawn to show the changes in the temperature of water as more temperature is added to the water.

### The Heating Curve of Water



The areas on the curve that appear flat indicate that energy is being used by the particles to change state from solid to liquid. This energy is going into the particles converting them to the next state of matter.

### Example

When water is heated to 100°C, it will begin to boil. The temperature will not rise beyond 100°C even if more heat is added to it. So where is all the heat energy going? It is going to the particles in the liquid form, causing them to move further apart and become gas. The energy that is causing this change of state from liquid to gas is called the **Enthalpy of Vaporisation**.

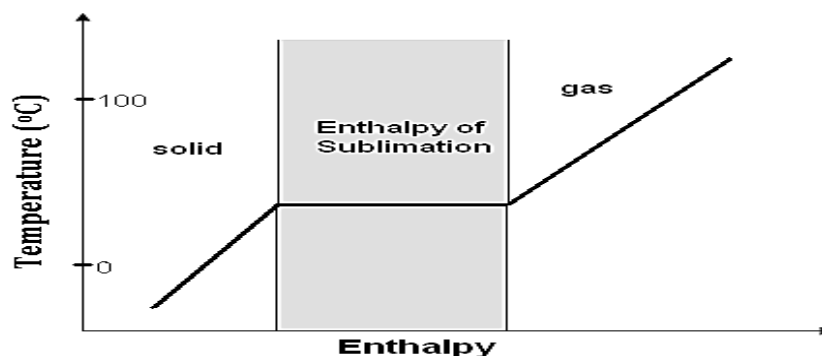
In the case of ice melting, the energy is causing the particles to move from their fixed position in the solid state, to be further apart and become liquid. This energy is called the **Enthalpy of Fusion**.

The changes discussed above are reversible changes. Thus, when water gains enthalpy, it vaporises and when it loses enthalpy, it condenses.

### The Heating Curve of Naphthalene

Naphthalene has only two states – solid and gas. This is because naphthalene **sublimes** (changes directly from solid to gas).

#### The Heating Curve of Naphthalene



The place on the graph that appears flat is where enthalpy increases but temperature does not. This means that heat is being added but no apparent change is seen.

So where is this heat energy going to?

*It is going to the naphthalene particles causing them to move further apart from each other and become gas.*

This change in enthalpy is called **Enthalpy of Sublimation**.

## Density

The density of an object is said to be how much matter is packed in that object. The more compactly matter is packed within an object, the higher the density of the object.

In solids, the particles are packed very closely together compared to particles in the liquid. Thus, solids are denser than liquids. This is evident when solids are placed in a liquid. They sink.

### Example

Consider a block of wood and a block of metal of the same size. When weighed in your hand, the block of metal would weigh more. If the two blocks were examined more closely, it will be found that the piece of wood has small air spaces within it while the particles in the metal block are more closely packed. This is why the block of metal is denser than the block of wood.

## Calculating Density

The density of a substance is calculated by dividing its mass by its volume.

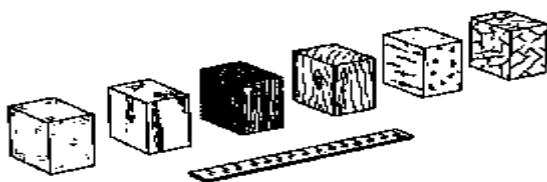
$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

The unit of density is grams per cubic centimetre ( $\text{g}/\text{cm}^3$ ) or grams per millilitres ( $\text{g}/\text{ml}$ ).

## Density of Solids

### 1. Regular solid

Measure the three dimensions of each regular object, repeating each measurement at two or more places. Depending on the size of the measurement, use the ruler, vernier callipers or a micrometer.



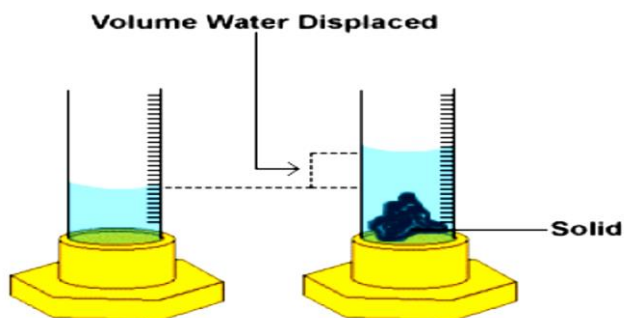
- Calculate the volume,  $V$ , of the object.
- Measure the mass,  $m$ , of the object, using the balance.
- Calculate the density,  $\rho$ , of each sample, using  $\rho = m/V$ .

### 2. Irregular Solids

- Place the irregular solid on the scale to find its mass. Note down the mass.
- Find a beaker/measuring cylinder that can contain your irregular solid. You can also use other containers that measure volume.
- Fill the beaker with enough liquid (such as water) that the irregular solid could be completely submerged.

Write down the volume of the liquid.

- iv. Place the irregular solid into the beaker and record the new volume.



- v. Subtract the new volume from the original volume to find the volume of the solid.

For example, if the original volume equals 40 ml and the new volume equals 65 ml, subtract 40 from 65 to get 15 ml as the volume of the solid.

Note: Each ml equals 1 cubic centimeter ( $\text{cm}^3$ ), so you can change the unit without changing the number. In this example, 15 ml becomes  $15 \text{ cm}^3$ .

Volume of water displaced = Final volume – Initial Volume

Volume of water displaced = Volume of solid

NB:  $1 \text{ ml} = 1 \text{ cm}^3$

- vi. Divide the irregular solid's mass by the irregular solid's volume to find its density.

For example, if your irregular solid weighs 45 g, divide 45 g by  $15 \text{ cm}^3$  to get a density of  $3 \text{ g/cm}^3$ .

## Density of Liquids

**Problem:** You are given an unknown liquid. Find the density.

**Materials:** 100 ml graduated cylinder, triple beam balance, calculator, unknown liquid.

### Procedure:

- 1) Find the mass of the empty graduated cylinder.
- 2) Pour unknown liquid 1 into the graduated cylinder to the 50 ml level.
- 3) Find the mass of the graduated cylinder with 50 ml of unknown liquid.

We can calculate density of a liquid using the formula:

**Density = Mass/Volume**

Where mass is that for just the liquid (you must subtract out the mass of the graduated cylinder).

## Class Activity

Given:

Mass of empty graduated cylinder = **78 g**

Mass of graduated cylinder with unknown liquid = **117.5 g**

### Find:

- a) Mass of the unknown liquid.
- b) Volume of the unknown liquid.
- c) Density of the unknown liquid.

### Exercise 2.1.1

1. A rock has a mass of 210 g and occupies a volume of 70 cm<sup>3</sup>. What is its density?
2. An unknown liquid occupies a volume of 5 mL and has a mass of 40 g. Find its density.
3. A rock has a density of 4 g cm<sup>-3</sup> and a mass of 16 g. What volume does this rock occupy?
4. An unknown substance from planet X has a density of 10 g cm<sup>-3</sup>. It occupies a volume of 80 mL. What is the mass of this unknown substance?
5. Pure water has a density of 1.0 g cm<sup>-3</sup> and ocean water has a density of 1.025 g cm<sup>-3</sup>. Why are the densities different?

### Effect of Temperature on the Density of a Substance

- When a liquid or gas is heated, the molecules move faster, bump into each other, and spread apart. Because the molecules are spread apart, they take up more space, thus becoming less dense.
- Cooling a substance causes molecules to slow down and get slightly closer together, occupying a smaller volume that results in an increase in density.
- Therefore temperature can affect density.
- Hot water is less dense and will float on room-temperature water.
- Cold water is denser and will sink in room-temperature water.

**As temperature increases, the density of liquids and gases decreases; as temperature decreases, the density increases.**

### Effect of Concentration on the Density of a Substance

Adding more solute to a solvent changes the composition of particles in a given volume of solution. This results in a change of the mass per unit volume of the solution. Hence, the density increases.

### Kinetic Theory of Gases

The theory states that a gas consists of molecules that are in constant random motion.

### Gas pressure

**Gas pressure** is the force exerted by the particles of gas per unit area on the walls of the container. The rapid motion and collisions of molecules with the walls of the container causes pressure (force on a unit area – Nm<sup>-2</sup> = Pa). When you blow air into a balloon, the balloon expands because the pressure of air molecules is greater on the inside of the balloon than the outside. Pressure is a property which determines the direction in which mass flows. If the balloon is released, the air moves from a region of high pressure to a region of low pressure.

## Diffusion of Gases

Diffusion is the movement of molecules from a region of high concentration to a region of low concentration. This allows them to spread out and mix with other particles.

### Examples

- The smell of aftershave or perfume diffuses and is detected by people on the other side of the room.
- Opening a bottle of cologne or perfume.
- Spraying air freshener.
- Smell of food being cooked.
- Smell of urine.

## Relationship between volume, pressure and temperature

Have you ever wondered how an air powered water gun works? It uses the fantastic properties of gases to make a summer day more enjoyable!

### Boyle's Law

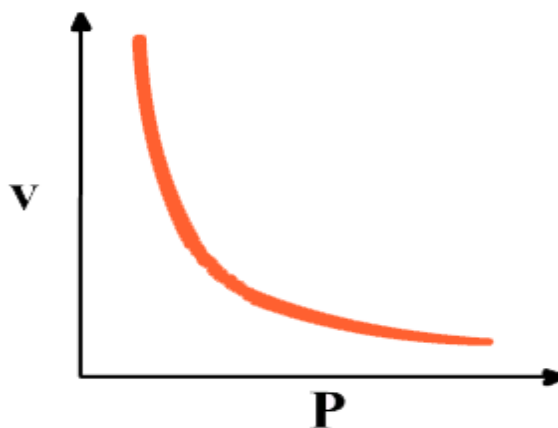
Boyle's Law is a basic law in chemistry describing the behaviour of a gas held at a constant temperature. The law, discovered by Robert Boyle in 1662, states that *at a fixed temperature, the volume of gas is inversely proportional to the pressure exerted by the gas*. In other words, when a gas is pumped into an enclosed space, it will shrink to fit into that space, but the pressure that the gas puts on the container will increase.

Boyle's Law can be written out mathematically as:

$$P \times V = \text{constant}$$

OR

$$P_1V_1 = P_2V_2; \text{ where } P = \text{pressure and } V = \text{volume}$$



### Real-World Examples

One example of Boyle's Law in action can be seen in a balloon. When air is blown into the balloon, the pressure of that air (a gas) pushes on the rubber, making the balloon expand. If one end of the balloon is squeezed, making the volume smaller, the pressures inside increases, making the un-squeezed part of the balloon expand out. There is a limit to how much the gas can be compressed. However, because eventually the pressure becomes so great that it causes the balloon (or any container) to break.

Another example is a syringe for taking blood. An empty syringe has a fixed amount of gas (air) in it. If the plunger is drawn back without the needle end being inserted into anything, the volume of the tube will increase and the pressure will drop, causing more air to move into the tube to equalise the pressure. If the syringe is inserted into a vein and the plunger drawn back, blood will flow into the tube since the pressure in the vein is higher than the pressure in the syringe.

## Charles's Law

Charles's law states that if a given quantity of gas is held at a constant pressure, its volume is directly proportional to the absolute temperature. Hence, as the temperature of the gas increases, the gas molecules will begin to move around more quickly and hit the walls of the container with more force, thus increasing the volume.

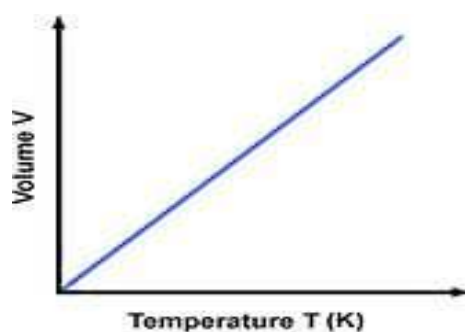


Jaqes Alaxandre Cesar Charles (1746-1823)

Charles's law can be written out mathematically as:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

**Graphically:**



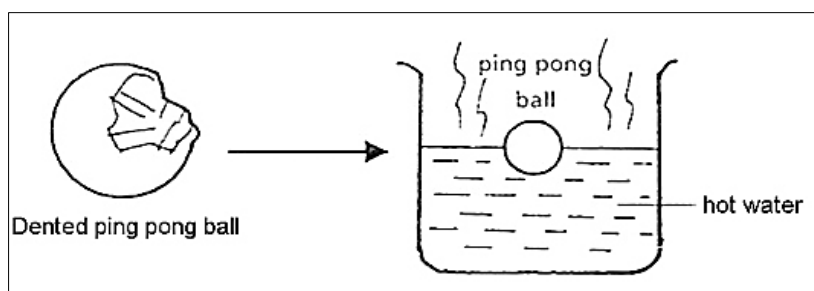
Note: Use only the *Kelvin* temperature scale when working with temperature in all gas law formulas.

### Helium Balloon on a Cold Day

On a cold morning, if a helium balloon is kept outside, the balloon will crumple. However, if the balloon is taken inside a warm home, the balloon returns to its original shape. This is because the gas takes up more space when it is warm.

### The Dented Ping Pong Ball

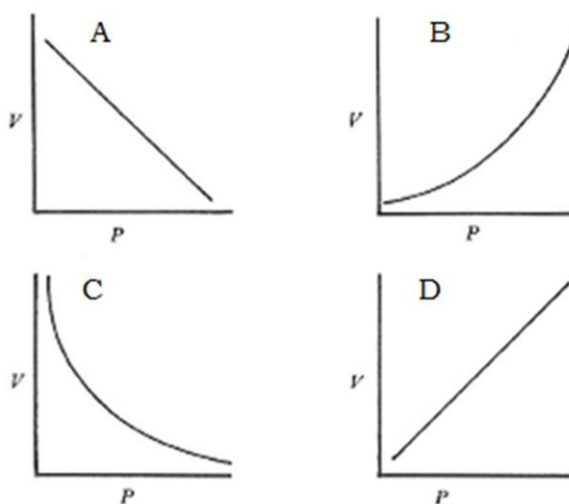
A dented ping pong (table tennis) can be restored by placing it into a saucepan half filled with water. Apply gentle heat to the saucepan, stirring constantly. If the ball is not cracked by the dent, the air inside will expand as it heats, pushing out the dent and restoring the ball to its original shape.



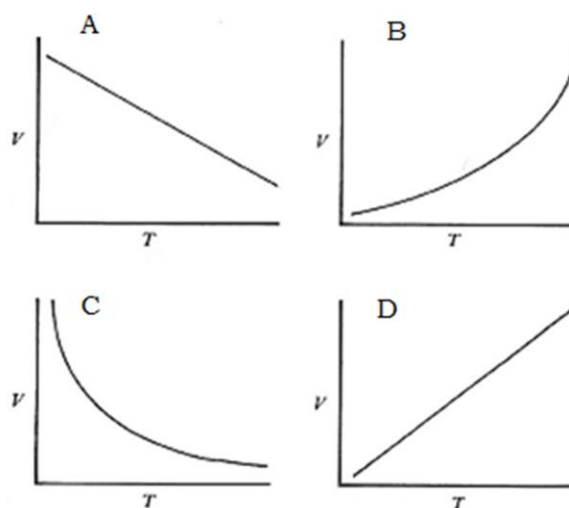
Source: <http://www.oldschool.com.sg>

### Exercise 2.1.2

- Enthalpy of fusion is the amount of heat energy needed to:
  - make the particles move.
  - change water from liquid to gas.
  - change water from solid at  $0^{\circ}\text{C}$  to liquid at  $0^{\circ}\text{C}$ .
  - change water from solid at  $0^{\circ}\text{C}$  to liquid at  $100^{\circ}\text{C}$ .
- Volume is related to pressure by Boyle's Law. Which of the graphs below represents Boyle's Law?



- Volume varies with temperature according to Charles' Law. Which of the graphs below represents Charles' Law?



- Gas pressure is caused by:
  - The weight (mass) of the molecules of gas.
  - The repulsion between gas molecules.
  - The collision of the molecules with the walls of the container.
  - The kinetic energy of the gas molecules.
- Mere noticed that as she spray-painted her bicycle, the spraycan grew steadily colder. How can this be explained in terms of the kinetic theory of gases?



## 2.2 Atomic Structure and Bonding

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Examine and present the electron configuration/arrangement of the first 20 elements.
- ✓ Explore and discuss the properties of elements in terms of their position in the periodic table.
- ✓ Investigate and explain the properties of any isotope.
- ✓ Evaluate and describe the stability of noble gases.
- ✓ Research and present on the findings of Dimitri Mendeleev in relation to the periodic table and John Dalton on atoms.
- ✓ Draw and explain the electron diagram illustrating the formation of ions and name them.
- ✓ Discover and write names and formula of ions.
- ✓ Describe and draw the formation of stable pairs and octets in bonding.
- ✓ Confirm by carrying out experiments on some general properties of ionic and covalent substances.

### Elements & Their Symbols

There are approximately 118 discovered and known elements. An element may be defined as 'a pure substance made up of only one type of atom'. For example, the element copper is made up of only copper atoms, and sulphur is made up of only sulphur atoms. Elements are distinguishable from each other by their atomic number, which is the number of protons in its nucleus.

The Periodic Table is a table of all the known elements. In the table, the elements are represented by their symbols. The symbol may consist of:

1. A single capital letter e.g. N for nitrogen and F for fluorine; OR
2. Two letters, one capital letter and one small lowercase letter e.g. Mg for magnesium and Pb for lead.

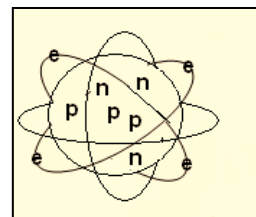
#### Exercise 2.2.1

Identify the first twenty elements of the Periodic table, including the common elements of bromine, iodine, iron, copper, iron, zinc, silver, barium and lead. Learn the symbols of these elements.

Hint: Create an acronym to help you remember the order and symbols of the first twenty elements.

Elements are made up of tiny particles called **atoms**. An atom is the smallest particle of matter. Atoms are made up of a central core called the nucleus and three subatomic particles:

1. **Protons** – They are positively charged particles and are found within the nucleus.
2. **Neutrons** - They have no charge, are neutral particles and found within the nucleus.
3. **Electrons** - They are negatively charged particles that are orbiting the nucleus.



All atoms are neutral because they contain the same number of protons (positive charge) as the number of electrons (negative charge).

On the Periodic Table, an element X appears as:  $\begin{matrix} A \\ Z \\ X \end{matrix}$

Where:

X: Represents the symbol of the element

Z: Represents the **Atomic Number** (it is always the smaller number).  
It is equal to the number of protons and is equal to the number of electrons.

A: Represents the **Atomic mass** (or Mass number) and it is the sum of number of protons and neutrons.

### Example

The atomic structure of an aluminium atom may be written as:

27
<b>Al</b>
13

Symbol:	Al
Atomic Number:	13
Mass Number:	27
Number of protons:	13
Number of electrons:	13
Number of neutrons:	14

### Isotopes

The atoms that make up an element all have the same number of protons. However, some elements have atoms whose number of neutrons differs. Therefore, these atoms have atomic masses that differ.

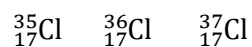
Atoms of the same element that have the same atomic number but different mass numbers are called **isotopes**.

An example is the isotopes of carbon as given in table given below:

Isotope	$^{12}_6\text{C}$	$^{13}_6\text{C}$	$^{14}_6\text{C}$
Mass Number	12	13	14
Atomic Number	6	6	6
Number of Neutrons	6	7	8

### Exercise 2.2.2

- The nucleus of an atom contains 9 protons and 10 neutrons.
  - What is its atomic number?
  - What is its atomic mass?
  - Name the element and give its symbol.
- What are isotopes?
- Isotopes of the same element should have the same number of
  - neutrons.
  - protons.
  - nucleus.
  - both neutrons and protons
- Isotopes of the same element should have different numbers of
  - neutrons.
  - protons.
  - electrons.
  - both protons and electrons.
- Consider the isotopes of chlorine given below and answer questions that follow:



- How many protons, electrons and neutrons are found in each isotope?
  - What structural characteristics do all the isotopes above have in common?
  - How does one isotope of chlorine differ from another isotope of chlorine?
- Complete the paragraph below.

The \_\_\_\_\_ number of an atom is the total number of \_\_\_\_\_ and protons present.

Atoms of an \_\_\_\_\_ with \_\_\_\_\_ numbers of neutrons are called \_\_\_\_\_.
  - Consider the three atoms,  ${}^1_1\text{H}$ ,  ${}^2_1\text{D}$  and  ${}^3_1\text{T}$ , where D stands for deuterium and T stands for tritium. What is the relationship between these atoms? Give as much information as you know about the relationship between these atoms. Which of the atoms do not have any neutron?
  - Briefly explain why the atoms of  ${}^{14}_6\text{C}$  and  ${}^{14}_7\text{N}$  have the same mass.

## The Periodic Table

The Periodic table is a table of all the known elements. Dmitri Mendeleev (1834-1907) was a Russian scientist who presented the elements in a table. He recognized that there was a pattern in the characteristics and behaviour of elements and that this pattern and behaviour was regularly repeated. Anything that is repeated regularly is said to be *periodic*. Mendeleev was able to place the elements into a table where the pattern the elements exhibited could be grouped. The modern Periodic Table shows this.

### Sections of the Periodic Table

The Periodic Table has eight groups, symbolized but roman numerals (Group I to Group VIII). The **groups** are the vertical columns on the table. The horizontal rows are called **periods**. For example, the second row elements are from Li to Ne.

Between Group II elements and Group III elements is a block of ten elements called the **transition** elements. Generally the metals are found on the left hand side of the periodic table, and the non-metals are on the right-hand side of the periodic table. The **metalloids** are found between the metals and non-metals. The Group VII elements are called the **halogens** or the salt-makers and the Group VIII elements are the **noble** or inert gases.

Periodic Table of Elements																				
1	2											10	11	12	13	14	15	16	17	18
IA	IIA											IIIA	IVA	VVA	VIA	VIIA	VIIIA	0		
1 H	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne		
11 Na	12 Mg	IIIB		IVB	VB	VIB	VIIA	VII		IB	IB	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar			
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr			
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe			
55 Cs	56 Ba	*La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn			
87 Fr	88 Ra	+Ac	Rf	Ha	106	107	108	109	110											

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

Legend - click to find out more...

<b>H - gas</b>	<b>Li - solid</b>	<b>Br - liquid</b>	<b>Tc - synthetic</b>
Non-Metals	Transition Metals	Rare Earth Metals	Halogens
Alkali Metals	Alkali Earth Metals	Other Metals	Inert Elements

Source: [paphysicalscience.wikispaces.com](http://paphysicalscience.wikispaces.com)

The elements in the Periodic Table are arranged in increasing order of their atomic number. In this way, some trends that have emerged in relation to the electron arrangement in the atoms of these elements are:

- All the elements in a group have the same number of electrons in their outermost (or valence) shell or the highest energy level. These valence electrons correspond to their group number. For example, lithium, sodium and potassium are Group I elements. In their outermost shells, they all have one electron. Beryllium, magnesium and calcium are Group II elements. They all have two electrons in their valence shells.
- All the elements in the same row or period have the same number of electron shells. For example, sodium with 11 electrons will have 3 electron shells, so will magnesium and aluminium. This is because all of them belong to Period 3.

## The modern version of the periodic table of the elements.

**Periodic table of the elements**

Legend:

- alkali metals
- alkaline earth metals
- transition metals
- other metals
- other nonmetals
- halogens
- noble gases
- rare earth elements (21, 39, 57-71)
- lanthanide elements (57-71 only)
- actinide elements

Periodic table of the elements showing groups (Ia to VIIIb, IIIa to VIIa) and periods (1 to 7). The lanthanide series (Ce to Lu) and actinide series (Th to Lr) are shown below the main table.

Source: <http://islam4jesus.org>

### Summary

Horizontal rows are called **Periods**.

Vertical columns are called **Groups**.

**Metals** are generally on the left side of the periodic table.

**Non-metals** are generally on the right side of the periodic table.

**Metalloids** which have properties of both metals and non-metals are found between the metal and non-metal blocks.

**Transition metals** are found in the centre of the table.

**Group VII elements** are the **halogens** or **salt-makers**

**Group VIII elements** are the **noble** or **inert** gases

### Trends in the Periodic Table

There are some trends seen in the physical properties of elements in their arrangement in the Periodic table.

- **Physical State** - Across the period, the state changes from solid to liquid to gas. All Group I elements are solid while all Group VIII elements are gases.
- **Melting point** - Is the temperature at which the elements change from solid to liquid. It is expressed in degrees Celsius and is shown in the table below.

H																	He
-259																	-272
Li	Be											B	C	N	O	F	Ne
181	128											2300	3550	-209	-218	-219	-248
Na	Mg											Al	Si	P	S	Cl	Ar
97.8	649											660.3	1410	44.09	112.8	-100	-189
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
63.3	839	1539	1660	1890	1857	1244	1535	1495	1453	1083	420	29.78	937	817	217	-7.2	-156

Source: <http://www.chemix-chemistry-software.com>

### Exercise 2.2.3

- What trend or pattern can you see in the melting points of the elements:
  - Across the period?
  - Down the group?
- What would be the cause of such trend?

- Boiling point** – is the temperature at which the element changes from liquid to gas.

H																	He
-252																	-268
Li	Be											B	C	N	O	F	Ne
1342	2970											2550	4827	-195	-182	-188	-246
Na	Mg											Al	Si	P	S	Cl	Ar
882.9	1090											2467	2355	280	444.6	-34.5	-185
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
759.9	1484	2832	3287	3380	2672	1962	2750	2870	2732	2567	907	2403	2830	613	684.9	58.8	-152

It is expressed in degrees Celsius as shown in the table below.

Source: <http://www.chemix-chemistry-software.com>

### Exercise 2.2.4

- What trend or pattern you can see in the boiling points of the elements:
  - Across the period?
  - Down the group?
- What would be the cause of such trend?

- Electrical conductivity** - it measures the elements ability to conduct an electric current. It is expressed in  $\times 10^6/\text{cm } \Omega$ .

Electrical Conductivity																	
H																	He
0																	0
Li	Be											B	C	N	O	F	Ne
0.108	0.313											0	0	0	0	0	0
Na	Mg											Al	Si	P	S	Cl	Ar
0.21	0.226											0.377	0	0	0	0	0
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
0.139	0.298	0.017	0.02	0.05	0.08	0.01	0.09	0.17	0.14	0.6	0.1	0.067	0	0.034	0	0	0

Source: <http://www.chemix-chemistry-software.com>

## Electron Arrangement

The electrons that are orbiting around the nucleus are in electron shells or energy levels. The **electron arrangement** (or configuration) describes the arrangement of electrons in shells or energy levels.

The table below shows the maximum number of electrons that can be accommodated by an electron shell.

Electron Shell (Energy level)	Maximum number of electrons that can be accommodated
1	2
2	8
3	8
4	18

### Example

Write the electron configuration for magnesium and potassium.

a. Magnesium has 12 electrons:

1<sup>st</sup> level: 2 electrons

2<sup>nd</sup> level: 8 electrons

3<sup>rd</sup> level: 2 electrons

Total 12 electrons

Electron configuration: **Mg (2, 8, 2)**

b. Potassium has 19 electrons:

1<sup>st</sup> level: 2 electrons

2<sup>nd</sup> level: 8 electrons

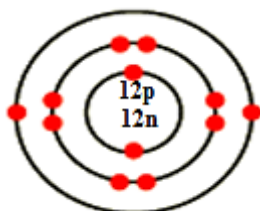
3<sup>rd</sup> level: 8 electrons

4<sup>th</sup> level: 1 electron

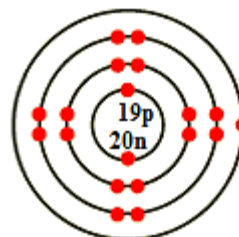
Total 19 electrons

Electron configuration: **K (2, 8, 8, 1)**

**Electron structure diagram**



**Electron structure diagram**



## Valence Electrons

The electrons in the outermost shell have the highest energy. They are called the **valence electrons**. In our examples of magnesium and calcium, magnesium has 2 valence electrons and potassium has 1 valence electron.

The valence electrons can be represented in an electron dot diagram.

### Example

Chlorine: Electron configuration: Cl (2, 8, 7)

Number of valence electrons: 7

Lewis structure diagram:





The table below shows some common metal and non-metal ions:

Name of Ion	Symbol	Name of Ion	Symbol
Bromide	Br <sup>-</sup>	Barium	Ba <sup>2+</sup>
Iodide	I <sup>-</sup>	Sodium	Na <sup>+</sup>
Iron	Fe <sup>2+</sup> , Fe <sup>3+</sup>	Chloride	Cl <sup>-</sup>
Copper	Cu <sup>+</sup> , Cu <sup>2+</sup>	Fluoride	F <sup>-</sup>
Zinc	Zn <sup>2+</sup>	Magnesium	Mg <sup>2+</sup>
Lead	Pb <sup>2+</sup>	Calcium	Ca <sup>2+</sup>
Oxygen	O <sup>2-</sup>	Potassium	K <sup>+</sup>
Lithium	Li <sup>+</sup>	Aluminium	Al <sup>3+</sup>

Monoatomic ions have only one type of atom. For example: Na<sup>+</sup>, Mg<sup>2+</sup> and Cl<sup>-</sup>. Polyatomic ions have two or more different types of atoms in the ion. For example: NH<sub>4</sub><sup>+</sup> (ammonium ion) and OH<sup>-</sup> (hydroxide ion).

## Bonding

When atoms gain electrons, these gained electrons must come from another atom. When atoms lose electrons, these lost electrons must go to another atom. In the case of losing or gaining, the electrons are *transferred* to another atom or shared between the two atoms.

### Ionic Bonding

Consider the substance sodium chloride, NaCl:

Na (11): 2,8, 1

Cl (17): 2,8,7

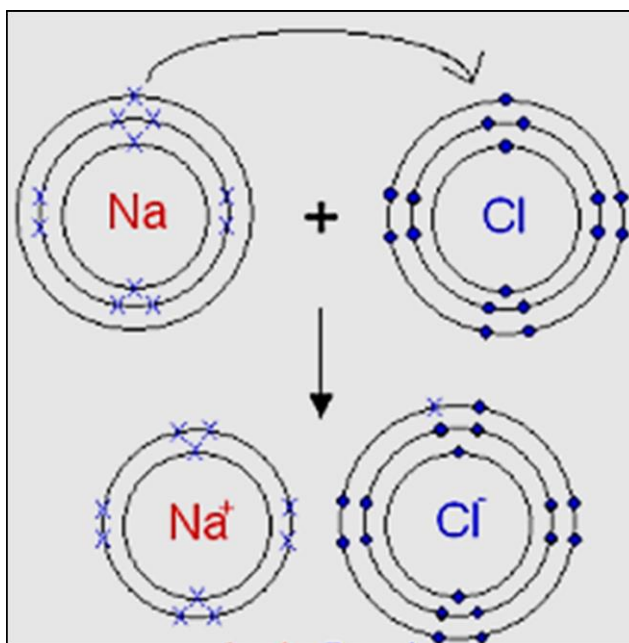
In the above example, in order for sodium (Na) to be stable, it must give away the 1 electron in its valence shell and in order for chlorine (Cl) to be stable; it needs 1 electron to have a fully filled valence shell.

Thus, the electron that sodium gives away will go to chlorine.

When sodium gives away its electron it will become Na<sup>+</sup> (sodium ion) while chlorine becomes a chloride ion (Cl<sup>-</sup>) upon receiving that electron.

In this case, the electron is **transferred** entirely from the sodium to the chlorine, and there is no sharing of electrons.

The force of attraction between positive and negatively charged particles or ions is called an **electrostatic force**. The bond that forms between positive and negative ions is called an **ionic bond**. There is a *transfer* of electrons during the formation of an ionic bond. Ionic bonds form between metals and non-metals.



Source: <http://www.gcscience.com>

### Summary

In ionic bonding, there is transfer of electrons and it occurs between metals and non-metals.

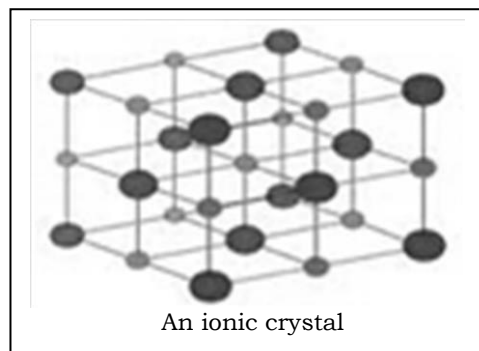
## Properties of Ionic Compounds

It is not difficult to identify ionic compounds because they are compounds that consist of a metal and a non-metal. Ionic compounds are compounds held together by ionic bonds.

Some properties of ionic compounds include:

1. They usually have a solid crystal structure. A crystal is an organised network of particles, having organised shapes where each ion that makes up the crystal is attached to another ion with the opposite charge.

The figure on the right shows a small portion of a sodium chloride crystal. Since ions are bound together by ionic bonds, the crystal is a very strong structure.



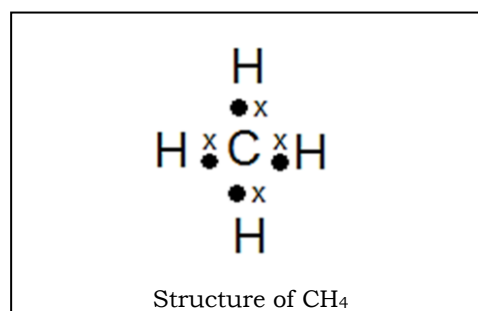
2. Due to the strong crystal structure of ionic compounds, they have very high melting and boiling points. This is because a lot of energy is needed to break the strong electrostatic force of attraction that holds the ions together.
3. The formula of the ionic compound is sometimes referred to as an **empirical formula**. It is a formula that gives the simplest ratio of positive ions to negative ions in the crystal. Empirical formulas are needed for ionic salts because there are millions of positive and negative ions and these cannot be represented in the formula. Thus, the formula NaCl means that for every sodium ion, there is a chloride ion.

## Covalent Bonding

Consider the compound methane, CH<sub>4</sub>.

C (2,4)

H (1)



In ionic compounds, ions are formed first then bonds form because oppositely charged particles are attracted to each other. Some elements like carbon and silicon do not form ions very easily as they have half-filled shells. Thus in order to fill their valence shells, these atoms must form **covalent bonds** where electrons in the valence shell are **shared**.

The structure of methane above shows the covalent bonds that form between the hydrogen and carbon atoms in the methane molecule. Carbon has four electrons in its valence shell and in order to become stable, it must *share* these electrons with four hydrogen atoms, each one having an electron and needing one more to become stable. This sharing of electrons results in the carbon and all the hydrogen atoms having fully filled shells.

Compounds like methane that have covalent bonds are called **molecules** or **molecular compounds**. Covalent bonds form between non-metals.

### Summary

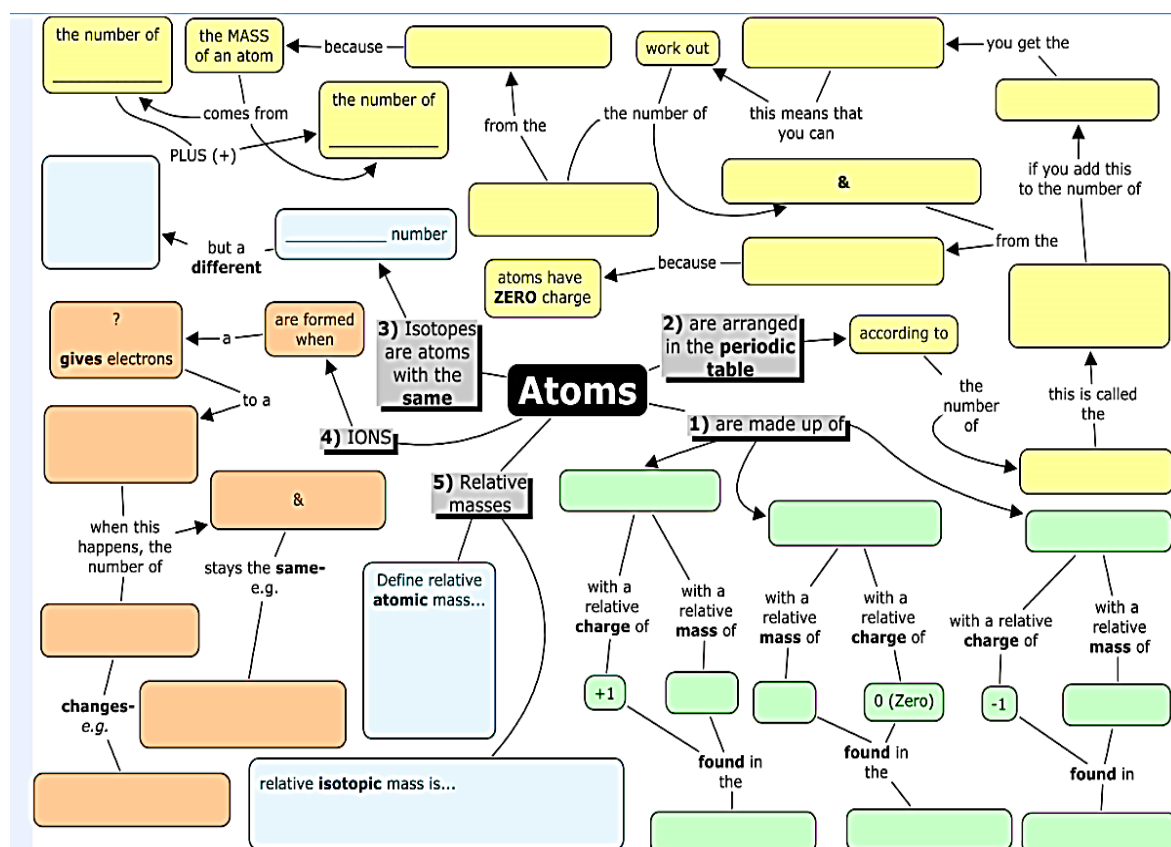
In covalent bonding there is sharing of electrons and it occurs between non-metals.

## Properties of Covalent Substance

1. Covalent bonding exists in solids, liquid and gases at room temperature.
2. Do not conduct electricity as they are made up of neutral particles.
3. The melting and boiling points are lower than those of ionic solids.
4. Lower solubility in water than ionic solids.

### Exercise 2.2.5

1. Silicon and carbon are in the same Group of the periodic table. This is best explained by the fact that both:
  - A. are non-metals.
  - B. are in Period 1 of the periodic table.
  - C. have the same number of valence electrons.
  - D. form ionic bonds.
2. An element has the electron configuration: 2,8,8,1. The element is
  - A. an alkali earth metal.
  - B. an alkali metal.
  - C. a transition metal.
  - D. a halogen.
3. The same number of electrons are present in each of:
  - A. an ion of Na, an ion of Mg and an ion of Ne.
  - B. an atom of Ar, an ion of Ca and an ion of Cl.
  - C. an atom of S, an ion of Ar and an ion of Ca.
  - D. an atom of S, an ion of O and an atom of H.
4. Complete the Concept map below.

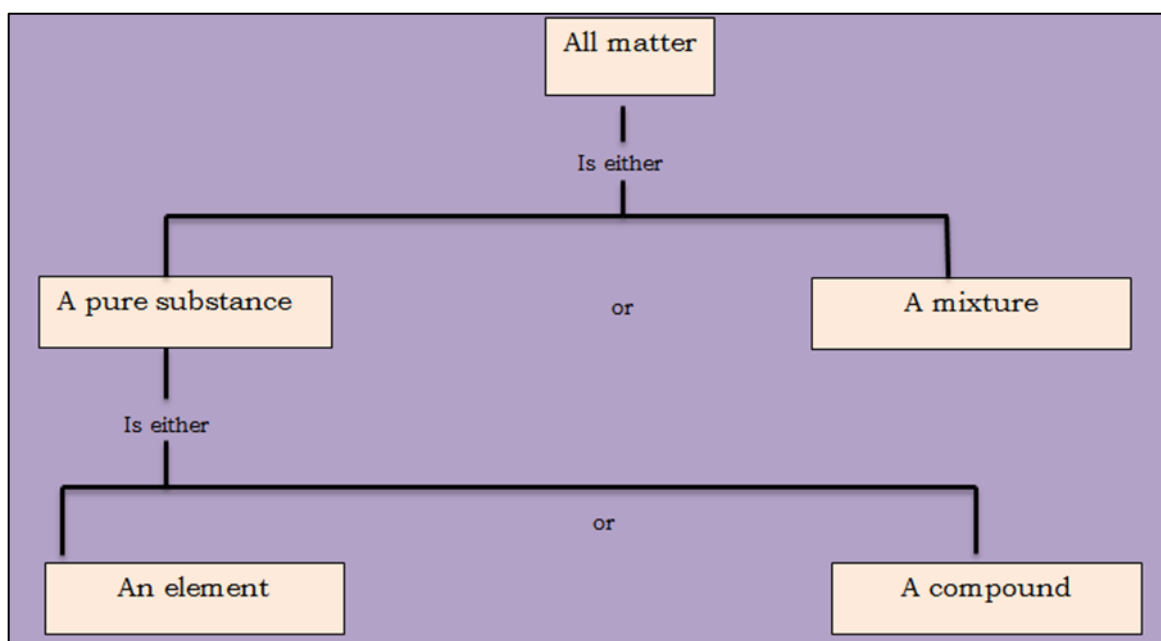


## 2.3 Elements, Compounds and Mixtures

### Achievement Indicators:

Upon completion of this sub-strand, students will be able to:

- ✓ Define and identify pure and impure chemicals in terms of formulae and samples into elements, compounds and mixtures.
- ✓ Show and carry out experiments illustrating physical change.
- ✓ Explore and discuss examples of chemical change.
- ✓ Identify and illustrate factors affecting solubility.
- ✓ Explain and draw the solubility curve of a salt at different temperatures and concentrations.
- ✓ Differentiate and carry out methods of separation and purification of mixtures.



### Elements

Elements are pure substances that are made up of only one type of atom. For example, carbon, hydrogen and oxygen. Carbon is made up of only carbon atoms, hydrogen is made up of only hydrogen atoms and likewise oxygen is made up of only oxygen atoms.

### Compounds

Compounds are pure substances that are formed when atoms of two or more different elements are chemically joined together in a definite ratio. For example, sodium chloride - NaCl (common salt) and sucrose -  $C_{12}H_{22}O_{11}$  (cane sugar) are examples of compounds. The properties of compounds are different from the properties of the substances it is made from.

### Mixtures

In mixtures different kinds of particles are simply mixed together. For example, Yaqona is a mixture of kava and water and likewise tea is a mixture of tea leaves, sugar and water. The properties of mixtures are similar to the properties of its parts.

## Pure chemicals

Chemical substances are often called 'pure' to set them apart from mixtures. In a pure chemical, all the particles are the same. A common example is pure water, it has the same properties and the same ratio of hydrogen to oxygen whether it is isolated from Rewa River or made in a laboratory. Limestone, gold and common salt are sometimes found in nearly pure forms.

Sodium chloride and magnesium is obtained from the sea. Sucrose is obtained from sugarcane and starch is found in foods such as dalo, roti and cassava. Similarly, nitrogen, oxygen and carbon dioxide are found in the air and bauxite, kerosene and other petroleum products are extracted from the ground.

## Physical and Chemical Change

A physical change is a change of form or state. For example, melting of ice is a physical change. A chemical change is any change that results in the formation of new chemical substance(s). For example burning of wood is an example of chemical change.

### Example of a Physical Change

#### *Hydrated and anhydrous CuSO<sub>4</sub>*

Copper sulphate (CuSO<sub>4</sub>) is a blue solid which is made up of crystals. When it is heated, the water is evaporated from the sample and is left with anhydrous copper sulphate (no water present) which is white in colour. Upon cooling after a few hours, the white copper sulphate absorbs water from the atmosphere to become hydrated copper sulphate again. Therefore, changing from hydrated to anhydrous CuSO<sub>4</sub> is a physical change since no new substance is formed.



### Examples of a Chemical Change


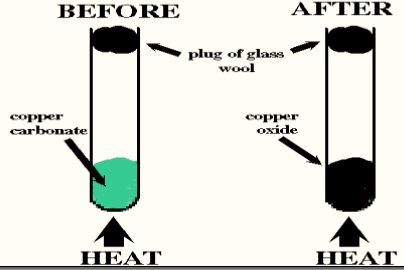
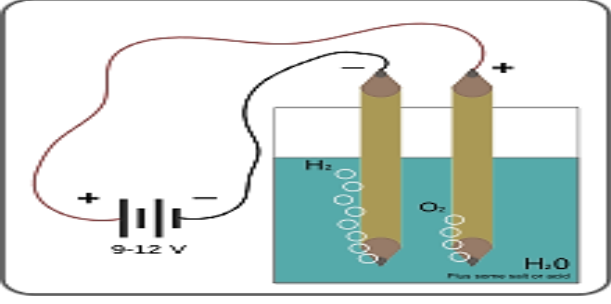
<p>When magnesium burns, it combines with oxygen to form magnesium oxide which is a new substance.</p>	
<p>When green copper carbonate (CuCO<sub>3</sub>) is heated, it decomposes to form two new substances namely copper oxide (CuO) and carbon dioxide (CO<sub>2</sub>).</p>	
<p>When acidified water is electrolysed, hydrogen gas and oxygen gas are produced.</p>	

Image source: <http://www.leekhigh.staffs.sch.u>

## Summary

Physical Change	Chemical Change
No new substance is formed	New substance is formed
Usually reversible	Usually non-reversible
Does not usually release heat	Often releases heat

### Exercise 2.3.1

1. Make a list of five elements, five compounds and five mixtures you have seen in your school laboratory or at home.
2. What is the difference between a compound and a mixture?
3. The difference between an element and a compound is that an element
  - A. is a diatomic molecule.
  - B. contains only one kind of atom.
  - C. can be separated by physical means.
  - D. is made from more than one kind of atom.
4. Which of the following is not a reversible change?
  - A. burning fuel
  - B. boiling water
  - C. dissolving sugar in water
  - D. mixing alcohol and water
5. Heating hydrated copper sulphate is a physical change.

With reference to the above statement, answer the following questions.

- i. What color are the copper sulphate crystals before they are heated?
  - ii. What color are the copper sulphate crystals after they are heated?
  - iii. What happens to the crystals when they are heated that causes them to change color?
  - iv. Why is this a reversible change?
6. Burning magnesium is a chemical change.

With reference to the above statement, answer the following questions.

- i. Briefly describe the appearance of magnesium wire.
  - ii. How would you describe the appearance of the substance that is formed when magnesium wire is burnt?
  - iii. What observation is made when magnesium wire is burnt?
  - iv. Is burning magnesium a reversible change or irreversible change. Justify your answer.
7. Define the following terms:
    - i. Element
    - ii. Hydrated
    - iii. Anhydrous
    - iv. Pure chemical

## Separating Substances

Most materials in our world are mixtures. Very few materials are pure substances. The art of separating mixtures is important because it enables us to isolate pure substances. For instance, salt is separated from sea water.

### Separating Insoluble Solids from Liquids

1. **Sedimentation** – Is a process in which insoluble particles can be removed from water or any other liquid. For example, when muddy water is stirred and left for a while, the dirt (impurities) in the water gradually settles down at the bottom of the container.
2. **Decantation** - Is pouring a solution from a container, leaving the sediments (precipitate) in the bottom of the container. For example, once the dirt has settled down, the water can be poured into another container.
3. **Filtration** – Is using a filter to remove solid particles from a liquid by causing the liquid to pass through the filter. For example, mixing yaqona is an example of filtration. The cloth traps the solid (kosa) and lets the liquid (yaqona) pass through.
4. **Centrifugation** – Uses centrifugal force to separate components of a substance. For example, separating dirt particles from clothes in a washing machine.

### Separating Solutes from Solvents

1. **Evaporation** – Separates a solute from a solvent by changing the solvent to a gas. For example, when salt water is heated, the water is evaporated (changes to gas) and the solute (salt) is left behind.
2. **Chromatography** – Is a process of separating out different parts of chemical mixtures onto an absorbent material that can then be individually analysed because different parts are caught on the material at different rates. For example, separation of plant pigments is the most common use of paper chromatography in biology.
3. **Distillation** – Is a way of separating the solvent from a solution by evaporation and condensation. For example, we can get pure water from sea water through the process of distillation.

### Separating Immiscible Liquids

Liquids are said to be immiscible if they do not mix together. A very good example is oil and water. Immiscible liquids can either be separated by decantation or by the use a separating funnel as shown in Figure below.



Source: SMC Lab Album

### Separating Miscible Liquids

Miscible liquids are liquids that can mix together. An example includes water and alcohol. Miscible liquids can be separated by fractional distillation. Fractional distillation separates liquids by their boiling points.

## Solutions

A **solute** is the solid part of a solution. For example, salt in salt solution.

A **solvent** is the liquid part of a solution. For example water in salt solution.

A **solution** is a homogeneous mixture of a solute and solvent. For example lemon juice is a mixture of lemon juice, sugar and water.

A **saturated solution** is a solution in which the maximum amount of solvent has been dissolved. Any more solute added will sit at the bottom of the container.

An **unsaturated** solution is a solution which can dissolve more solute. For example, a glass of water to which only one teaspoon of sugar is added is an unsaturated solution.

A **supersaturated solution** is a solution that is holding more dissolved solid than it can normally hold at that temperature. For example, supersaturated solutions of sugar and water are commonly used to make rock candy.

## Separating Solids from other Solids

If there are two solids and one is soluble and the other is not, they can be separated by dissolving the soluble solid. For example, in a mixture of salt and sand, sand can be obtained by dissolving the salt in water and then filtering the solution.

Sublimation can also be used to separate a solid that sublimates from the one that does not. For example, a mixture of ammonium chloride and sodium chloride can be separated by sublimation because ammonium chloride can sublime.

## Solubility

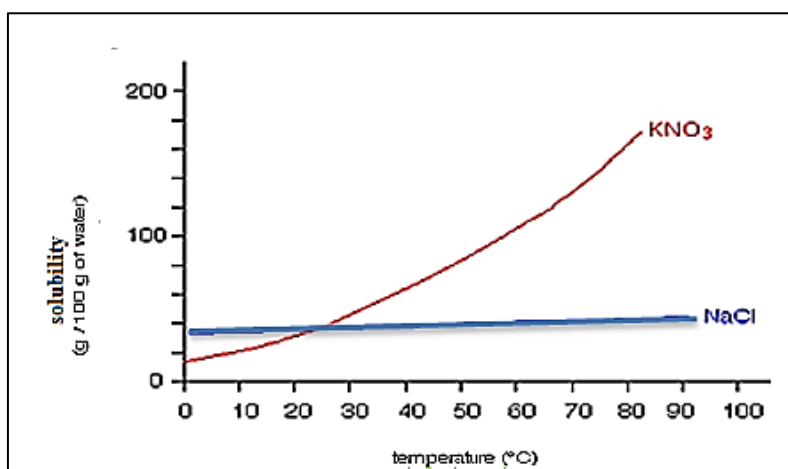
Solubility is the quantity of solute that dissolves in a given quantity of solvent to form a saturated solution. It is the maximum amount of solute that will dissolve. The solubility of a substance depends on the temperature.

### Effect of Temperature on Solubility

As the temperature increases, so does the solubility of salts. Most substances are more soluble in hot water than in cold water. A solubility curve is a graph of solubility versus temperature.

#### Example

#### The solubility curve of $\text{KNO}_3$ and $\text{NaCl}$



Source: <http://www.chemguide.co.uk>

The solubility curve above shows that the solubilities of potassium nitrate and sodium chloride increase with temperature.

### Exercise 2.3.2

1. Complete the statements below using the words given:

Compound, distillation, crystallisation, water

The two types of pure chemicals are elements and \_\_\_\_\_. Some solutions like seawater are separated by \_\_\_\_\_. The formation of a substance from a solution is called \_\_\_\_\_. A universal solvent is \_\_\_\_\_.

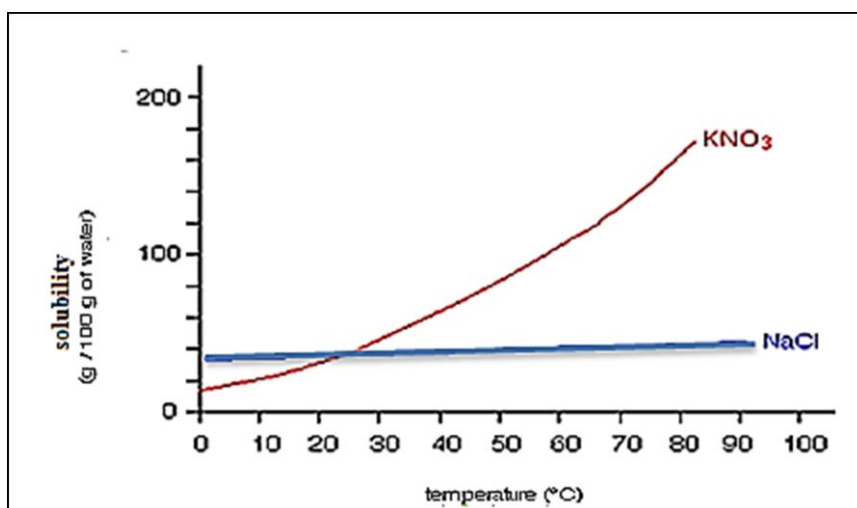
2. Briefly describe how you would separate:

- i. Alcohol from a mixture of alcohol and water.
- ii. Sand from a mixture of sand and salt.
- iii. Salt from a mixture of salt and water.
- iv. Different pigments in the leaves.
- v. Pure water from sea water
- vi. Oil from a mixture of oil and water.
- vii. Clean water from a mixture of mud and water.
- viii. Sugar from a mixture of sugar and naphthalene.
- ix. Tea from a mixture of tea leaves and tea.

3. When salt is dissolved in a beaker of water, a salt solution is made.

- i. What is a solution?
- ii. What is a solute? What is the solute in the salt solution?
- iii. What is a solvent? What is the solvent in the salt solution?

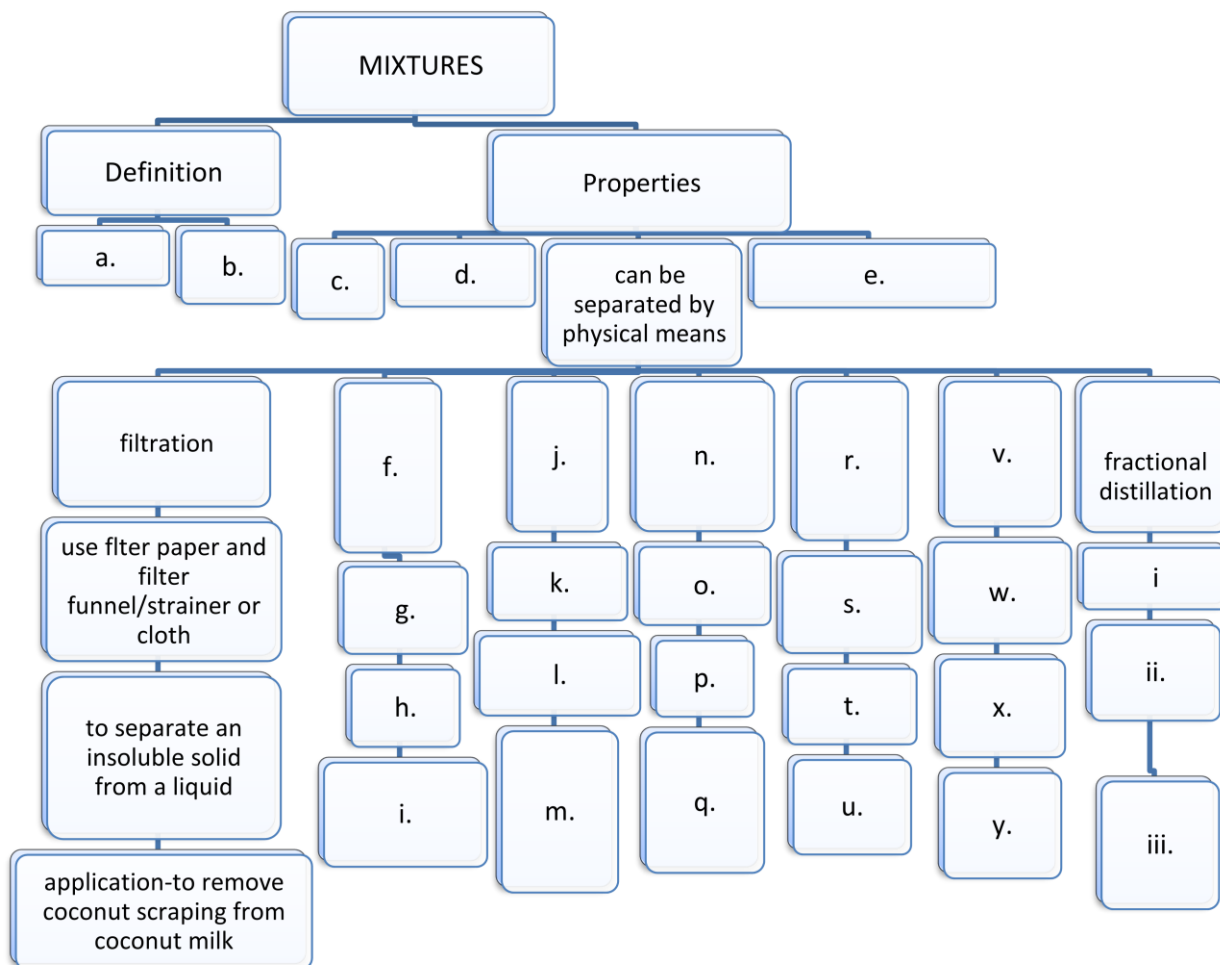
4. Consider the graph given below and answer the questions that follow.



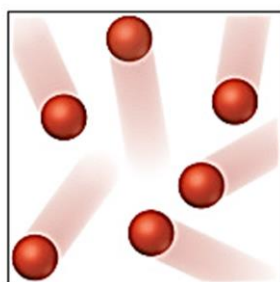
Source: <http://www.chemguide.co.uk>

- i. What name is generally given to such graphs?
- ii. Describe the relationship between the solubility of potassium nitrate (KNO<sub>3</sub>) and temperature.

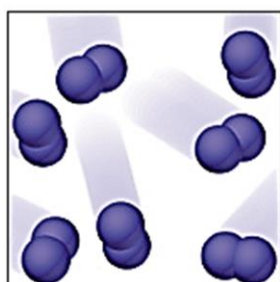
5. Complete the concept map given below.



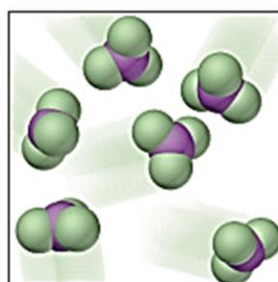
### Summary



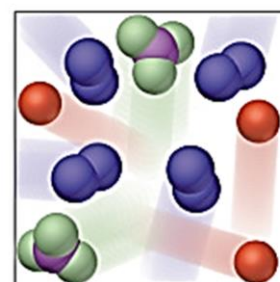
(a) Atoms of an element



(b) Molecules of an element



(c) Molecules of a compound



(d) Mixture of elements and a compound

Source: <http://wps.prenhall.com/>

**Strand Outcome:** Demonstrate an understanding of the chemical principles that involve changes during chemical reactions.

## Sub-strands

3.1 Chemical Equations and Calculations

3.2 Types of Reactions

3.3 Acids, Bases and Salts

## 3.1 Chemical Equations and Calculations

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Write the names and formula of common ionic and covalent compounds.
- ✓ Work out the percentage composition from formula mass and molecular mass.
- ✓ Write balanced equations from chemical statements and vice versa.
- ✓ Define and verify the Law of Conservation of Mass by carrying out experiments.
- ✓ Explore and carry out calculations to verify the Law of Definite Composition.

### Chemical formula and names of common substances

#### Chemical Formulas

Chemical formulas are used to express the composition of molecules and ionic substances in terms of chemical symbols.

#### Example

Chemical Formula	Name
NH <sub>3</sub>	Ammonia
H <sub>2</sub> O	Water
O <sub>2</sub>	Oxygen
NaCl	Sodium chloride
CaCO <sub>3</sub>	Calcium carbonate

Jons Jakob Berzelius invented a system of chemical notation in 1811. The system is based on the "law of definite proportions", which states that all samples of a given chemical compound have the same elemental composition. In other words, this law means that every chemical compound contains fixed and constant proportions (by weight) of its constituent elements.



Source: <http://crescentok.com>

## Writing Chemical Formula

The simplest chemical formulas are for binary compounds - compounds made up of two elements.

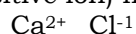
Note: The sum of the charges in a compound must always equal to zero.

### Example 1

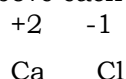
Write the chemical formula of calcium chloride.

#### Solution

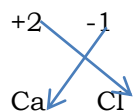
**Step 1:** Write the cation (positive ion) first followed by the anion (negative ion).



**Step 2:** Write the charges above each symbol.



**Step 3:** Cross the charges.



**Step 4:** Put the charges as subscript (do not write the positive and negative signs)



**Note:** Subscripts of 1 are NEVER written. They are understood.

**This method is commonly called the "Crossover method"**

### Other Examples

Write the chemical formula of the following:

1. Calcium hydroxide
2. Magnesium chloride
3. Aluminium oxide

#### Solution

1.  $\text{Ca}(\text{OH})_2$
2.  $\text{MgCl}_2$
3.  $\text{Al}_2\text{O}_3$

### Exercise 3.1.1

1. In the table given below, write the symbol of the positive and negative ions respectively (*The symbols of the first two ions are done for you*). Following that, write the correct formula of each compound that would form from the two ions in the appropriate box (*The first one is done for you*).

Name	Anions	chloride	oxide	hydroxide	nitrate	sulphate	carbonate
<b>Cations</b>	<b>Symbol of ions</b>	Cl <sup>-</sup>					
<b>sodium</b>	Na <sup>+</sup>	NaCl					
<b>potassium</b>							
<b>magnesium</b>							
<b>calcium</b>							
<b>aluminium</b>							
<b>ammonium</b>							
<b>lead (II)</b>							
<b>iron (II)</b>							
<b>iron (III)</b>							
<b>zinc</b>							
<b>copper (II)</b>							

2. Name the following substances:

- AgNO<sub>3</sub>
- PbSO<sub>4</sub>
- Al(OH)<sub>3</sub>

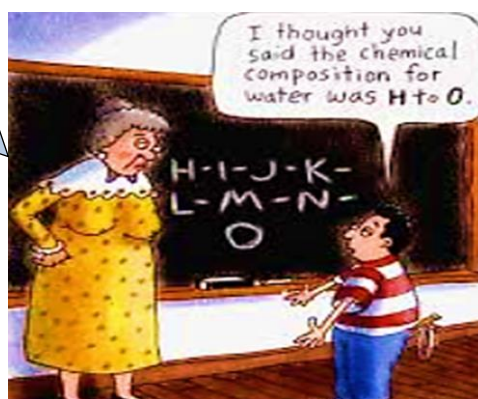
3. Write the formula of the following substances:

- Copper sulphate
- Zinc nitrate
- Calcium carbonate

4. The formula : (NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub> has

- One Nitrogen atom
- Four hydrogen atoms
- One sulphur atom
- Eight oxygen atoms

Importance of learning the correct chemical formula



## Molecular Mass (Mr)

- Used to express the mass of molecules. The mass is obtained by adding the relative atomic masses of the atoms in the molecule.
- There is no unit for molecular mass; however atomic mass unit (a.m.u) can be used.

### Example

Calculate the molecular mass of CO<sub>2</sub>.

### Solution:

$$\begin{aligned} \text{C: } & 1 \times 12 = 12 \\ \text{O: } & 2 \times 16 = 32 \\ \text{Molecular mass} & = 44 \text{ a.m.u} \end{aligned}$$

## Formula Mass

- Used to express the mass of **ionic** compounds. The mass calculation is similar to Mr.
- There is no unit for formula mass; however atomic mass unit (a.m.u) can be used.

### Example

Calculate the formula mass of Ca(NO<sub>3</sub>)<sub>2</sub>

$$\begin{aligned} \text{Ca: } & 1 \times 40 = 40 \\ \text{N: } & 2 \times 14 = 28 \\ \text{O: } & 6 \times 16 = 96 \end{aligned}$$

$$\text{Formula Mass} = 164 \text{ a.m.u}$$

### Exercise 3.1.3

Calculate the molecular/formula mass of the given substances.

- a. AgNO<sub>3</sub>      b. PbSO<sub>4</sub>      c. N<sub>2</sub>O<sub>3</sub>      d. NH<sub>3</sub>      e. PBr<sub>3</sub>      f. B<sub>2</sub>F<sub>6</sub>

## Percentage Composition

Percent composition gives the percentage of each element by mass present in a compound. This information is important because chemical compound is always the combination of two or more elements. Therefore, when studying a chemical compound, determining percent composition of a certain element within that chemical compound becomes necessary.

$$\text{Percent Composition} = \frac{\text{Mass of sample}}{\text{Total mass of compound}} \times 100$$

### Steps in determining the percent composition of the elements in a compound.

1. Determine the formula mass
2. Find the fraction mass of each element in the compound.
3. Convert to percentage.

### Example 1

Find the percent composition of red M&M's in a bag of 200 M&M's when there are 26 red M&M's altogether.

### Solution

$$\frac{26 \text{ red}}{200 \text{ total}} \times 100 = 13\%$$

### Example 2

Find the percent composition of copper and bromine in  $\text{CuBr}_2$ .

### Solution

i. % of Copper (Cu)

$$\begin{aligned} &= \frac{\text{Mass of Cu in the CuBr}_2}{\text{Total Mass of CuBr}_2} \times 100 \\ &= \frac{64}{224} \times 100 \\ &= 28.6\% \end{aligned}$$

ii. % of Bromine (Br)

$$\begin{aligned} &= \frac{\text{Mass of Br in the CuBr}_2}{\text{Total Mass of CuBr}_2} \times 100 \\ &= \frac{160}{224} \times 100 \\ &= 71.4\% \end{aligned}$$

**Hint:** The % composition of each element in a substance always adds up to 100%.

### Exercise 3.1.3

Calculate the % composition of each element in the following:

- NaOH
- $\text{KMnO}_4$
- $\text{Al}_2(\text{SO}_4)_3$

## Chemical Equations

- Chemical reactions convert reactants to products, whose properties differ from those of the reactants.
- They have the general form: Reactant  $\rightarrow$  Product

Equations are always written in the same format.

- The left side of the equation lists all the reactants.
- The right side of the equation lists all the products.
- The "+" means "reacts with".
- The arrow means "to produce".

## 1. Word Equations

Chemists use word equations to describe what has happened in a chemical reaction in a sentence/literal form.

### Example

Magnesium + Oxygen → Magnesium oxide

The above word equation can be read as follows: magnesium and oxygen are reactants and magnesium oxide is the product or magnesium and oxygen react to form magnesium oxide.

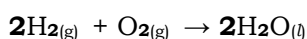
## 2. Chemical Equations

Chemical reactions are represented in a concise way by **chemical equations**.

### Example 1

For the word equation in (1), the chemical equation is:  $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$

### Example 2

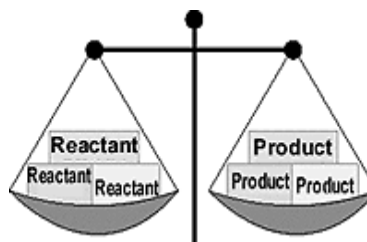
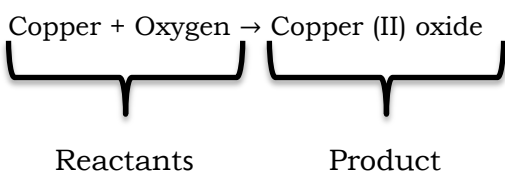


- Numbers in front of the formulas are **coefficients**. For example, the number 2 in front of  $\text{H}_2$  and  $\text{H}_2\text{O}$ . Coefficients indicate the relative number of molecules or ions of each kind involved in the reaction. Coefficients of 1 are never written - they are understood.
- Numbers to the lower right of chemical symbols in a formula are **subscripts**. For example the number 2 in  $\text{H}_2$  and  $\text{O}_2$ . Subscripts indicate the specific number of atoms of the element found in the substance. Subscripts of 1 are never written - they are understood.
- The **physical state** of each substance in a reaction may be shown in an equation by placing the following symbols to the right of the formula.
  - (g) for gas
  - (l) for liquid
  - (s) for solid
  - (aq) for aqueous (water) solution

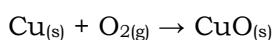
A chemical equation must have the same number of atoms of each element on both sides of the arrow. When this condition is met, the equation is said to be balanced.

## Balancing a Chemical Equation

Refer to the word equation:

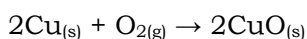


In the example above, if we replace the words shown above with the correct chemical formulae, we will get an unbalanced equation, as shown below:



To make the equation balanced, we need to adjust the number of units of some of the substances until we get equal numbers of each type of atom on both sides of the arrow.

Here is the balanced chemical equation:



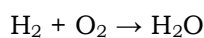
You can now see that there are two copper atoms and two oxygen atoms on each side.

### Example

Write the balanced chemical equation for the reaction between hydrogen and oxygen to produce water.

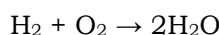
#### Step 1

Identify the reactants and products and write the formulae for each substance in an equation form.



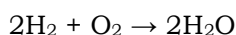
#### Step 2

Check for an unbalanced element. Adjust the number of each atom or molecule needed, but never change the formula. In this case it is O is unbalanced. We need two O atoms on each side.



#### Step 3

Check for another unbalanced element. In this example, there are 2 H atoms on the left and 4 H atoms ( $2 \times 2$ ) on the right. So we need to double the number of hydrogen molecules on the left.



**Hint: Never change subscripts to balance chemical equations**



Source: [www.clipartsheep.com](http://www.clipartsheep.com)

Here are some other examples of balanced equations. Check that you understand why they are balanced.

- $\text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2$
- $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$
- $4\text{Fe} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3$
- $4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}$
- $2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$

### Exercise 3.14

1. Write chemical equations for the following word equations:

- Ammonia gas plus oxygen gives nitrogen monoxide gas plus water vapour.
- Calcium hydroxide solution and carbon dioxide produces solid calcium carbonate and liquid water.
- Magnesium carbonate solution plus aqueous hydrochloric acid gives magnesium chloride plus liquid water and carbon dioxide gas.

2. Balance the following equations:

- $\text{N}_2 + \text{H}_2 \rightarrow \text{NH}_3$
- $\text{N}_2 + \text{O}_2 \rightarrow \text{N}_2\text{O}$
- $\text{C}_2\text{H}_5\text{OH} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- $\text{CaC}_2 + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_2 + \text{Ca}(\text{OH})_2$
- $\text{Fe}(\text{OH})_3 \rightarrow \text{Fe}_2\text{O}_3 + \text{H}_2\text{O}$

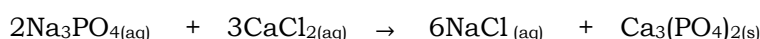
## Ionic Equation

The ionic equation is used to describe the chemical reaction while also clearly indicating which of the reactants and/or products exist primarily as ions in aqueous solution.

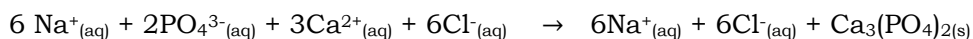
### Steps of writing a net ionic equation:

1. Start with a *balanced* molecular equation.
2. Break all soluble strong electrolytes (compounds with (aq) beside them) into their ions.
  - Indicate the correct formula and charge of each ion.
  - Indicate the correct number of each ion.
  - Write (aq) after each ion.
3. Bring down all compounds with (s), (l) or (g) unchanged.

#### Example

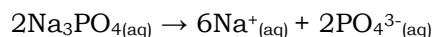


For the reaction given above, the complete ionic equation is:

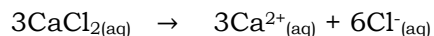


The following shows how each ion is achieved; Consider each reactant or product separately:

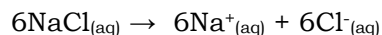
- 1 mole of  $\text{Na}_3\text{PO}_4$  contains 3 moles of  $\text{Na}^+$  and 1 mole of  $\text{PO}_4^{3-}$ . Since the balanced equation shows that two moles of  $\text{Na}_3\text{PO}_4$  are involved in the reaction, a total of 6 moles ( $2 \times 3$ ) of  $\text{Na}^+$  and 2 moles ( $2 \times 1$ ) of  $\text{PO}_4^{3-}$  are formed. Take note that the subscript "4" in the formula for the phosphate ion is not used when determining the number of phosphate ions present. The subscript is part of the formula for the phosphate ion itself.



- 1 mole of  $\text{CaCl}_2$  contains 1 mole of  $\text{Ca}^{2+}$  and 2 moles of  $\text{Cl}^-$ . Remember, the subscript "2" indicates the number of chloride ions. You will never have a diatomic chlorine ion (i.e.  $\text{Cl}_2^-$  or  $\text{Cl}_2^{2-}$ ) in aqueous solution. Since the balanced equation shows that 3 moles of  $\text{CaCl}_2$  are involved in the reaction a total of 3 moles ( $3 \times 1$ ) of  $\text{Ca}^{2+}$  and 6 moles ( $3 \times 2$ ) of  $\text{Cl}^-$  are formed.



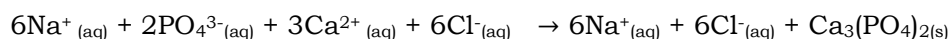
- 1 mole of  $\text{NaCl}$  contains 1 mole of  $\text{Na}^+$  and 1 mole of  $\text{Cl}^-$ . Since the balanced equation shows that 6 moles of  $\text{NaCl}$  are produced by the reaction, 6 moles ( $6 \times 1$ ) of  $\text{Na}^+$  and 6 moles ( $6 \times 1$ ) of  $\text{Cl}^-$  will be formed.



- Since calcium phosphate is an insoluble solid (indicated by the (s) beside its formula), it will not form ions in water. It is brought down unchanged into the complete ionic equation.

## Obtaining Net Ionic Equation

In the previous example, the complete ionic equation for the reaction between sodium phosphate and calcium chloride was:



If you look at the ions present on the reactant and product sides of the equation, you will notice that some of the ions are exactly the same.

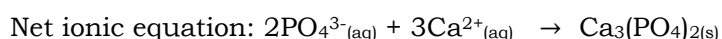
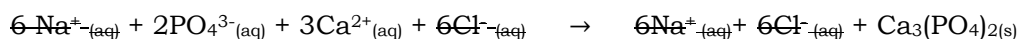
For example,  $6\text{Na}^+_{(\text{aq})}$  and  $6\text{Cl}^-_{(\text{aq})}$  are present on both sides of the equation. That means that the sodium ions and the chloride ions were present in the reaction mixture but *did not participate* in the reaction. The sodium and chloride ions in this reaction are referred to as **spectator ions**.

Spectator ions are ions that are present in the reaction mixture but do not participate in it. They "sit around and watch the reaction take place" just like a spectator at a rugby match watches the players in the game but doesn't play the game himself/herself. You can recognize spectator ions by looking for ions that are present on both sides of the equation. They will always have the same formula, charge and physical state. They will also be present in exactly the same number on both sides of the equation.

To get a net ionic equation:

1. Write the balanced molecular equation.
2. Write the balanced complete ionic equation.
3. Cross out the spectator ions that are present.
4. Write the "leftovers" as the **net ionic equation**.

For the above reaction, the net ionic equation is found by crossing out the spectator ions from the complete ionic equation:



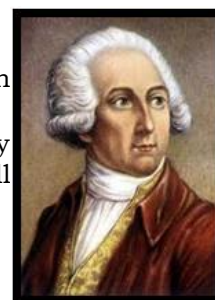
### Exercise 3.1.7

1. Write balanced net ionic equations for the following:

- i.  $\text{Mg}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g})$
- ii.  $\text{Zn}(\text{s}) + \text{CuSO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{Cu}(\text{s})$
- iii.  $2\text{NaOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
- iv.  $\text{Na}_2\text{CO}_3(\text{aq}) + 2\text{HNO}_3(\text{aq}) \rightarrow 2\text{NaNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

## Law of Conservation of Mass

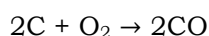
- The law of conservation of mass was established in 1789 by French Chemist Antoine Lavoisier
- This law states that mass is neither created nor destroyed in any ordinary chemical reaction. Therefore the mass of the reactants will always be equal to the mass of the products.



Source: www.catawiki.com

### Example 1

Carbon combines with oxygen at low concentration to form carbon monoxide.

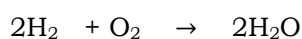


This equation shows that the total mass of the reactants and products are equal. The number of atoms of carbon and oxygen in the reactant and product sides are also equal.

Mass of Reactants	Mass of Products
<p>Mass of one Carbon atom = 12.010 g</p>	<p>Mass of one Carbon monoxide molecule = 12.010 + 15.999 = 28.009 g</p>
<p>Mass of two Carbon atoms = 2 x 12.010 = 24.020 g</p>	
<p>Mass of one Oxygen atom = 15.999 g</p>	<p>Total mass of the products = 2 x 28.009 g = <b>56.018 g</b></p>
<p>Mass of one Oxygen molecule = 2 x 15.999 g = 31.998 g</p>	
<p>Total mass of reactants = 24.020 + 31.998 = <b>56.018 g</b></p>	

### Example 2

Consider the formation of water molecule:

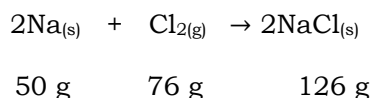


This equation also shows that the total mass of the reactants and products are equal. The number of atoms of hydrogen and oxygen in the reactant and product sides are also equal.

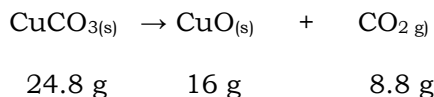
Mass of Reactants	Mass of Products
<p>Mass of one Hydrogen atom = 1.008g</p>	<p>Mass of one Water molecule = (2 x 1.008) + 15.999 = 18.015 g</p>
<p>Mass of one Hydrogen molecule = 2.016 g</p>	
<p>Mass of one Oxygen atom = 15.999 g</p>	<p>Total mass of the products = 2 x 18.015 g = <b>36.030 g</b></p>
<p>Mass of one Oxygen molecule = 2 x 15.999 g = 31.998 g</p>	
<p>Total mass of reactants = 4.023 + 31.998 = <b>36.030 g</b></p>	

### Examples showing the Law of Conservation of Mass

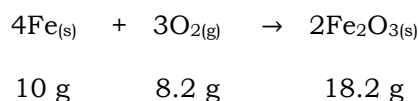
1. In an experiment where 50 g of sodium was reacted with 76 g of chlorine to form sodium chloride salt, it was found that 126 g of the salt was formed.



2. When 24.8 g of copper carbonate is strongly heated, it produces 16 g of copper oxide and 8.8 g of carbon dioxide gas.



3. When a 10 g sample of iron reacts with oxygen to form 18.2 g of ferric oxide, 8.2 g of oxygen was needed.

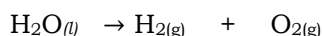


#### Exercise 3.1.8

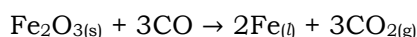
1. Consider the reaction of methane with oxygen:  $\text{CH}_{4(g)} + 2\text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + 2\text{H}_2\text{O}_{(l)}$

According to the law of conservation of mass:

- A. The mass of oxygen equals the mass of methane.
  - B. The mass of oxygen equals the mass of carbon dioxide.
  - C. The combined mass of methane and oxygen will be equal to the combined mass of carbon dioxide and water.
  - D. The mass of the reactants remains the same.
2. If 178.8 g of water is separated into hydrogen and oxygen gas and the hydrogen gas has a mass of 20.0 g, what is the mass of the oxygen gas produced?



3. From a laboratory process, a student collects 28.0 g of hydrogen and 224.0 g of oxygen. How much water was originally involved in the process?
4. Consider the reaction equation below and answer the question that follows.



If 160 g of  $\text{Fe}_2\text{O}_3$  reacted with CO to form 112 g of Fe and 132 g of  $\text{CO}_2$ , how many grams of CO was used in this reaction?

- 5.
- i. Write a balanced equation for the reaction of carbon with oxygen to form carbon dioxide.
  - ii. If 6 g of carbon reacts with oxygen to make 22 g of carbon dioxide, what mass of oxygen was used in the reaction?

## Law of Definite Composition

- The law of definite composition was proposed by Joseph Proust based on his observations on the fixed composition of chemical compounds.
- This law states that chemical compounds are composed of a fixed ratio of elements as determined by mass.
- Elements combine in whole numbers: it is not possible to have a compound with portion an atom.

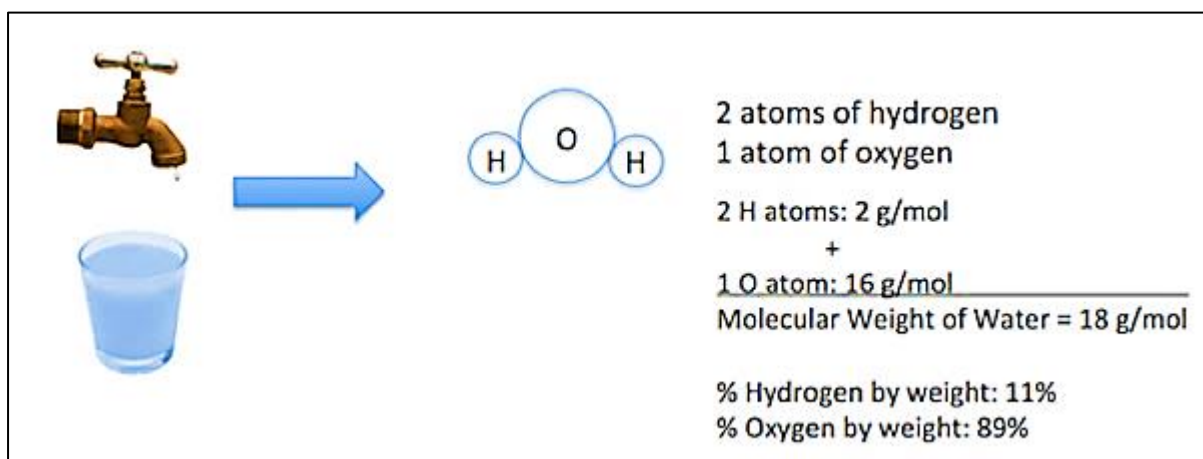


Source: [www.daviddarling.info](http://www.daviddarling.info)

### Example 1

The compound water ( $H_2O$ ) is always a chemical combination of hydrogen and oxygen in a 1:8 ratio by mass. This means that whatever the source of water, its composition is “two atoms of hydrogen and one atom of oxygen”. If a mixture of hydrogen and oxygen in some other ratio, say 1:2, were reacted, there would be water formed, but there would also be some un- reacted oxygen, because water always forms in the 1:8 ratio by mass.

Look at the illustration below:



Source:<http://study.com/>

Therefore by weight, there are 11% hydrogen and 89% oxygen in one molecule of water. This gives a ratio of 1:8 of hydrogen to oxygen in water.

### Example 2

Glucose has the chemical formula  $C_6H_{12}O_6$ . This means that for glucose to be formed, it must be comprised of 6 atoms of carbon, 12 atoms of hydrogen, and 6 atoms of oxygen. The carbon makes up 40 % of glucose; the hydrogen makes up 6.7% of glucose and the oxygen makes up 53.3% of glucose. This means that if the carbon, hydrogen and oxygen are combined in any other proportion than stated, glucose will not be formed.

### Example 3

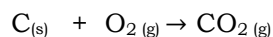
Ammonia, commonly used in household cleaning agents is made up of the elements hydrogen and nitrogen. It has the chemical formula;  $NH_3$  indicating that there is one atom of nitrogen combined with 3 atoms of hydrogen. This means that ammonia contains 82% nitrogen and 18% hydrogen by mass. Therefore, any other combination of hydrogen and nitrogen would result in an entirely different chemical compound.

## Calculations on Law of Definite Proportion

$$\% \text{ of element} = \frac{\text{Mass of the element}}{\text{Total mass of the compound}} \times 100$$

### Example 1

What is the experimental percentage of oxygen in  $\text{CO}_2$  if 42.0 g of carbon reacted completely with 112.0 g of oxygen?



$$\begin{aligned} \% \text{ O} &= \frac{\text{Mass of oxygen}}{\text{Mass of carbon dioxide}} \times 100 \\ &= \frac{112 \text{ g}}{112 \text{ g} + 42 \text{ g}} \times 100 \\ &= 72.7\% \end{aligned}$$

### Example 2

In a 80 gram sample of sodium hydroxide ( $\text{NaOH}$ ), calculate the mass of each element.

#### Solution

First find the % composition of each element in  $\text{NaOH}$  ( $M_r = 40$ )

$$\text{Na: } \frac{23}{40} \times 100 = 57.5\% \quad \text{O: } \frac{16}{40} \times 100 = 40\% \quad \text{H: } \frac{1}{40} \times 100 = 2.5\%$$

Find the mass from the % composition.

$$\begin{aligned} \text{i. Na} &= 57.5\% \\ &= \frac{57.5}{100} \times 80 \text{ g} \\ &= \underline{46 \text{ g}} \end{aligned}$$

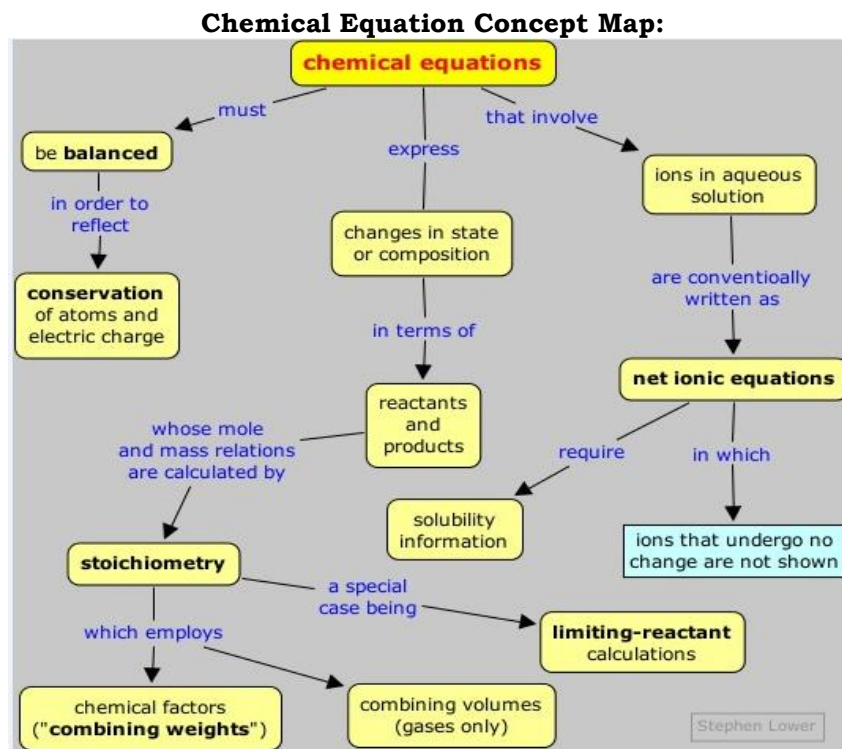
$$\begin{aligned} \text{ii. O} &= 40\% \\ &= \frac{40}{100} \times 80 \text{ g} \\ &= \underline{32 \text{ g}} \end{aligned}$$

$$\begin{aligned} \text{iii. H} &= 2.5\% \\ &= \frac{2.5}{100} \times 80 \text{ g} \\ &= \underline{2 \text{ g}} \end{aligned}$$

Check: The total mass adds up to 80 g. This indicates that the calculations are correctly done.

### Exercise 3.1.8

1. Find the mass of each element in 50 g sample of calcium hydroxide ( $\text{Ca}(\text{OH})_2$ ).
2. Calculate the mass of copper and bromine in 112 g of  $\text{CuBr}_2$ .



Who will win the race below?

### Balancing Equations Race

- 1)  $\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- 2)  $\text{Al} + \text{Fe}_3\text{N}_2 \rightarrow \text{AlN} + \text{Fe}$
- 3)  $\text{Na} + \text{Cl}_2 \rightarrow \text{NaCl}$
- 4)  $\text{H}_2\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{O}_2$
- 5)  $\text{C}_6\text{H}_{12}\text{O}_6 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{CO}_2$
- 6)  $\text{H}_2\text{O} + \text{CO}_2 \rightarrow \text{C}_7\text{H}_8 + \text{O}_2$
- 7)  $\text{NaClO}_3 \rightarrow \text{NaCl} + \text{O}_2$
- 8)  $(\text{NH}_4)_3\text{PO}_4 + \text{Pb}(\text{NO}_3)_4 \rightarrow \text{Pb}_3(\text{PO}_4)_4 + \text{NH}_4\text{NO}_3$
- 9)  $\text{BF}_3 + \text{Li}_2\text{SO}_3 \rightarrow \text{B}_2(\text{SO}_3)_3 + \text{LiF}$
- 10)  $\text{C}_7\text{H}_{17} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- 11)  $\text{CaCO}_3 + \text{H}_3\text{PO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + \text{H}_2\text{CO}_3$
- 12)  $\text{Ag}_2\text{S} \rightarrow \text{Ag} + \text{S}_8$
- 13)  $\text{KBr} + \text{Fe}(\text{OH})_3 \rightarrow \text{KOH} + \text{FeBr}_3$
- 14)  $\text{KNO}_3 + \text{H}_2\text{CO}_3 \rightarrow \text{K}_2\text{CO}_3 + \text{HNO}_3$
- 15)  $\text{Pb}(\text{OH})_4 + \text{Cu}_2\text{O} \rightarrow \text{PbO}_2 + \text{CuOH}$
- 16)  $\text{Cr}(\text{NO}_2)_2 + (\text{NH}_4)_2\text{SO}_4 \rightarrow \text{CrSO}_4 + \text{NH}_4\text{NO}_2$
- 17)  $\text{KOH} + \text{Co}_3(\text{PO}_4)_2 \rightarrow \text{K}_3\text{PO}_4 + \text{Co}(\text{OH})_2$
- 18)  $\text{Sn}(\text{NO}_2)_4 + \text{Pt}_3\text{N}_4 \rightarrow \text{Sn}_3\text{N}_4 + \text{Pt}(\text{NO}_2)_4$
- 19)  $\text{B}_2\text{Br}_6 + \text{HNO}_3 \rightarrow \text{B}(\text{NO}_3)_3 + \text{HBr}$
- 20)  $\text{ZnS} + \text{AlP} \rightarrow \text{Zn}_3\text{P}_2 + \text{Al}_2\text{S}_3$

## 3.2 Types of Reactions

### Achievements Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Distinguish and describe different types of reactions based on chemical statements and balanced chemical equations.
- ✓ Describe the reactions of oxidation and reduction in terms of transfer of atoms and electrons.
- ✓ Study and write simple oxidation and reduction reactions involving atoms and electrons.
- ✓ Show that electrolysis of molten and aqueous salt experimental set-up involves oxidation and reduction.

### Chemical Reactions

Chemical reactions are processes that will cause change in the properties of the substances involved. Most reactions are chemical changes and are *irreversible* and some are *reversible*. Chemical reactions in living things are termed biochemical including photosynthesis, respiration and digestion of food in the gut of animals.

Chemical reactions are part of our daily lives in our homes and communities.

#### Example



A combustion reaction

The study of chemical reactions is an integral aspect of chemistry. It equips us with the ability to understand and explain the chemical principles that involves changes.

These chemical reactions are:

1. Combustion
2. Synthesis
3. Decomposition
4. Neutralisation
5. Double Displacement
6. Precipitation
7. Oxidation-Reduction

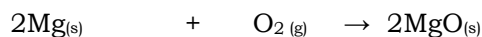
## 1. Combustion

It is the chemical term for the burning of substances in oxygen to form compounds called *oxides*. Though oxygen does not burn, it is used as it supports combustion.

Metals will burn completely in oxygen to form metallic oxides. The oxides are *ionic* compounds and are *basic* in nature.

### Example

Magnesium + Oxygen → Magnesium oxide



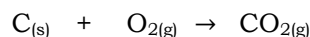
Combustion of metals may be used to distinguish some common metals as metals burn with distinctive flame.

Non-metals burn completely in oxygen to form non-metal oxides. These oxides are *molecular* substances and are *acidic* in nature. Most are gases at room temperature.

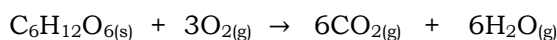
Organic compounds are used as fuels for its high carbon content. It burns completely in oxygen to produce carbon dioxide and water. A lot of energy is released. Incomplete combustion will form harmful products such as carbon monoxide, soot (unburnt carbon) and less heat is released.

### Examples

1. Carbon + Oxygen → Carbon dioxide



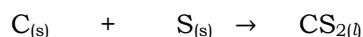
2. Glucose + Oxygen → Carbon dioxide + Water



## 2. Synthesis

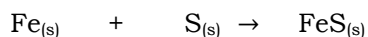
Naturally occurring elements combine chemically to form compounds. When two non-metals combine, a covalent substance is formed. However, metals combine with a non-metal to form ionic compounds.

### Example 1 Combination of two non-metals.



Carbon disulphide

### Example 2 Combination of a metal and a non-metal



Iron sulphide

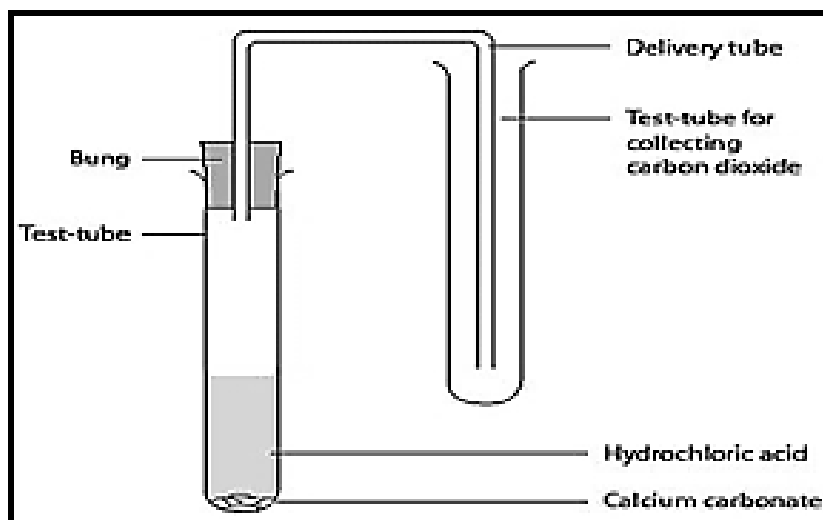
### Example 3 Formation of oxides

All combustion of elements is synthesis reaction.

## 3. Decomposition

Some carbonates and nitrates are decomposed by heat. Carbonates are decomposed to form carbon dioxide and the oxide of the metal.

The set up below shows the laboratory preparation of carbon dioxide by the decomposition of marble chips,  $\text{CaCO}_3$ .



The presence of the carbon dioxide formed can be tested by passing it through lime water.

### Example

Calcium carbonate  $\rightarrow$  Calcium oxide + Carbon dioxide



#### *Did you know?*

White wash for painting decorative garden stones is calcium oxide produced in coastal Fijian villages by heating sun dried coral in earth oven, 'lovo'.

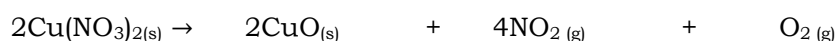
Calcium oxide was also used to form a paste to straighten/ stiffen wavy hair.

Green copper carbonate crystals will decompose to form a black solid, copper oxide ( $\text{CuO}$ ). White lead carbonate is decomposed to form the yellow solid, lead oxide ( $\text{PbO}$ ).

The nitrates of copper, lead and zinc will change to liquid by gentle heating. The salt had dissolved in its water of crystallisation. When the liquid is strongly heated, a brown gas is given off and a solid deposit is also formed. The other gaseous product is oxygen.

### Example

Copper nitrate  $\rightarrow$  Copper oxide + Nitrogen dioxide + Oxygen



Blue copper nitrate crystals decomposes to form a black powdered (copper oxide), zinc nitrate crystals will decompose to form white powder (zinc oxide) and lead nitrate will decompose to form yellow lead oxide.

### Exercise 3.2.1

- For each reaction below:
  - Write a balanced equation.
  - Classify the type of reaction.
    - Burning of sulphur
    - Burning of magnesium
    - Formation of ammonia from nitrogen gas and hydrogen gas
- Balance the equations given below:
  - $\text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3$
  - $\text{SO}_2 + \text{O}_2 \rightarrow \text{SO}_3$
  - $\text{KMnO}_4 \rightarrow \text{K}_2\text{MnO}_4 + \text{MnO}_2 + \text{O}_2$

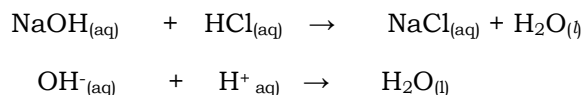
## 4. Neutralisation

In a neutralisation reaction, *acids react with bases to form salt and water*. The hydrogen ion of the acid will react with the hydroxide ion of the base to form water which is neutral. Substances which are acidic in nature include mineral acids, citrus juice, vinegar, gastric juice and non-metal oxides. Common bases are oxides and hydroxides of metals, ammonia and carbonates.

The reaction is also known as the **acid-base reaction**.

### Example 1

Sodium hydroxide + Dilute hydrochloric acid  $\rightarrow$  Sodium chloride + Water



**Note:** Carbonates react with dilute acids to form a salt, water and carbon dioxide.



*Did You Know?*

Rust on iron surfaces is removed by soaking the metal in lemon juice or vinegar.

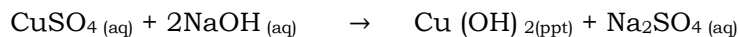
## 5. Precipitation

It is the formation of an *insoluble* salt from the mixture of two different *clear* solutions. The *insoluble salt* formed is the *precipitate* (ppt). In earlier science classes we learnt about the formation of scum by mixing soap solution and hard water. Scum is a precipitate. Some precipitate may settle at the bottom of the test tube; others will form a suspension.

The illustration below shows the formation of the precipitate copper hydroxide from mixing copper sulphate solution with caustic soda, NaOH<sub>(aq)</sub>



### Reaction Equation



### Writing Ionic Equations

An ionic equation is used to emphasise the species involved in a reaction.

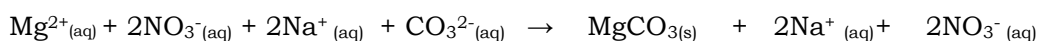
The following steps are used to determine the ionic equation from a balanced chemical equation.

1. Write the balanced equation for the reaction.

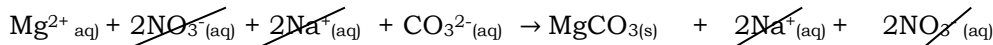
#### Example



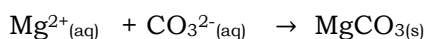
2. Write all the ions present in aqueous solutions in the reaction separately. Ions contained in solids, liquids and gases will not be separated.



3. Cancel all the ions that appear on both sides of the equation, *spectator ions*. These ions are not involved in the reaction being emphasised.



4. Write the balanced ionic equation from the remaining species in Step 3.



*The balanced ionic equation shows that the ions involved in forming the precipitate are magnesium, Mg<sup>2+</sup> and carbonate ions, CO<sub>3</sub><sup>2-</sup>.*

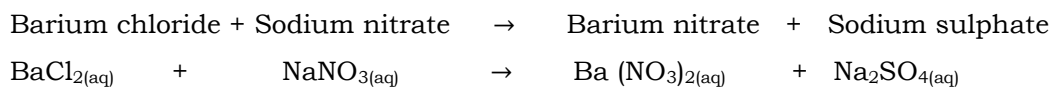
### Exercise 3.2.2

1. For each reaction below:
  - i. Write a balanced chemical equation.
  - ii. Classify the type of reaction and give a reason for your choice.
    - a. Precipitation of silver chloride by reacting barium chloride with silver nitrate.
    - b. Formation of solid barium sulphate by reacting barium chloride with dilute sulphuric acid.
    - c. Release of carbon dioxide by reacting sodium carbonate with dilute sulphuric acid.
    - d. Copper metal formed as zinc granules is placed into a test tube containing copper sulphate solution.
2. Balance the equations given below:
  1.  $\text{MgCO}_3 + \text{HNO}_3 \rightarrow \text{Mg(NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O}$
  2.  $\text{MgO} + \text{CH}_3\text{COOH} \rightarrow \text{Mg(CH}_3\text{COO)}_2 + \text{H}_2\text{O}$
  3.  $\text{NH}_4\text{OH} + \text{Al}_2(\text{SO}_4)_3 \rightarrow \text{Al(OH)}_3 + (\text{NH}_4)_2\text{SO}_4$
  4.  $\text{K}_2\text{O} + \text{H}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O}$

## 6. Double Displacement

When two different salt solutions react forming a clear solution. The resultant salts formed are both soluble in water. It is termed double displacement as the anions are exchanged between the two cations.

### Example

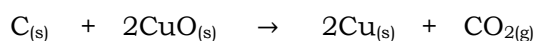


## 7. Oxidation-Reduction

Oxidation is the gain of oxygen. For example, combustion and corrosion reactions are oxidation reactions. Reduction is the loss of oxygen. Oxidation and reduction reactions occur simultaneously. As a substance is reduced, the other reactant will be oxidised.

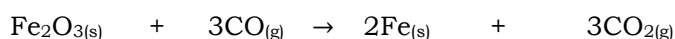
### Example 1

In the extraction of metal from metal oxides using carbon, the metal oxide is reduced to the metal and carbon is oxidised to carbon dioxide.



### Example 2

Iron metal is produced in the Blast Furnace by the reduction of iron (III) oxide by carbon monoxide. The carbon monoxide is oxidised to carbon dioxide.

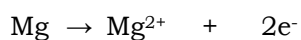
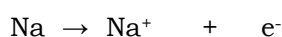


The collective term for oxidation-reduction reaction is **redox**. Some of these changes may not involve the movement of oxygen, so the two reactions are also defined by the transfer of hydrogen and electrons.

Oxidation is the loss of hydrogen or electrons. Reduction is the gain of hydrogen or electrons.

### Example 1: Oxidation reaction

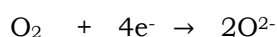
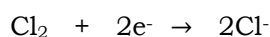
Ionisation of metal atoms to positive metal ions (cations) is an **oxidation** reaction as electrons are *lost* from its valence shell.



**Note:** Electron(s) lost will appear on the right side of the equation.

### Example 2: Reduction reaction

Non-metals *gaining* electron(s) to form stable negative ions (anions) are **reduction** reactions.

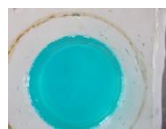


**Note:** Electron(s) gained will appear on the left side of the equation.

### Example 3

Displacement reactions of metals are redox reactions. The more active metal is oxidised to its ions and the less active metals is reduced from its ionic form (in aqueous solution) to the metal.

The figure below shows iron filings placed in a beaker of copper sulphate solution reduces the copper ions (light blue solution) to copper metal (reddish brown)



**Copper sulphate solution**



**Iron filings**



**Mixture of Iron and copper sulphate solution**



**Reddish brown deposits formed**

Oxidation Reaction:  $\text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^-$

Reduction Reaction:  $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

#### **Example 4**

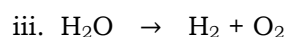
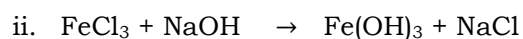
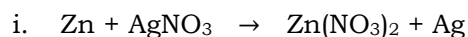
Displacement reactions of active metals with dilute acid are redox reactions. The more active metal is oxidised to its ions and the hydrogen ions in acid is reduced to hydrogen gas (bubbles evolved). (*Refer to Chapter 2*)

Oxidation Reaction:  $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$

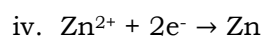
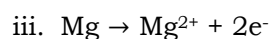
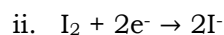
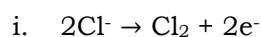
Reduction Reaction:  $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$

#### **Exercise 3.2.3**

1. Balance the equations given below and identify the type of reaction shown by each equation:



2. Identify the following equations as either oxidation or reduction.



3. The reaction of lead oxide with carbon forms lead metal and carbon dioxide.

i. Write a balanced chemical equation to represent the reaction above.

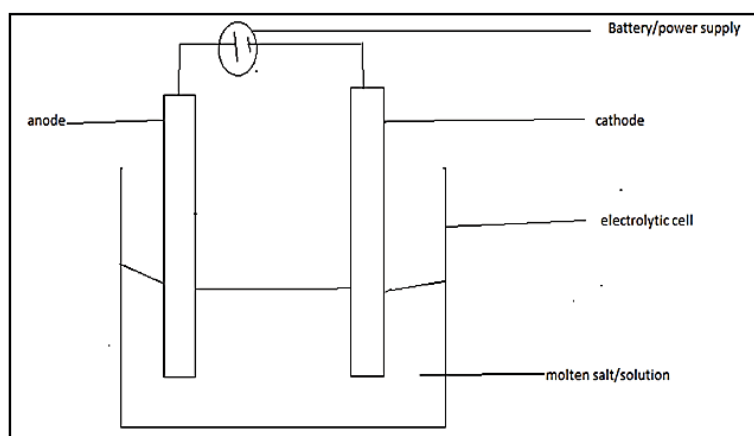
ii. From the equation, determine which reactant is oxidised and which is reduced.

iii. Explain why the reaction between lead oxide and carbon is called a redox reaction.

## Electrolysis

Redox is commercially used in a process called electrolysis. An English chemist, Humphry Davy in the 1800s found out that electricity could be used to make reactions. Electrolysis is the decomposition of an electrolyte by passing an electric current through it. An electrolyte is a molten salt or solution that conducts electricity.

Electrolysis is carried out in an electrolytic cell, as in Figure below.



The components of an electrolytic cell are:

1. **Electrolyte** – molten or solutions of ionic compounds. The mobile/free ions are the carriers of electric current. Examples include:  $\text{NaCl}_{(l)}$ ,  $\text{NaCl}_{(aq)}$ ,  $\text{H}_2\text{O}_{(l)}$ ,  $\text{MgCl}_{2(aq)}$ ,  $\text{CuSO}_{4(aq)}$ .
2. **Batteries/Direct Current, DC power supply** – source of current, creates or discharge ions in the electrolyte. The electrode potential should be large enough to drive the reactions.
3. **Electrodes** – connects batteries/DC power supply to electrolyte. The two types are **anode** (positively charged) and **cathode** (negatively charged). Electrodes are usually inert or unreactive and a conductor of electricity. A common electrode is carbon (graphite) as it is inert and a conductor. Less reactive metals such as copper, iron and zinc, are used in electroplating.

The electrolytic cell is a complete circuit. Electrons move from the anode to the cathode. As electrons moves away from the anode, the anode becomes positive and as electrons are deposited on the cathode, the cathode becomes negative.

### Electrolysis of a salt solution

The degree of electrolysis of an aqueous solution depends on the concentration, ions present and nature of the electrodes.

Water could be oxidised or reduced. Its presence complicates the electrolysis of aqueous solutions.

#### Anode Reaction (Oxidation)

The anions that are easily oxidised instead of water include chloride ion ( $\text{Cl}^-$ ), bromide ion ( $\text{Br}^-$ ) and iodide ion ( $\text{I}^-$ ). Polyatomic ions are not discharged. An example of such ions include sulphate ion ( $\text{SO}_4^{2-}$ ). Instead, oxygen gas is evolved due to oxidation of water.

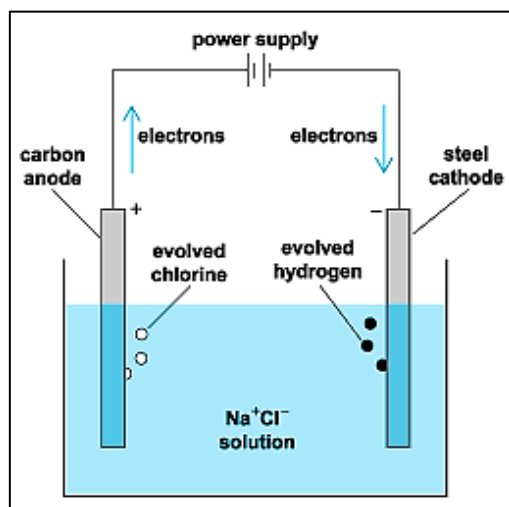
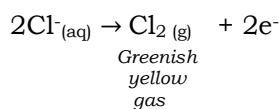
#### Cathode Reaction (Reduction)

A solution that contains cation of an element below aluminium in the Activity Series will be reduced instead of water. Cations of salt solutions of very active elements will not be reduced as water is reduced.

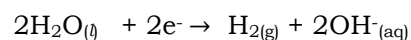
The figure below shows the electrolysis of brine, concentrated  $\text{NaCl}_{(\text{aq})}$ ; it is an important industrial application as it produces much needed chlorine. The salt solution contains the electrolyte, sodium ions, chloride ions and water.

Water is reduced at the cathode instead of sodium ions as it was **easier** to reduce; at the anode, chloride ion was easier to oxidise than water.

### Anode Reaction



### Cathode Reaction



**Note:** Hydroxide ion formed ( $\text{OH}^{-}$ ) is picked by sodium ion in the solution producing sodium hydroxide,  $\text{NaOH}$

Source: <http://look4chemistry.blogspot.com>

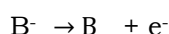
### Electrolysis of a molten salt

Electrolysis of a molten salt is less complicated than electrolysis of ionic solutions as it involves only two species. In solutions, the presence of water may increase the number of species that may be oxidised or reduced.

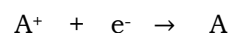
The power supply is connected to the electrodes. The positive electrode (anode) attracts the negatively charged ions (anions) in the electrolyte and the negative electrode (cathode) attracts the positively charged ions (cations) in the electrolyte.

At the anode, the anion will lose electron(s) and be oxidised to the respective element. At the cathode, the cation will gain electron(s) and be reduced to the respective metal.

*Anode Reaction*



*Cathode Reaction*

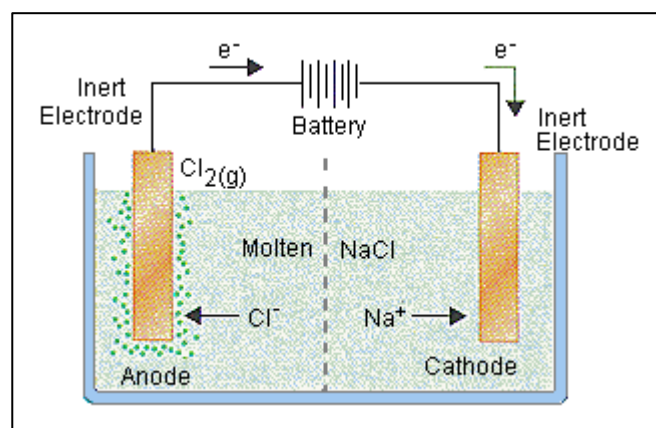
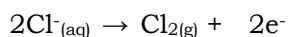


The net effect is the decomposition of the ionic compound forming two separate elements.

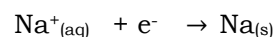
### Example

### Electrolysis of molten $\text{NaCl}$

### Anode Reaction



### Cathode Reaction



Source: <http://chemed.chem.purdue.edu>

Sodium ions ( $\text{Na}^{+}$ ) gain electrons and are reduced to sodium metal at the cathode. Chloride ions move to the anode and are oxidised by losing electrons to form chlorine gas. The anode is separated from the cathode by a diaphragm. It prevents the greenish yellow chlorine gas

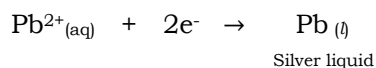
produced at the anode from reacting with the liquid silver sodium metal formed at the cathode. This is because the reaction is explosive.

## Other Examples

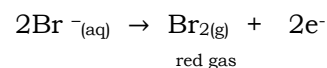
### 1. Electrolysis of molten Lead Bromide Using Carbon Electrodes

Lead bromide melts at 373°C to form molten lead bromide (PbBr<sub>2</sub>), which is made up of mobile lead ions and bromide ions. During electrolysis, lead ions are attracted to the cathode and are reduced to silver lead metal and bromide ions are attracted to the anode forming bromine, a red coloured gas.

#### Cathode Reaction



#### Anode Reaction

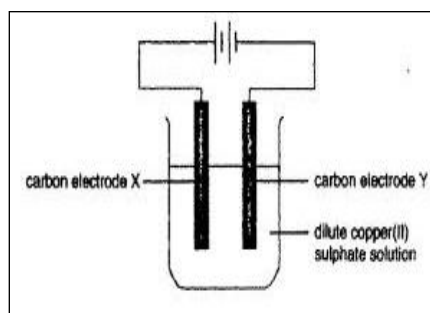
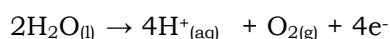


### 2. Electrolysis of Copper Sulphate solution Using Carbon

Copper sulphate solution contains the electrolytes; copper ions, sulphate ions and water.

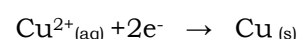
Copper ions are attracted to the cathode and are reduced to reddish brown copper metal. Sulphate ions are attracted to the anode. However, water is oxidised as sulphate ions cannot be oxidised. It releases oxygen gas, so a colourless gas is formed at the anode.

#### Anode Reaction



Source: [www.gcestudybuddy.com](http://www.gcestudybuddy.com)

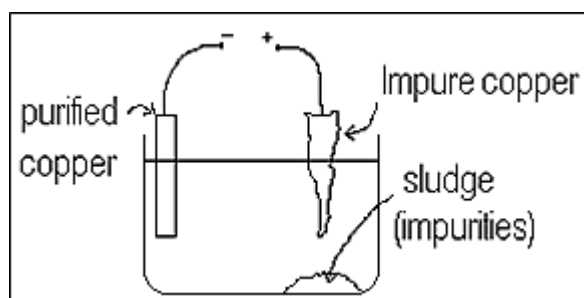
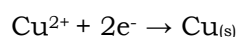
#### Cathode Reaction



### 3. Electrolysis of Copper sulphate solution using copper metal electrodes

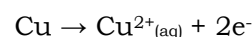
This set up is used in the *industrial production of copper metal*. In the figure below, impure copper is the anode and pure copper is the cathode.

#### Cathode Reaction



Source: <http://faculty.chem.queensu.ca/>

#### Anode Reaction



#### Did you know?

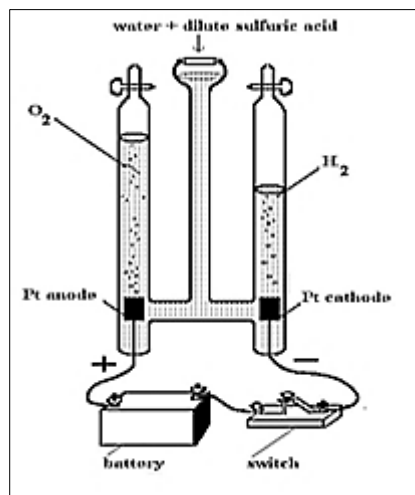
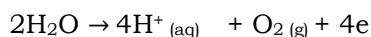
Impure gold present in the copper anode forms a sludge/deposit below the anode. This process is called electroplating. It is used to either protect a metal surface from corroding by coating it with a more active or less active metals or to cover a cheaper metal with an expensive metal. E.g. Tin coated iron cans, galvanised iron, gold/silver coated iron chain. The metal to be plated or coated is the anode; the metal to be used as coat is the cathode.

#### 4. Electrolysis of water using carbon electrodes

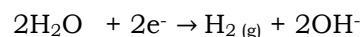
Hydrogen gas is used as rocket fuel as it burns explosively (self-combust) in oxygen releasing a lot of energy. The product formed is water. The *Energy Industries* around are researching how it can be used as fuel for vehicles, especially buses as its product, water is environment friendly. The risk is the explosive nature of the reaction.

Electrolysis of water is used industrially to produce hydrogen gas. Water reduced at the cathode produces hydrogen gas and water oxidised at the anode forms hydrogen ions.

##### Anode Reaction



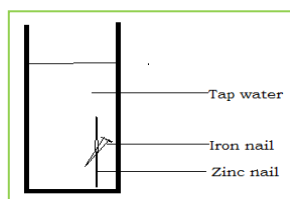
##### Cathode Reaction



Source: [www.uq.edu.au](http://www.uq.edu.au)

#### Exercise 3.2.4

1. Why is graphite used as an electrode?
2. Why does an open can rust faster than an unopened can?
3. Is water an electrolyte? Give a reason for your answer.
4. Suggest why the anode of an electrolytic cell is positive.
5. Suggest why the cathode of an electrolytic cell is negative.
6. How does electroplating prevent corrosion?
7. Study the set up given below. The iron nail is touching the zinc nail as they are tied together.



- i) Will the iron corrode? Give a reason for your answer.
  - ii) The zinc nail was removed. State an observation that you will make after a few days.
8. Suppose you are given the following materials:
- A silver table spoon, copper electrodes, electrical wires, 1L beaker, 500 mL 1.0 moles/litre copper sulphate solution, AC/DC Transformer*
- i. Draw a diagram of the electrolytic cell you would construct using the given materials in order to plate the silver spoon with copper.
  - ii. Why was the silver spoon placed as the cathode?
  - iii. Will the copper sulphate solution decolourise? Give a reason for your answer.
9. Draw a concept map on the sub-strand "Types of Reactions".

## 3.3 Acids, Bases and Salts

### Achievement Indicators

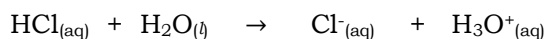
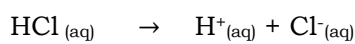
Upon completion of this sub-strand, students will be able to:

- ✓ Describe the differences between acids and bases.
- ✓ Illustrate the differences between diluted and concentrated solutions, weak and strong acids and alkalis using the pH scale.
- ✓ Research and explain the corrosive nature of acids and bases and the safety precautions when handling them.
- ✓ Study and show examples of neutralization reactions with balanced chemical and ionic equations.
- ✓ Investigate and prepare solutions of required concentrations in g/L.
- ✓ Explore and perform a simple acid-base titration considering the role of indicators.
- ✓ Work out and write the formula, chemical names, common names and uses of some salts.
- ✓ Determine and discuss some properties of the salts.
- ✓ Evaluate and present the percentage of water of crystallization in a hydrated salt.
- ✓ State and illustrate the solubility rules of salts and use it to predict the formation of a precipitate.

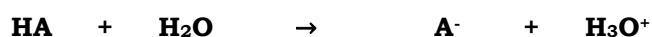
### Acids

Acids form hydrogen ions /protons ( $H^+$ ) or hydronium ions ( $H_3O^+$ ) in solution.

### Examples



In general:



### Properties of Acids

- Have a low pH (below 7).
- Neutralises bases to form water and a salt.
- Have a sour taste.
- Changes blue litmus to red.
- React with many metals to produce hydrogen gas.

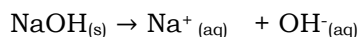
### Examples of Acids:

- ✓ Hydrochloric acid (HCl), sulphuric acid ( $H_2SO_4$ ) and nitric acid ( $HNO_3$ )
- ✓ Citric acid (in orange juice or lemon juice)
- ✓ Acetic acid (in vinegar)
- ✓ Phosphoric acid (in Coke)
- ✓ Ascorbic acid (in vitamin C tablets)
- ✓ Uric acid (in urine)
- ✓ Stearic and lauric acid (in cosmetics)

## Bases/Alkalis

Form hydroxyl ions (OH<sup>-</sup>) in solution.

### Example



In general:



### Properties of Bases/Alkalis

- Have a high pH (above 7).
- Changes red litmus to blue.
- Neutralises acids to form water and a salt.
- Have a bitter taste.
- Feels slippery.

### Examples of Bases/Alkalis

- ✓ Ammonia
- ✓ Calcium hydroxide (caustic lime/lime water)
- ✓ Lithium hydroxide
- ✓ Potassium hydroxide (caustic potash)
- ✓ Sodium hydroxide (caustic soda)
- ✓ Many bleaches, soaps, toothpastes and cleaning agents
- ✓ Window cleaners may contain ammonia

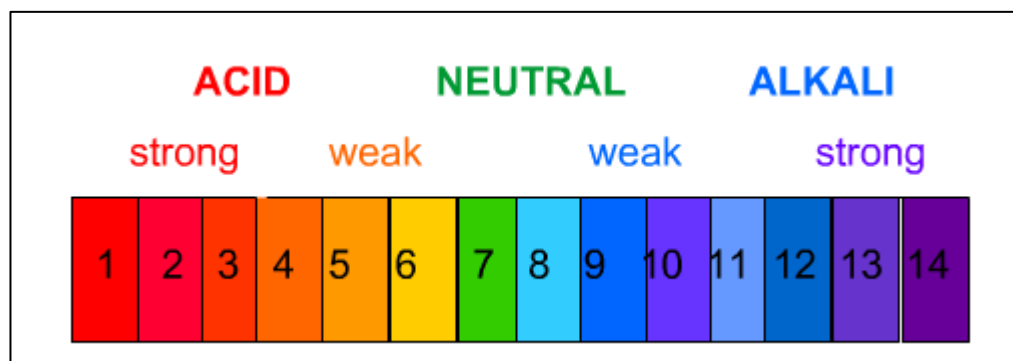
#### Exercise 3.3.1

#### Fill in the blanks

Acids have a \_\_\_\_\_ taste, react with metals to produce \_\_\_\_\_ gas, turn \_\_\_\_\_ different colours according to pH and are \_\_\_\_\_ because their water solutions conduct electricity. On the other hand, bases have a \_\_\_\_\_ taste, feel \_\_\_\_\_, turn \_\_\_\_\_ different colours according to pH and are \_\_\_\_\_ because their water solutions conduct electricity.

## The pH Scale

pH is a measure of how acidic or how alkaline a solution is. In other words, it is a scale for measuring hydrogen ion concentration. In the term pH, p stands for '*potenz*' in German meaning power.



Source: <http://hlashy12.weebly.com>

## Measuring the pH of a solution using the Universal Indicator

The *Universal Indicator* is in liquid form and changes colour at different pH.

Steps in measuring the pH of a solution using the Universal Indicator.

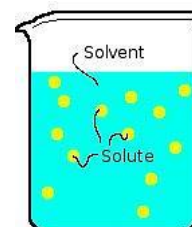
1. Take the test solution in a test tube. If there is a solid substance such as a salt then dissolve the solid by adding distilled water to it.
2. Place a drop of the Universal Indicator using a fine dropper into the solution.
3. Observe the colour produced and match it with the different colour shades of the standard colour pH chart.
4. Note down the pH of the colour chart that matches most closely with the colour produced on the pH paper.

## Strengths of Acids and Bases

Acids	Bases
<b>Strong Acids</b>	<b>Strong Bases</b>
Are acids which completely dissociates into ions in aqueous solution. Examples include: HCl, H <sub>2</sub> SO <sub>4</sub> , and HNO <sub>3</sub> .	Are bases which completely dissociates in aqueous solution. Examples include: NaOH and KOH.
<b>Weak Acids</b>	<b>Weak Bases</b>
Are acids which only partially ionise or dissociates in aqueous solution. Examples include: Acetic acid (CH <sub>3</sub> COOH) and carbonic acid (H <sub>2</sub> CO <sub>3</sub> ).	Are bases which only partially ionise or dissociates in aqueous solution. Examples include: Ammonia (NH <sub>3</sub> ) and ammonium hydroxide (NH <sub>4</sub> OH)

## Solutions

In the laboratory, we will often deal with solutions. A solution consists of a liquid (the solvent) with a substance (the solute) dissolved in it. We counter many solutions in our everyday life. For example, milk is a solution consisting of water (the solvent) with lactose and salts dissolved in it. Ocean water is another type of solution that has many dissolved ions in water.

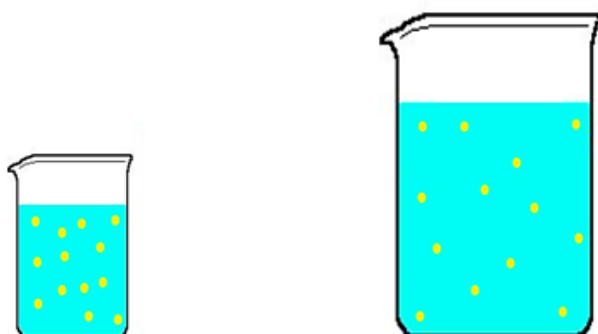


A solution

## Dilute and Concentrated solutions

**Dilution** involves adding more solvent to a solution so that the concentration of the solute becomes lower. The total number of solutes in the solution remains the same after dilution, but the volume of the solution becomes greater, resulting in a lower concentration.

### Example

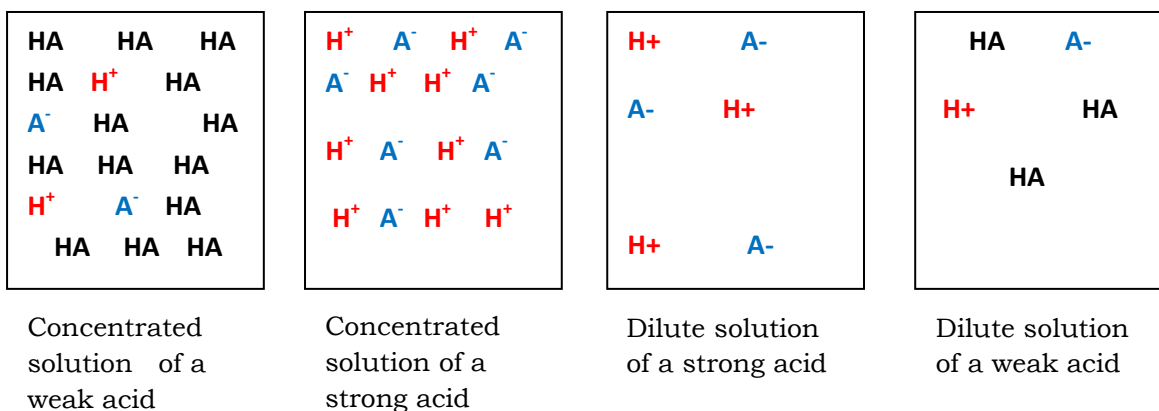


A concentrated solution

A dilute solution

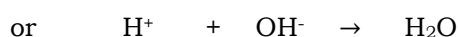
In this illustration, the solutes are shown as yellow dots and the solvent as blue. The 1 L beaker on the left shows the initial concentration, which we might represent as 13 dots/L. The beaker on the right is the result of dilution of the left beaker. More solvent is added so that the solution's total volume is 3 L. As a result, the concentration of the diluted beaker is 13 dots/3 L or 4.3 dots/L.

## A Summary of Concentrated/Dilute solution of Strong/Weak Acid



### Neutralization reaction – the formation of a salt

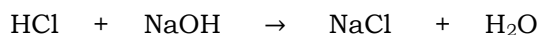
Acids dissolve in water to produce solutions containing hydronium ions/hydrogen ions ( $\text{H}_3\text{O}^+/\text{H}^+$ ). Bases dissolve in water to produce hydroxide ions ( $\text{OH}^-$ ). When an acidic solution and a basic solution are mixed, hydroxide ions react with hydronium ions to produce water molecules:



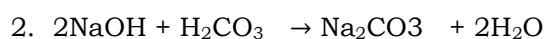
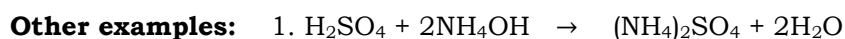
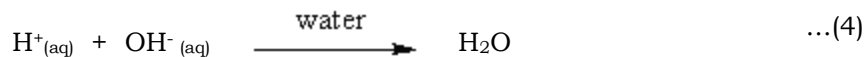
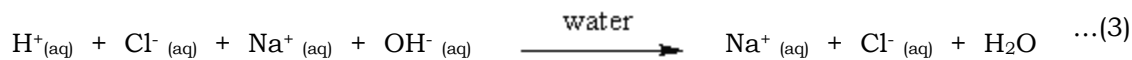
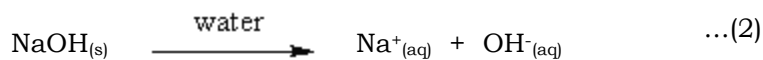
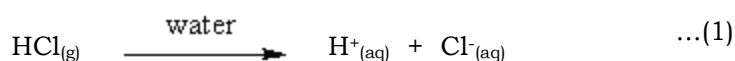
The other substance formed is the **salt**.

In general:  $\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$

#### Example



The following shows the ionic equations:



## Common salts, their names and important uses

Chemical Name	Formula	Common name	Uses
Ammonium chloride	$\text{NH}_4\text{Cl}$	Salt ammoniac	Electrolyte in dry cell
Copper sulphate	$\text{CuSO}_4$	Blue stone	Drying agent, fungicide
Calcium carbonate	$\text{CaCO}_3$	Marble/chalk	Blackboard chalk, dietary supplement
Calcium sulphate	$\text{CaSO}_4$	Plaster of Paris	Making Plaster of Paris, food additive
Calcium oxychloride	$\text{CaOCl}_2$	Bleaching powder	Commercial bleach, water treatment
Magnesium sulphate	$\text{MgSO}_4$	Bitter salt	Laxative, as fertilizer
Sodium chloride	$\text{NaCl}$	Salt	Flavor enhancer, preservative
Sodium hydroxide	$\text{NaOH}$	Caustic soda	Making paper, manufacturing textile
Sodium sulphate	$\text{Na}_2\text{SO}_4$	Glaubers salt	Dyeing textile
Sodium bicarbonate	$\text{NaHCO}_3$	Baking soda	Antacid, soothe insect bite

**Note:** There are a number of uses of the salts. Given in the table are the common ones.

### Properties Salts

- Salts are product of an acid-base reaction.
- Many salts are soluble:
  - For this reason, most drugs are salts (most often sodium). This gets the drug into the bloodstream as all sodium compounds are soluble
  - If a drug cannot be turned into a salt, it will not dissolve in water and is useless to a patient as it will not enter the bloodstream.
- Salts are ionic compounds
- Salts are hygroscopic: they absorb water very well to form hydrates.

### Summary



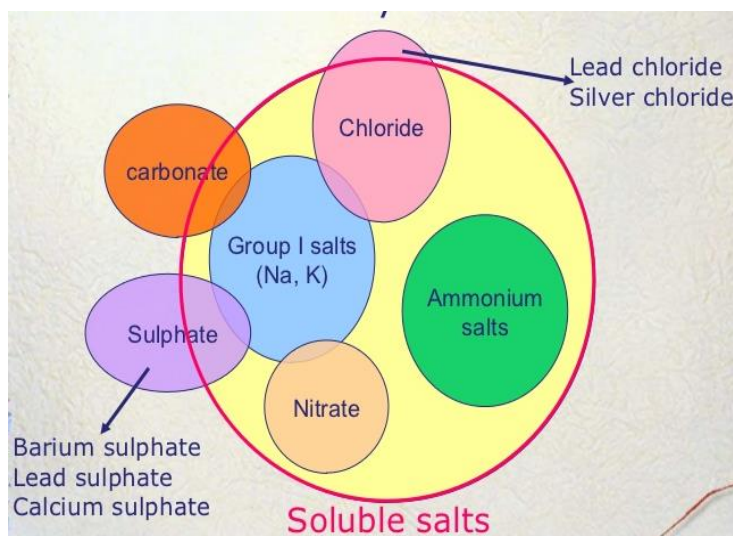
## Preparing soluble and insoluble salts

To enable one to decide whether a salt/precipitate will form if two solutions are mixed together, one needs to know the following solubility rules:

### Solubility Rules – solubility of the salt in water

- ❖ All nitrates are **soluble**
- ❖ All sulphates are **soluble** except barium,  $\text{Ba}^{2+}$ , calcium,  $\text{Ca}^{2+}$  and lead,  $\text{Pb}^{2+}$
- ❖ All chlorides are **soluble** except those of silver,  $\text{Ag}^+$  and lead,  $\text{Pb}^{2+}$ .
- ❖ All carbonates are **insoluble** except those of  $\text{K}^+$ ,  $\text{Na}^+$  and  $\text{NH}_4^+$ .
- ❖ All hydroxides are **insoluble** except those of  $\text{K}^+$ ,  $\text{Na}^+$  and  $\text{NH}_4^+$ .

### Summary:



## Preparation of Salts

### 1. Reaction of Acids

Salts can be prepared by reacting acids with; a metal, an insoluble base or a carbonate.

#### Example

- Acid + Metal  $\rightarrow$  Salt + Hydrogen  
E.g.  $2\text{HCl} + \text{Zn} \rightarrow \text{ZnCl}_2 + \text{H}_2$
- Acid + Insoluble bases  $\rightarrow$  Salt + Water  
E.g.  $2\text{HNO}_3 + \text{CuO} \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{H}_2\text{O}$
- Acid + Carbonate  $\rightarrow$  Salt + Water + Carbon dioxide  
E.g.  $\text{H}_2\text{SO}_4 + \text{CaCO}_3 \rightarrow \text{CaSO}_4 + \text{H}_2\text{O} + \text{CO}_2$

**The following steps show the procedures to be followed in obtaining a salt:**

#### 1. Dissolving

- ✓ Warm the acid.
- ✓ Add solid base/carbonate/metal till no more can be added.

#### 2. Filtration

This is done to remove the un-reacted solid.

- ✓ Heating till saturation – heat the filtrate to evaporate off half the original volume to get a saturated solution.

#### Note:

Residue = solid  
Filtrate = salt solution

### 3. Crystallization

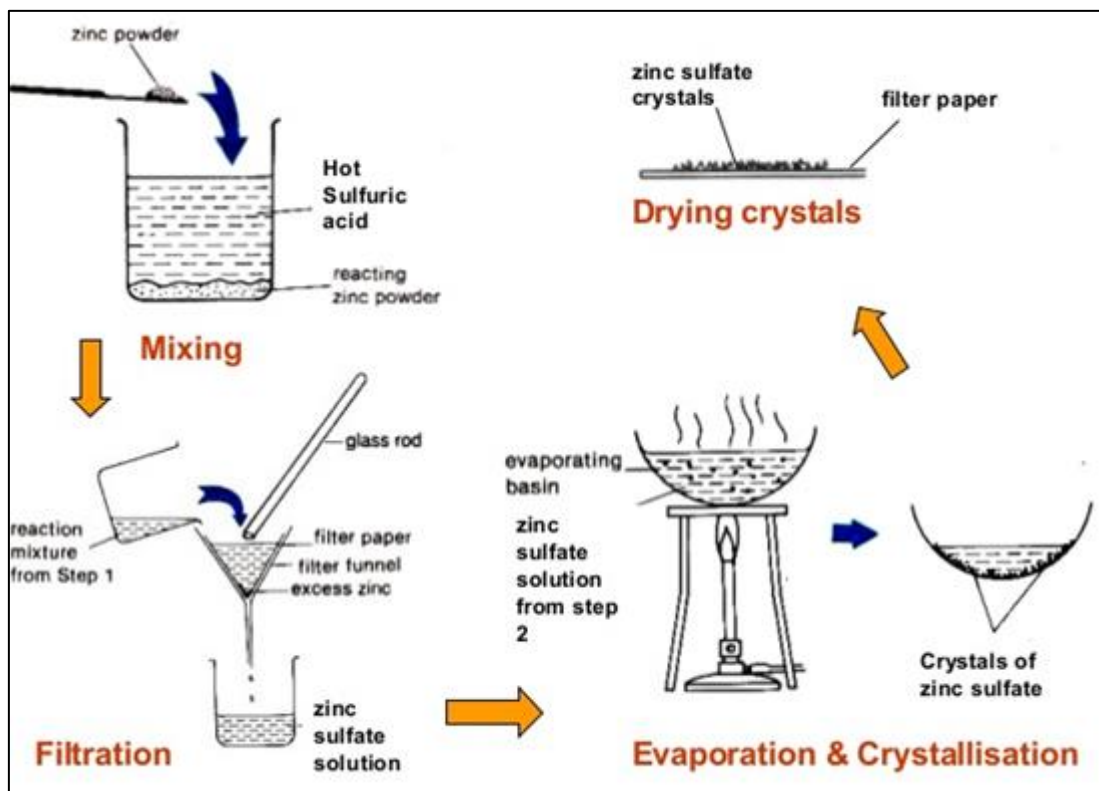
- ✓ This allows the hot saturated solution to cool for pure crystals to grow

### 4. Collecting the crystals

- ✓ Collect crystals formed by the filtrate. Dry the crystals between the pieces of dry filter paper.

#### Example

#### Preparing Zinc sulphate crystals using H<sub>2</sub>SO<sub>4</sub> and Zinc powder



Source: <http://www.slideshare.net>

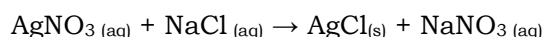
## 2. Precipitation

This is another way of getting a salt. An insoluble salt can be made by mixing two solutions of soluble salts in a process called precipitation.

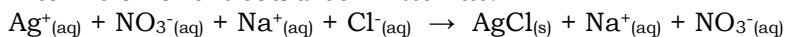
- ✓ One solution contains the first required ion and the other solution contains the second required ion.
- ✓ The two solutions of **soluble** compounds are mixed together so the **insoluble** salt (precipitate) is formed.
- ✓ The precipitated salt can then be filtered out with a filter funnel and paper.
- ✓ The collected solid is washed with distilled water to remove any remaining soluble salt impurities and carefully removed from the filter paper to be dried.

#### Example

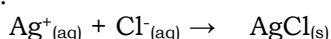
Silver chloride is made by mixing solutions of silver nitrate and sodium chloride.



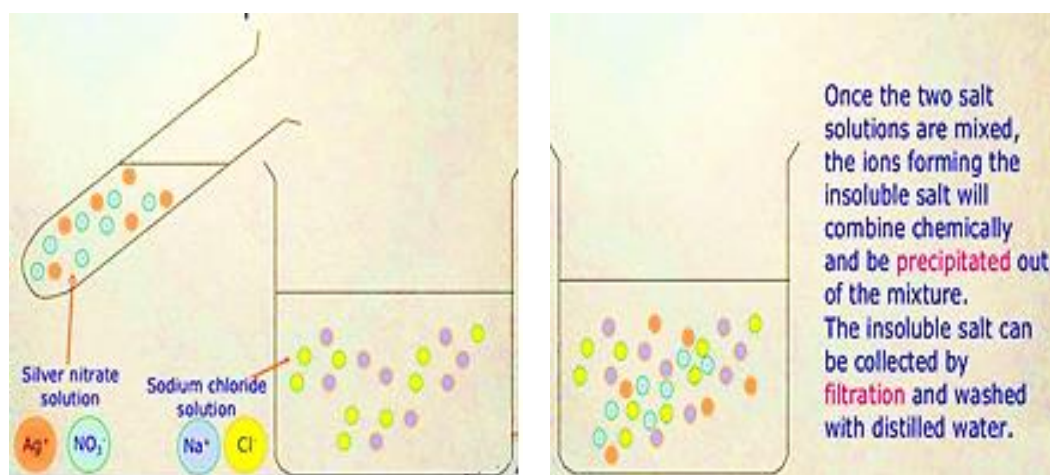
In terms of ions it could be written as:



**Note:** Spectator ions are nitrate (NO<sub>3</sub><sup>-</sup>) and sodium (Na<sup>+</sup>) which do not change at all. So the net ionic equation is simply:



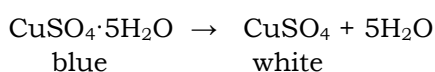
An illustration of the formation of AgCl is shown below



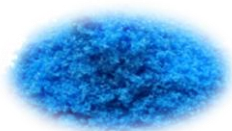
### Anhydrous and Hydrated salts

Hydrates are generally crystalline salts which contain a specific ratio or definite proportion of water in a compound. This water is called the **water of crystallization**. Hydrates are formed by slowly evaporating the water from the salt solutions. An example of a hydrate is copper sulphate pentahydrate crystals (CuSO<sub>4</sub>·5H<sub>2</sub>O) crystals.

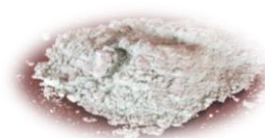
Anhydrous compounds are formed when hydrates are strongly heated and the water of hydration is driven off. For example, copper sulphate pentahydrate can be converted into anhydrous copper sulphate. When the blue crystalline copper sulphate pentahydrate is heated it changes to a white, powdery, anhydrous salt.



Diagrammatically:

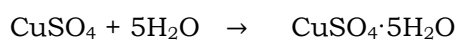


Hydrated copper sulphate crystals

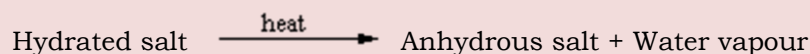


Anhydrous (dry) copper sulphate solid

It is also possible to reverse the above process by adding water to the copper sulphate powder. This is shown in the equation below:



Generally:



### Calculating the water of crystallization (w.o.c)

$$\text{W.O.C} = \frac{\text{Molecular mass of water}}{\text{Formular mass of hydrated salt}} \times \frac{100}{1}$$

#### Example

Calculate the percentage w.o.c in  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .

#### Solution

$$\begin{aligned} \text{W.O.C} &= \frac{\text{Molecular mass of water}}{\text{Formular mass of hydrated salt}} \times \frac{100}{1} \\ &= \frac{90}{250} \times \frac{100}{1} \\ &= \mathbf{36\%} \end{aligned}$$

#### Common hydrated salts:

- Sodium carbonate crystals (washing soda) -  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
- Barium iodide crystals -  $\text{BaI}_2 \cdot 7\text{H}_2\text{O}$
- Magnesium sulphate crystals -  $\text{MgSO}_4 \cdot 5\text{H}_2\text{O}$

#### Solutions and Concentration

Concentration is a measure of the amount of dissolved solute in a given volume of solution.

The formula for calculating the concentration of a solution is:

$$c = \frac{m}{V}$$

Where  $c$  = concentration in gram per litre (g/L).  
 $m$  = mass of solute in grams  
 $V$  = volume of solution in litres.

**Note:** Grams per litre represent the mass of solute (in grams) divided by the volume of solution (in litres).

#### Example

Calculate the concentration (in g/L) of the following solution:

10 g of NaOH dissolved in enough water to make 2 L of the solution.

#### Solution

$$\begin{aligned} c &= \frac{m}{V} \\ &= \frac{10 \text{ g}}{2 \text{ L}} \\ &= 5 \text{ g/L} \end{aligned}$$

### Exercise 3.3.2

- What is the percentage water of crystallisation in the following compounds?
  - $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
  - $\text{BaI}_2 \cdot 7\text{H}_2\text{O}$
  - $\text{MgSO}_4 \cdot 5\text{H}_2\text{O}$
- A solution with a pH of 9 is:
  - a neutral solution
  - an acidic solution
  - an alkaline solution
  - a dilute solution
- Which of the following is a strong base?
  - NaOH
  - $\text{NH}_3$
  - $\text{Na}_2\text{CO}_3$
  - $\text{NH}_4\text{OH}$
- Which of the following is a weak acid?
  - $\text{H}_2\text{SO}_4$
  - $\text{H}_2\text{CO}_3$
  - $\text{HNO}_3$
  - HCl
- What is the difference between a strong acid and a weak acid?
- What is the difference between a strong acid and a concentrated acid solution?
- Determine the concentration in each of the following cases.
  - When 20 g of NaOH is dissolved in enough water to make 1 litre of NaOH solution.
  - When 45 g of glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$  is dissolved in enough water to make 0.5 litres of glucose solution.
  - When 116 grams of KF is dissolved in enough water to make 4 L of KF solution.
- In which of the following reaction will a precipitate form?
  - KOH reacting with HCl
  - $\text{AgNO}_3$  reacting with NaCl
- Give the formula of the salts below:
  - Copper sulphate
  - Sodium bicarbonate
  - Sodium sulphate
- Complete and balance the reaction equations below.
  - $\text{HCl} + \text{Zn} \rightarrow \text{_____} + \text{H}_2$
  - $\text{HNO}_3 + \text{_____} \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{H}_2\text{O}$
  - $\text{H}_2\text{SO}_4 + \text{CaCO}_3 \rightarrow \text{_____} + \text{H}_2\text{O} + \text{CO}_2$

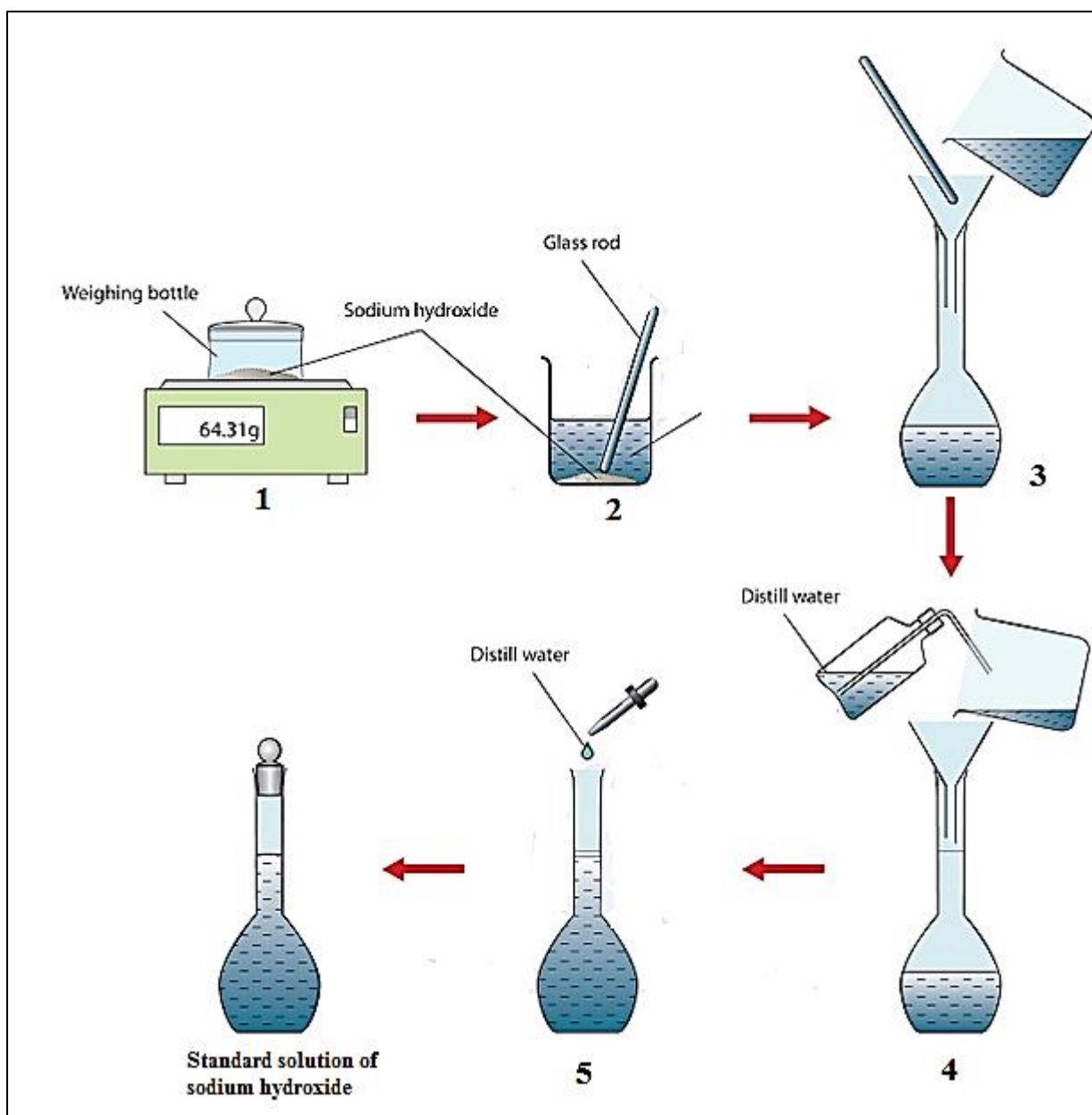
## Preparing a Standard Solution

1. Take a watch glass (or any weighing container) and place it on the electric balance. Tare the balance (set it to zero). Carefully weigh out the required mass of the salt.

Note: If an electronic balance is unavailable then correctly take the mass (in grams) using any weighing balance.

2. Transfer the salt to a beaker. Add distilled water from a wash bottle to dissolve it. Use some water to rinse all the salt off the watch glass. Do this at least twice. Stir with a glass rod until the entire solid is dissolved.
3. Transfer the solution to the volumetric flask. Use more water from the wash bottle to rinse out the beaker and the glass rod. Do this at least twice.
4. Add water to just below the line on the volumetric flask. Add the final drops with a pasture pipette or dropper to ensure that the bottom of the meniscus is on the line.
5. Put the lid on the flask and turn the flask over a couple of times to mix the solution. Label the solution with your name, the date, and the contents.

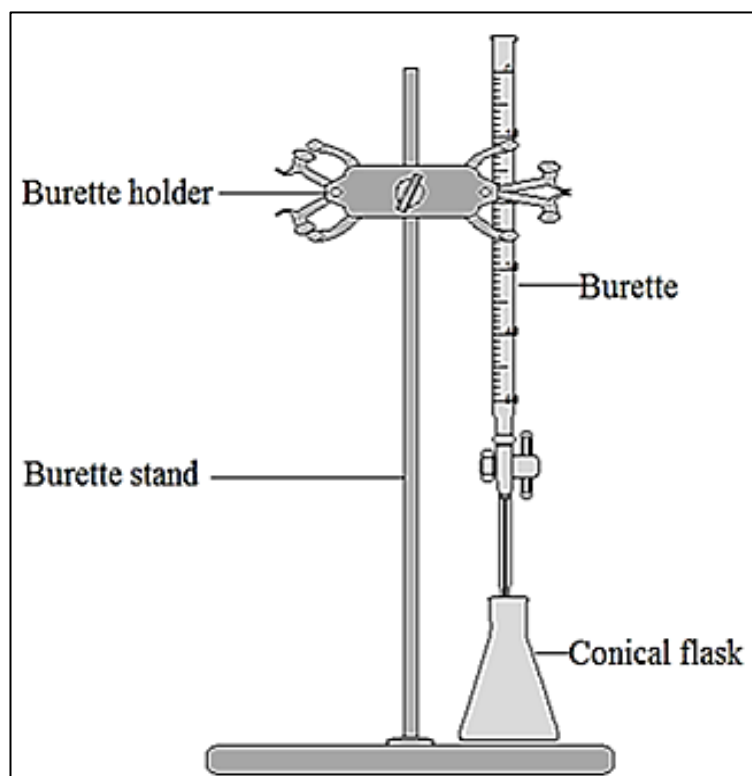
### Example: Preparation of a standard solution of sodium hydroxide



Source: <http://spmchemistry.onlinetuition.com.my>

## Titration

Titration is a technique in which a solution of known concentration is used to determine the unknown concentration of another solution. The solution whose concentration is accurately known is called a standard solution. Titration involves the slow addition of one solution to another solution until the reaction complete or reaches neutralization. This happens when the indicator changes colour.



**A titration set-up**

## Indicators

These are substances which indicate the acidic or basic nature of a solution by their colour change.

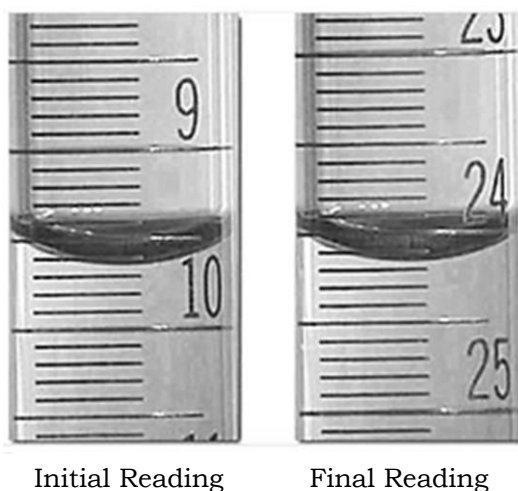
During a titration reaction, two to three drops of a suitable indicator is added to the conical flask that contains the solution of known concentration to show when neutralisation will take place. As the solution from the burette is added to the conical flask, the indicator changes colour. At the end point the desired colour is noted.

Common indicators used in the laboratory:

Indicator	Colour		
	Basic	Endpoint	Acidic
<b>Phenolphthalein</b>	Pink	Light/faint/pale pink	Colourless
<b>Methyl orange</b>	Yellow	Orange	Red
<b>Bromothymol blue</b>	Blue	Green	Yellow

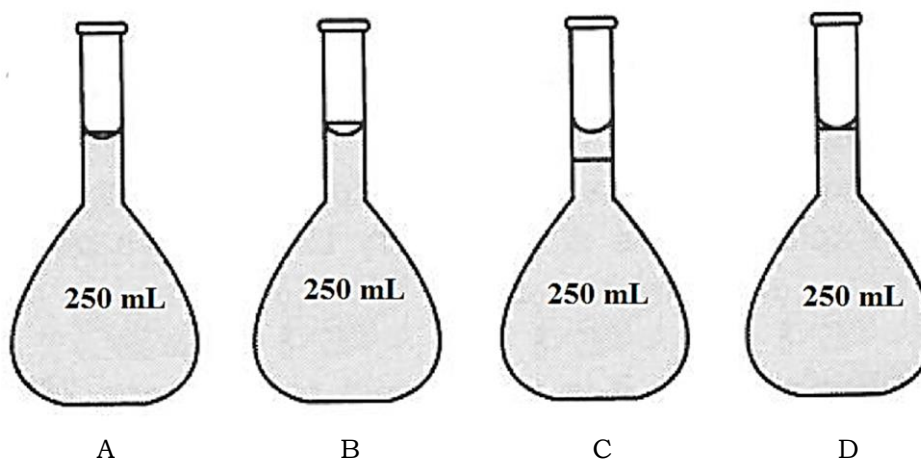
### Exercise 3.3.5

1. Define/describe the following terms.
  - i. Titration
  - ii. Standard solution
  - iii. Indicator
2. How many grams of KOH is used to make 2.5 L of a 50 g/L solution of KOH.
3. Preparing a standard solution requires a series of steps. Arrange the steps given below in the order you would follow to prepare a standard solution of sodium carbonate.
  - A. Rinse the funnel and the beaker with distilled water.
  - B. Measure the mass of sodium hydroxide.
  - C. Pour the solution into a standard flask.
  - D. Dissolve the sodium hydroxide in 100 mL of water.
  - E. Make the volume to exactly 250 mL.
4. Given below are the initial and final readings of a burette during a titration experiment.



Source: <http://jr.stryker.tripod.com>

- i. Explain why it is important to have the meniscus at eye level when reading a burette.
  - ii. Calculate the titre used in the titration in litres.
5. Which standard flask contains exactly 250 mL of standard solution?



**Strand Outcome:** Investigate the physical properties of materials and how these influence the way they are used.

## Sub-strands

- 4.1 Metals and Non-Metals
- 4.2 Oxides of Metals and Non-metals
- 4.3 Organic Substances

### 4.1 Metals and Non-Metals

#### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Examine and explain the physical and chemical properties and uses of selected metals.
- ✓ Study and describe the production and uses of iron.
- ✓ Confirm and discuss the physical and chemical properties and uses of selected non-metals.
- ✓ List and describe the allotropes of sulphur, carbon, phosphorous, oxygen and their structures.
- ✓ Prepare samples of the allotropes of sulphur, state their uses and explain their properties.
- ✓ Consider and present on some common and traditional uses of Carbon.

#### Metals

Metals are group of elements which includes: Alkali metals (Group I elements), Alkaline earth metals (Group II elements), aluminium and transition elements. These elements are generally shiny/lustrous, good conductors of electricity, malleable, ductile and some are magnetic. Metals are commonly found in the earth's crust as metal deposits or ores.

The table below shows some common metals, their physical features and common uses.

<b>Metal</b>	<b>State</b>	<b>Density (g/cm<sup>3</sup>)</b>	<b>mpt (°C)</b>	<b>bpt (°C)</b>	<b>Natural Source</b>	<b>Use</b>
 <p><b>Copper</b></p>	Solid	8.96	1083	2562	Copper is present in the earth's crust as copper sulphide, chalcopyrite and chalcocite	Very good conductor of electricity so it is used in electrical wiring for home appliances, telecom lines and cooking pot.
 <p><b>Lead</b></p>	Solid	11.34	327.5	1749	Found in ore with silver, copper and zinc. Main mineral is galena which is iron sulphide.	Ballast keel of sailboats, scuba diving weight belts, lead – acid car batteries.
 <p><b>Calcium</b></p>	Solid	1.54	842	1484	Calcium carbonate - Chalk, marble, limestone.  About 4.2% of the earth's crust contains calcium.	White wash, cleaning powder, toothpaste and antacid.
 <p><b>Zinc</b></p>	Solid	7.14	419	907	Very low concentration of zinc is present in the earth's crust, chalcophile	Galvanise iron, zinc alloy, brass is used widely
 <p><b>Aluminium</b></p>	Solid	2.38	659	2470	Third most abundant element in the earth's crust. Common ore bauxite, aluminium oxide.	Cooking foil, electric wire for power lines
 <p><b>Sodium</b></p>	Solid	0.97	97.8	882.9	Sixth most abundant element in the earth's crust. Most common source is sea water. It is found as sodium chloride.	Plant micronutrient, needed for neuron function and osmoregulation in cells.

 <p><b>Magnesium</b></p>	Solid	1.74	651	1091	Present in sea water as magnesium ions, $Mg^{2+}$ . Common ores are magnesite, dolomite and mineral water	Alloy is used in aerospace industry
 <p><b>Mercury</b></p>	Liquid	13.53	-38.82	356.76	Very rare element in the earth's crust. Commonly known as <i>quicksilver</i>	Used in mercury thermometers
 <p><b>Iron</b></p>	Solid	7.87	1538	2862	Sixth most abundant element and 2 <sup>nd</sup> most abundant metal. Common ores are iron oxides, haematite ( $Fe_2O_3$ ) and magnetite ( $Fe_3O_4$ ).	Most commonly used metal. It's available cheaply and has a lot of strength so it is used in the building industries, making tools and body of ships.

#### Did you know?

- The use of metals depends on its physical properties. For example, copper is used in electrical appliances as it is a good conductor of electricity.
- Mercury readily combines with aluminium to form amalgam, destroying the aluminium oxide layer and exposing the metal to corrosion. Possession of mercury product is banned from aboard the aeroplane.

#### Exercise 4.1.1

1. Name three general features of metals.
2. Suggest two reasons why gold is a metal.
3. Why are our daily eating utensils made up of tin or iron or aluminium instead of gold and silver?
4. Why is mercury used in thermometers?
5. Copper is a better conductor than aluminium, however is more widely used in electrical wiring than copper. Suggest two reasons to explain this.
6. Why is sodium kept under oil or paraffin in the laboratory?

## The Activity Series

It is the order of oxidation of metals from the most easily oxidised to the least. Group 1 metals are the most easily oxidised and heavy metals are the least. The decreasing order of reactivity is shown in the figure below:

Potassium	Most reactive	K	
Sodium	↓	Na	
Calcium		Ca	
Magnesium		Mg	
Aluminium		Al	
Zinc		Zn	
Iron		Fe	
Tin		Sn	
Lead		Pb	
Copper		Cu	
Silver		Ag	
Gold		Au	
Platinum		Least reactive	Pt

## Chemical properties of metals

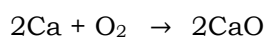
Exposed metal surface will react with air, water and dilute acids. The rate of reaction of different metals is determined by the activity series. Its reactivity may also determine the uses of the metal.

### Reactions with air

Exposed metal react with oxygen in air to form metal oxides. Some of these metal oxides may combine with water vapour in the atmosphere to form hydroxides.

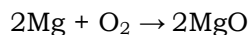
#### Calcium

Calcium, when exposed to air easily develops a greyish white layer of calcium oxide over the metal. Through strong heating, it burns with a small white flame to form a white solid, calcium oxide.



#### Magnesium

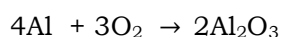
A layer of greyish black magnesium oxide will form on the metal surface after long period of exposure to air. Magnesium burns readily in oxygen with a brilliant flame giving a white smoke and a white powder which are both magnesium oxide.



#### Aluminium

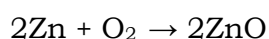
Aluminium does not corrode as a very thin layer of aluminium oxide coats the exposed metal surface preventing any further reaction.

In the absence of the oxide layer, aluminium corrodes rapidly to form aluminium oxide, a white powder.



#### Zinc

Zinc reacts very slowly with air. It burns with a green flame to form a white powder, zinc oxide.



## Iron

Iron will not corrode under the following conditions:

- ✓ Kept at room temperature in very dry conditions.
- ✓ Iron is pure.
- ✓ Placing it in boiled water and the water is covered with paraffin oil.

Iron rusts in tap water. In the presence of a lot of moisture, a film of moisture covers the surface of the metal. Rusting is rapid. Presence of dissolved ions in water (sea water) or carbonates (dissolved carbon dioxide) accelerates the rate of corrosion.

Rusting of iron is destructive as the dirty yellow flakes of iron oxide formed falls off the metal causing the structure to collapse.

Iron burns in oxygen giving off sparks to form iron oxides.

Iron surfaces can be protected from corrosion by the following methods:

- ✓ **Electroplating** – The surface of the metal is electroplated with tin, a less reactive metal. The iron surface is protected as tin does not corrode. Cutting/scratching of the coat will expose the iron to rusting conditions.
- ✓ **Galvanising** – The surface of the metal is coated with zinc, a more reactive metal. The iron surface is protected as the zinc will corrode first.
- ✓ **Greasing/Painting** – The layer of grease or paint are water and air proof so the iron surface is dry and not exposed to air. Rusting will not take place.
- ✓ **Bluing** – When strongly heated, a blue iron oxide forms. It forms a layer on the surface of the iron preventing it from corrosion. Gardening implements such as spades, knife and fork are protected by this method.

## Copper

Exposed shiny copper surface kept in dry air at room temperature turns dull but no further reaction will occur. In damp air condition, a green coat, copper hydroxide with other copper salts slowly forms on the surface of the metal.

### Displacement Reaction

More active metals will displace the less active metals from its solution. The active metal will form the solution and the ions of the less active metal will form the metal.

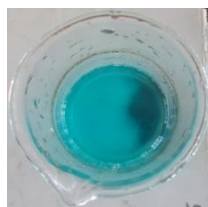
**The illustration below shows iron displacing copper from copper sulphate solution.**



Iron Powder



Copper Sulphate

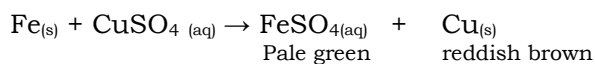


Mixture



Reddish brown deposits of copper metal

### Chemical Equation



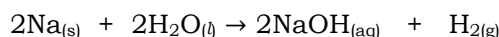
## Reaction of metals with water

Highly reactive metals such as sodium and potassium are kept under oil in the laboratory as they react with water vapour in the atmosphere. They react *explosively* with water so they are not to be reacted in the school laboratory.

Metals that react with water will release hydrogen gas and the hydroxide of the metal will form. The hydroxide is basic in nature.

### 1. Sodium

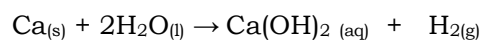
Sodium is a highly reactive metal. It is kept under oil/paraffin in the laboratory to protect the metal from reacting with moist air. Sodium reacts explosively with water to form strong alkali solution, caustic soda, sodium hydroxide and hydrogen gas is released. The hydrogen gas released instantly combust in air.



Source: [ewart.org.uk](http://ewart.org.uk)

### 2. Calcium

It reacts vigorously with water as shown in the figure below. The hydrogen released is tested by placing a lit splint at the mouth of the jar. The gas burns with a 'pop' sound. The solution formed turns phenolphthalein pink rapidly.



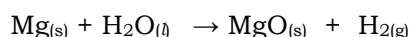
Source: [aisphysicalscience.pbworks.com](http://aisphysicalscience.pbworks.com)

### 3. Aluminium

It does not react with water as the surface of the metal is protected by coating of aluminium oxide.

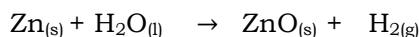
### 4. Magnesium

It reacts slowly with water. Water bubbles may form on the surface of the metal and change in colour of the phenolphthalein indicator is detected after sometime around the metal.



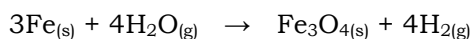
## 5. Zinc

It does not react with cold water but red hot zinc will react with steam to form zinc oxide and hydrogen gas.



## 6. Iron

It does not react with cold water, but steam passed over heated iron will form black magnetic oxide of iron.



## 7. Lead and Copper

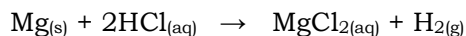
Both metals do not react with water.

### Reaction of metal with dilute acids

Potassium, sodium and calcium react explosively with dilute acids and magnesium show vigorous reaction with acids. Iron and zinc react slowly with dilute acids and aluminium may show no reaction due to the presence of the oxide coat which will first react with the acid.

The reaction of dilute acid with metals releases hydrogen gas.

### Example



### Exercise 4.1.2

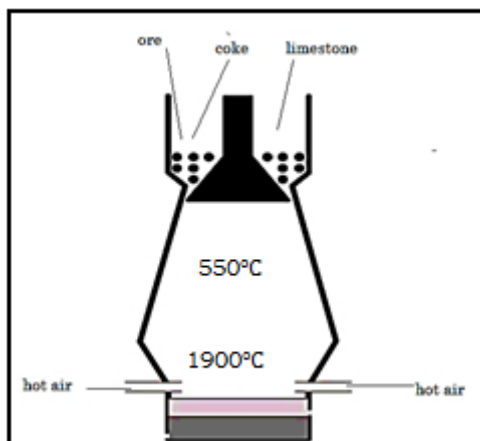
- Write equations for the reaction of sodium with water and with dilute acids.
- Metals X, Y and Z were reacted with metals and the gases collected were tested for hydrogen. The results were shown in the table below.

Observations			
Metals	Effect on water	Effect of solution on phenolphthalein	Test for Hydrogen Gas
<b>X</b>	Reacted vigorously forming bubbles rapidly	Colour changed from colourless to pink	Gas burns with a pop sound
<b>Y</b>	Very slow reaction. Bubbles formed on the surface of the metal	Solution around the metal turned pink	No change
<b>Z</b>	No reaction	No change	No change

- Predict what will happen to the size of metal X and give a reason for your answer.
- Why did the phenolphthalein change colour?
- What can you conclude from the pop sound given off?
- Explain why the pop sound was not produced in the other two test tubes.

## Production of Iron

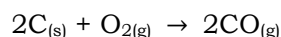
Iron is the second most abundant metal in the earth's crust and the fourth most abundant element. The two common iron ores are haematite and magnetite. Iron is industrially produced using the Blast Furnace.



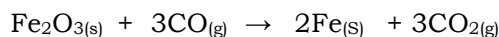
Source: [www.micromountain.com/sci](http://www.micromountain.com/sci)

The raw materials used are the iron ore, haematite (iron III oxide,  $\text{Fe}_2\text{O}_3$  and the main impurity silica,  $\text{SiO}_2$ ), limestone and coke (carbon).

Hot air at a temperature of  $1900^\circ\text{C}$  is introduced at the bottom of the furnace heated coke to form carbon monoxide.



As the hot air moves up the furnace, the temperature decreases to  $550^\circ\text{C}$ . At this temperature, iron III oxide is reduced by the carbon monoxide to molten iron.



Limestone (calcium carbonate,  $\text{CaCO}_3$ ) decomposes to form calcium oxide. The oxide will react with silica to form calcium silicate,  $\text{Ca}_3(\text{SiO}_4)_2$ , slag. Molten slag floats on the surface of molten iron; they are removed separately through separate taps.

The molten iron is run into casts called Cast iron or Pig iron. It is impure as it contains 3 – 5% carbon and small amounts of silicon and phosphorus. Pig iron is brittle and cannot be welded. It is used as fire grates, stoves and other uses.

The impurities in Pig Iron are removed by heating it with an iron ore to molten form. The oxygen contained in the ore combines with the non-metal impurities to form such as carbon monoxide. It is removed as gases or slag. The iron formed is called Wrought iron.

It is softer and less brittle than pig iron, could be welded and rust resistant. Wrought iron is used to make ornaments and farming implements and machinery. The use of mild steel containing 1% carbon had replaced wrought iron as it is cheaper.

## Alloy

It is homogenous mixture of two or more metals. The metals are mixed in definite proportions. The property of one of the metal is improved by the alloy.

The table below shows the common alloys, its components, properties and their uses.

<b>Name</b>	<b>Composition</b>	<b>Properties</b>	<b>Use</b>
<b>Brass</b>	Cu (60 - 80%), Zn (40 -20%)	More colourful, high resistant to corrosion, more malleable, low melting point	For making household utensils
<b>Bronze</b>	Cu (75-90%), Sn (25-10%)	Lightweight, high resistant to corrosion, stronger	For making coins, idols, utensils, decorative ornaments, springs, turbines, sculpture
<b>Solder</b>	Sn (50-75%), Pb (50-25%)	More tensile strength	Soldering of metals
<b>Duralumin</b>	Al (96%), Cu (3.5%), Mg (0.5%)	Strong, hard, light weight	In aircraft manufacturing
<b>Steel</b>	Fe (99%), C (1%)	High strength and hard	For making nails, screws, bridges
<b>Stainless Steel</b>	Fe (74%), Cr (18%), Ni (8%)	Highly resistant to corrosion, does not react with acid, shiny	For making cooking utensils, knives, surface for making food



### **A Case Study on Copper - Uses of Copper Compounds**

#### **1. Agricultural Uses**

Copper compounds have their most extensive employment in agriculture where the first recorded use was in 1761, when it was discovered that seed grains soaked in a weak solution of copper sulphate inhibited seed-borne fungi. By 1807 the steeping of cereal seeds in a copper sulphate solution for a limited time and then drying them with hydrated lime became the standard farming practice for controlling stinking smut or bunt of wheat, which by then was endemic wherever wheat was grown. Flour milled from bunted wheat had to be fed to animals or sold cheaply to ginger bread makers who had found a way of masking its bad taste and color with ginger and treacle. Within a few decades, so general and effective had become the practice of treating seed grains with copper sulphate that the appearance of more than a few bunted ears in a field of wheat was looked upon as a sign of neglect on the part of the farmer. So well have copper compounds controlled bunt that today this seed-borne disease is no longer of any economic importance.

The greatest breakthrough for copper salts undoubtedly came in the 1880's when the French scientist Millardet, while looking for a cure for downy mildew disease of vines in the Bordeaux district of France, chanced to notice that those vines, bordering the highways and which had been daubed with a paste of copper sulphate and lime in water in order to make the grapes unattractive to passers-by, appeared freer of downy mildew. This chance observation led to experiments with mixtures of copper sulphate, lime and water and in 1885 Millardet announced to the world that he had found a cure for the dreaded mildew. This mixture became known as Bordeaux mixture and saw the commencement of protective crop spraying.

Within a year or two of the discovery of Bordeaux mixture, Burgundy mixture, which also takes its name from the district of France in which it was first used, appeared on the scene.

Burgundy mixture is prepared from copper sulphate and sodium carbonate (soda crystals) and is analogous to Bordeaux mixture.

Trials with Bordeaux and Burgundy mixtures against various fungus diseases of plants soon established that many plant diseases could be prevented with small amounts of copper applied at the right time and in the correct manner. From then onwards copper fungicides have been indispensable and many thousands of tons are used annually all over the world to prevent plant diseases.

As a generalisation, soils would be considered copper deficient if they contain less than two parts per million available copper in the context of plant health. However, where the soil contains less than five parts per million available copper, symptoms of copper deficiency may be expected in animals. The increasing use of chemical fertilizers which contain little or no copper are denuding soils of readily available copper and creating a deficiency of the element in plants and through them in animals. Copper compounds are now being added to the ever increasing copper deficient soils either direct or in combination with commercial fertilizers. This is particularly the case where the fertilizers are rich in nitrogen and phosphorus. Animals grazing on copper deficient pastures or obtaining an inadequate amount of copper through their normal diet will benefit from mineral supplements containing copper.

Copper sulphate, because of its fungicidal and bactericidal properties, has been employed as a disinfectant on farms against storage rots and for the control and prevention of certain animal diseases, such as foot rot of sheep and cattle.

## **Benefits and Toxicity of Copper to Humans**

Copper is a key mineral in many different body systems. It is central to building strong tissue, maintaining blood volume and producing energy in your cells. Yet, for all its critical importance, you don't have much copper in your body—barely more than the amount found in a single penny. And those pennies in your pocket are only 2.5% copper by weight.

In the foods we commonly eat, there are only very small amounts of copper. As much as any dietary mineral, the amount of copper you eat is directly related to the amounts of minimally processed plant foods you get every day.

### **Role in Health Support**

#### **1. Antioxidant Protection**

Copper is one of the co-factors for one form of an enzyme called superoxide dismutase (SOD). SOD is one of the major antioxidant enzymes in the body. As a measure of how important SOD is, amyotrophic lateral sclerosis—also known as Lou Gehrig's disease—is thought to be the result of an under functioning (SOD) enzyme.

From recent studies where young volunteers were fed a copper-depleted diet, reduced SOD function was an early result. In fact, these changes were apparent within the first month of the experimental diet.

In more advanced cases of copper deficiency, including people who have undergone gastric bypass surgery, this loss of antioxidant protection over a period of years can lead to irreversible damage to the nervous system. However, this does not appear to occur without the types of unusual deficiency risks detailed below.

#### **2. Bone and Tissue Integrity**

Copper is required to manufacture collagen, a major structural protein in the body. When copper deficiency becomes severe, tissue integrity - particularly bones and blood vessels can begin to break down.

Luckily, it appears at the present time that a very severe and prolonged dietary deficiency of copper is necessary to lead to overt problems. For example, premature babies with immature gastrointestinal tracts can develop bone problems related to copper deficiency.

At least one recent author has speculated that the marginal copper status of the diets of about one-quarter of adults in the U.S. is related to eventual development of osteoporosis in some members of this group. For adults with borderline copper intake from food, deficient intake of nutrients like calcium and vitamin D is still likely to put them at greater risk than borderline intake of copper. Still, this low copper intake may be increasing their risk of osteoporosis and is very likely to be the subject of future research.

### **3. Energy Support**

Copper plays two key roles in energy production. First, it helps with incorporation of iron into red blood cells, preventing anaemia. Second, it is involved with generation of energy from carbohydrates inside of cells.

Each of these uses of copper also requires iron, and for this reason, the symptoms of copper deficiency can mimic those of low iron intake. Lentils and sesame seeds are just a few examples of World's Healthiest Foods rich in both iron and copper.

### **4. Cholesterol Balance**

Animal studies have demonstrated that copper-deficient diets lead to increases in blood cholesterol levels. In humans, this appears to be true in some situations, but not all. This should not be a surprise, as human diets are much more varied than those of laboratory animals. Interestingly, the effect of copper deficiency appears to be through increased activity of an enzyme called *HMG-CoA reductase* - the same enzyme targeted by the most commonly prescribed cholesterol medications.

### **Summary of Food Sources**

With the single exception of shrimp, all of the very good or excellent sources of copper among the World's Healthiest Foods are plant foods. These best copper sources are varied and come from many different food groups.

Our top three sources of copper are sesame seeds, cashews, and soybeans. Any of these three foods will bring at least three-quarters of your daily copper requirement. Shiitake and crimini mushrooms are also excellent copper sources and will provide 40 to 75% of your daily need.

Many of the excellent food sources of copper are leafy greens, including turnip greens, spinach, Swiss chard, kale, and mustard greens. Asparagus and summer squash are two other excellent vegetable sources of copper.

The good and very good sources of copper include many legumes, whole grains, nuts, and seeds. For example, flax seeds, walnuts, and garbanzo beans are rated as very good sources of copper.

Combining a grain- or legume-based recipe with an excellent vegetable source of copper could very easily provide the entire daily requirement of this mineral. For example, 7-Minute Sautéed Crimini Mushrooms would meet or exceed your daily Recommended Dietary Allowance (RDA) for copper.

### **Risk of Dietary Deficiency**

Between one-quarter to one-half of Americans fail to reach Dietary Reference Intake (DRI) recommendations for copper on a daily basis. In fact, deficient diets in the experimental research where scientists intentionally created copper-composition of those diets were quite similar to the average U.S. diet. These copper-depleted diets were based largely around meats, refined grains and dairy foods.

About 5% of U.S. adults eat a diet with less copper than was used in these studies. In fact, this 5% of U.S. adults obtain less copper from their diets on a daily basis than would be found in a single serving of navy beans - a food not even close to the best source of copper in our rating system.

According to a statistical analysis published in 2011, copper deficiency risk has risen substantially over the past 75 years. This is probably most related to modern food processing methods, although copper depletion of soils may also contribute to some extent.

### **Other Circumstances that Might Contribute to Deficiency**

Most of the non-dietary factors that contribute to copper deficiency tend to involve somewhat uncommon medical conditions. Gastric by-pass surgery and stomach surgeries are two examples. Certain cancers - like pancreatic cancer can increase risk of copper deficiency, as can celiac disease when it is poorly managed or untreated.

### **Relationship with Other Nutrients**

Prolonged supplementation with doses of zinc that go beyond normal dietary intake ranges can interfere with copper absorption and utilization, leading to copper deficiency.

### **Risk of Dietary Toxicity**

*Most U.S. adults struggle to achieve the Dietary Reference Intake (DRI) for copper intake, so the risk of dietary toxicity from copper is really only seen in a person with one of two issues.*

The first issue would be a genetic condition that impairs the ability to clear copper from the body, leading to a build-up to toxic levels. The most likely reason for this is a condition called Wilson's disease, an inherited genetic mutation. Wilson's disease is both rare (as few as one case per 100,000 people) and very severe. People with this condition and other similar genetic mutations that affect copper metabolism are usually diagnosed by the time they reach adulthood.

A more common reason to see risk of copper toxicity is due to excessive exposure from the water supply. This is not generally caused by excessive amounts in city water supplies—these are monitored by the Environmental Protection Agency (EPA), but by leaching from old copper pipes and fittings.

The amount of copper that is leached into water from old pipes can be significant, but it varies widely. If you have concern about the amount of copper in your tap water, you can take some simple steps to help reduce the exposure risk. First, the amount of leaching is directly related to the amount of time the water spends in the copper pipe. Use the first gallon or so of water in the morning for non-cooking tasks (for example, cleaning or watering plants). In fact, anytime you are getting drinking water from your tap, you can let the water run until you feel it get noticeably colder. Second, hot water will leach more copper than cold water, so if you want hot water for a beverage, you can use cold water and then heat it up rather than getting hot water out of your tap. Finally, you could install a water filter to remove much of the copper. Both activated charcoal and reverse osmosis filters should remove significant amounts of copper from your water. However, before taking any of these steps, make sure that toxicity risk is a greater risk for you than deficiency risk! You don't want to be lowering the amount of copper in your drinking water if you actually need more copper than you are getting from your food.

### Exercise 4.1.3

- Which of the following metals is less dense than water?
  - Aluminium
  - Magnesium
  - Sodium
  - Lead
- From the following list, choose a metal(s) that is appropriate for each use given below. Give reasons for your choice.

Aluminium	Lead	Copper	Duralumin
-----------	------	--------	-----------

  - Making a cooking pot
  - Making electrical wires
  - Building an aeroplane
  - Weights for a fishing net
- Out of the following four elements, which is the most reactive?
  - Zinc
  - Gold
  - Sodium
  - Magnesium
- Write balanced chemical equations for the reactions of the following metals with dilute HCl, oxygen and water. *If a metal does not react with either dilute HCl, oxygen or water, write 'no reaction'*
  - Aluminium
  - Copper
  - Magnesium
  - Zinc
  - Calcium
- What is the use of the following substances in the Blast Furnace?
  - Coke
  - Limestone
- State the composition of the following alloys:
  - Brass
  - Bronze
  - Solder
  - Stainless steel
- Identify a use of each of the alloys in (6) above and explain how it is better than the constituent element(s).
- The chemical process by which iron is produced from iron ore is called
  - reduction.
  - combustion.
  - melting.
  - oxidation.
- Inside a blast furnace, carbon monoxide reduces iron oxide to iron metal by
  - removing impurities.
  - melting the iron oxide.
  - adding oxygen to iron oxide.
  - removing oxygen from iron oxide.

## Non- Metals

Hydrogen, noble gases and Groups IV – VII elements are non-metals. These elements are generally dull, brittle and non-conductors of electricity. Some are solids, others are gases and bromine is a liquid at room temperature.

Some non-metals may form allotropes. This means that they may exist in more than one physical form.

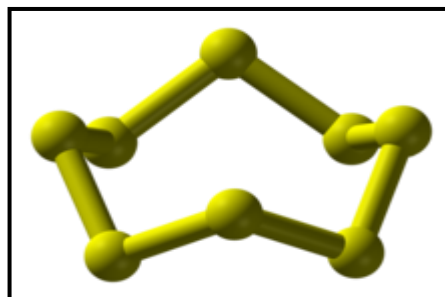
The table below summarises the physical properties of common non-metals.

Element	State	m.pt (°C)	b.pt (°C)	Density (gcm <sup>-3</sup> )	Conductivity
<b>Hydrogen (H<sub>2</sub>)</b>	Colourless Gas	- 259.2	- 252.9	0.0899	Non conductor
<b>Oxygen (O<sub>2</sub>)</b>	Colourless Gas	-218.4	-183.0	1.43	Non conductor
<b>Nitrogen (N<sub>2</sub>)</b>	Colourless Gas	-209.9	-195.8	1.25	Non conductor
<b>Chlorine (Cl<sub>2</sub>)</b>	Yellow Gas	-101.0	-34.6	3.21	Non conductor
<b>Neon (Ne)</b>	Gas	-248.6	-246.1	0.901	Non conductor
<b>Bromine (Br<sub>2</sub>)</b>	Red Liquid	-7.2	58.8	3.12	Non conductor
<b>Phosphorus (P)</b>	White Solid	44.1	280.0	1.82	Non conductor
<b>Sulphur (S)</b>	Yellow Solid	112.8	444.6	2.07	Non conductor
<b>Carbon (C)</b>	Black Solid	3500.0	4827.0	2.62	Conductor

## Sulphur

The 16<sup>th</sup> element in the Periodic Table is sulphur. It is a yellow solid ring structure made up of eight sulphur atoms.

Sulphur is found in nature as pure sulphur. Sulphides and sulphates also exist. It is also found as an impurity in natural gas or petroleum. There are references of the elements in ancient texts such as the Bible that referred to it as brimstone. It is an indication of extensive use of the element and its chemistry. In the commercial world, sulphur is used primarily for the production of sulphuric acid, the main raw material in manufacture of fertilisers. It is also found in matches, fungicide and insecticide.

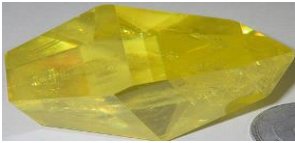




Source: en.wikipedia.org/wiki/

The figure below shows the main allotropes of sulphur: Rhombic, monoclinic and plastic sulphur.



Source: learnthings.co.za

<b>Rhombic Sulphur</b>	<b>Monoclinic Sulphur</b>	<b>Plastic Sulphur</b>
It is the most common form of sulphur. A yellow, odourless and brittle solid that burns readily in air with a blue flame.	When rhombic sulphur is slowly heated to a temperature above 94.5°C, it will change to needle like allotropes called monoclinic.	It is a non-crystal form of sulphur formed by pouring liquid sulphur into cold water. It has a texture similar to rubber.
		
Source: www.flickr.com	Source: site.google.com	Source:www.wonderwhizkids.com

## Chemical Properties of Sulphur

### (i) Reaction with water

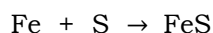
Sulphur is a non-polar solid so it is insoluble in water. Balls of plastic sulphur are readily formed by pouring liquid sulphur into cold water.

### (ii) Effect of heat

On slow heating, sulphur will melt to form a yellow viscous liquid. It changes from rhombic to monoclinic sulphur. If heating is continued, the yellow liquid will change to an orange red colour and becomes more viscous. At a temperature of 444°C, the orange liquid vaporises to a reddish brown vapour. The vapour burns with a blue flame forming a strong smelling gas, sulphur dioxide.

### (iii) Formation of iron sulphide

A mixture of powdered sulphur and iron filings will form a black iron sulphide when heated.



## Allotropes of Carbon

Carbon, a Group IV element, forms compounds that are the basis of living cells. The allotropes of carbon are graphite and diamond.



Diamond



Graphite

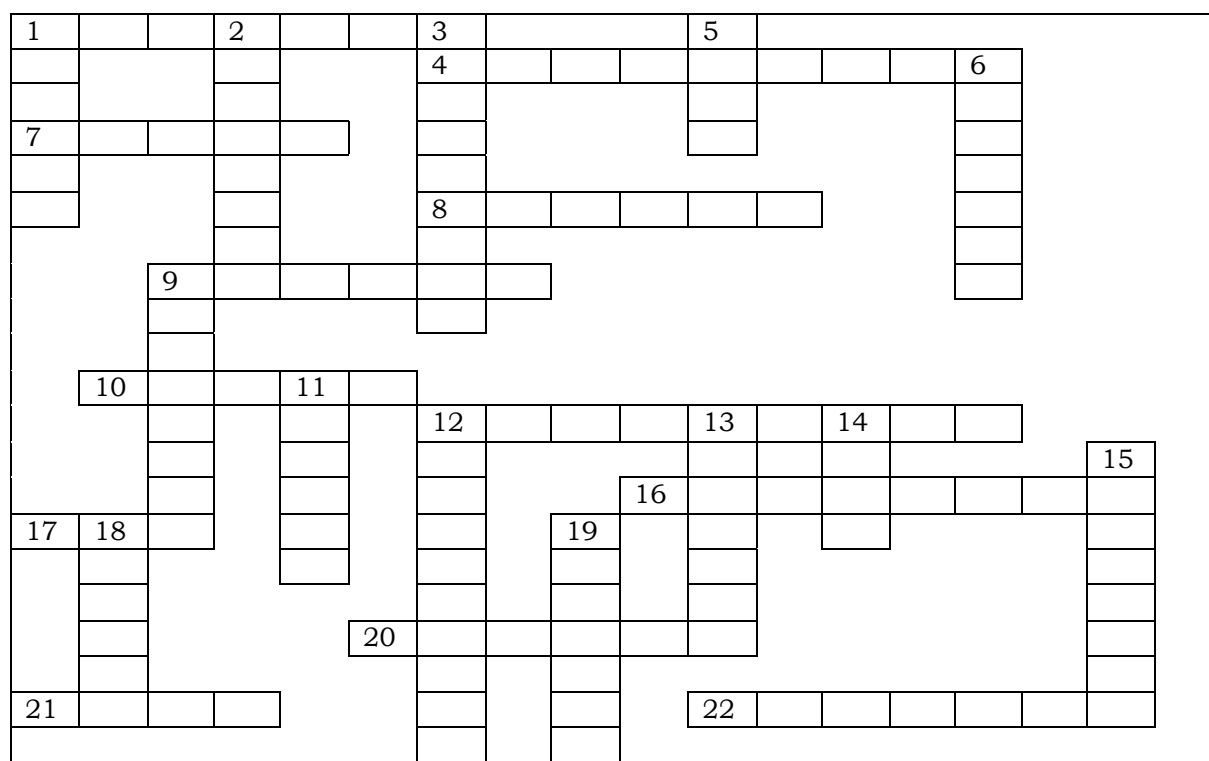
Source: shelf3d.com

### A comparison of Diamond and Graphite

	<b>Diamond</b>	<b>Graphite</b>
<b>Appearance</b>	Hard, transparent, colourless solid	Soft, black solid
<b>Melting / boiling point</b>	3550°C/4827°C	3720°C (sublimes)
<b>Use</b>	Jewellery and industrial cutting tools	Lubricant, pencil lead, electrolysis electrodes

### Exercise 4.1.4 Element Crossword

1. Each clue describes a naturally occurring element. Complete the cross word puzzle by writing the name of the element.



#### ACROSS

1. Makes strong bones and teeth
4. Bodies of vehicles, boats, and planes
7. Makes glass
8. 2<sup>nd</sup> place getter
9. Filling for weather balloon
10. Filling for tube light
12. Makes up NPK fertiliser
16. Added to water for dental health
17. A common metal for electroplating
20. Present in electrical appliances
21. Line sinker
22. Found in torch batteries

#### DOWN

1. Diamond is a form of it
2. Purify drinking water
3. Burns with bright flame
5. Galvanises iron
6. Quicksilver
9. Rocket fuel
11. Supports life and burning
12. Makes up NPK fertiliser
13. Yellow powder
14. Corrodes to form rusts
15. Fourth element of the Periodic Table
18. Abundant in sea food
19. Superman's element

2. What are allotropes?
3. Sulphur is the primary source in the production of sulphuric acid.
  - (i) Sulphur has many allotropes. Name any two allotropes of sulphur and describe their structures.
  - (ii) State one use of sulphur in agriculture.

## 4.2 Oxides of Metals and Non-Metals

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Work out and write the names and formula of oxides of metals and non-metals.
- ✓ Distinguish the formation of oxides of metals from the formation of non-metals oxides with balanced chemical equations.
- ✓ Describe the properties of metallic oxides and non-metallic oxides with balanced equations considering the type of bonding in them.
- ✓ State the sources and properties of the oxides found in the atmosphere.
- ✓ Conduct an experiment to prepare and test a sample of oxide gas.
- ✓ Describe the effect of carbon monoxide on human beings.

Metals are reactive so they react naturally with oxygen in the air to form their oxides. Metal oxides are the coating on the metal which tarnishes their appearance. Metal oxides are ionic compounds thus having ionic bond which holds the metal ion and the oxide ion. The ionic bond is the result of the transfer of electrons from the metal to oxygen.

### Example

1.  $4\text{Na}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{Na}_2\text{O}_{(s)}$
2.  $2\text{Mg}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{MgO}_{(s)}$

### Exercise 4.2.1

Complete the table given below for the formation of metal oxides.

Metal	Equation
<b>Aluminium</b>	
<b>Copper</b>	
<b>Zinc</b>	
<b>Iron</b>	
<b>Lead</b>	

Oxides of non-metals are covalent compounds formed when non-metals burn in oxygen. The covalent bond is formed from the non-metal and oxygen sharing electrons. Phosphorous react with oxygen on its own.

### Example

1.  $2\text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{H}_2\text{O}_{(g)}$
2.  $\text{P}_{4(s)} + 5\text{O}_{2(g)} \rightarrow \text{P}_4\text{O}_{10(s)}$

### Exercise 4.2.2

Complete the table below for the formation of non-metal oxides

Non-Metal	Equation
<b>Carbon</b>	
<b>Nitrogen</b>	
<b>Sulphur</b>	

## Metal Oxides

Name	Formula	State	Colour	Diagram	Solubility in Water
<b>Sodium oxide</b>	Na <sub>2</sub> O	Solid	White		Soluble
<b>Magnesium oxide</b>	MgO	Solid	White		Very slightly soluble
<b>Calcium oxide</b>	CaO	Solid	White		Soluble
<b>Aluminium oxide</b>	Al <sub>2</sub> O <sub>3</sub>	Solid	White		Insoluble
<b>Copper oxide</b>	CuO	Solid	Black		Insoluble
<b>Zinc oxide</b>	ZnO	Solid	White		Insoluble
<b>Iron oxide</b>	Fe <sub>2</sub> O <sub>3</sub>	Solid	Red Brown		Insoluble
<b>Lead oxide</b>	PbO	Solid	Yellow		Very slightly soluble

## Non- Metal Oxides

Name	Formula	State	Colour	Solubility in Water
<b>Carbon monoxide</b>	CO	Gas	Colourless	Soluble
<b>Carbon dioxide</b>	CO <sub>2</sub>	Gas	Colourless	Soluble
<b>Sulphur dioxide</b>	SO <sub>2</sub>	Gas	Colourless	Soluble
<b>Sulphur trioxide</b>	SO <sub>3</sub>	Gas	Colourless	Soluble
<b>Nitrogen dioxide</b>	NO <sub>2</sub>	Gas	Colourless	Soluble
<b>Phosphorous oxide</b>	P <sub>4</sub> O <sub>10</sub>	Solid	White	Soluble

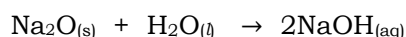
## Reactions of metal oxides

### 1. Reaction with water

Only highly reactive metal oxides react with water to produce metal hydroxides which are basic in nature. Metal oxides turn moist red litmus paper blue confirming that they are basic in nature.

The less reactive metal oxides do not react with water but they can react with dilute hydrochloric acid to form metal-chloride and water. There would be no change to the litmus paper since metal chlorides and water are neutral.

#### Example



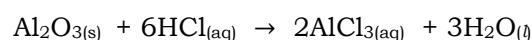
#### Exercise 4.2.3

Complete the reactions given below.

Metal Oxide	Reaction with water
MgO	
CaO	

### 2. Reaction with Dilute Hydrochloric acid

#### Example



#### Exercise 4.2.4

1. Complete the reactions given below.

Metal Oxide	Reaction with dilute hydrochloric acid
CuO	
ZnO	
Fe <sub>2</sub> O <sub>3</sub>	
PbO	

2. What type of bonding is found in:

- i. MgO
- ii. CO<sub>2</sub>






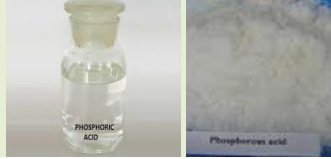
2. Name the following oxides.

- i. Al<sub>2</sub>O<sub>3</sub>
- ii. P<sub>4</sub>O<sub>10</sub>

Non-metals oxides are acidic in nature because they react with water to form acidic solutions.

### Exercise 4.2.5

Complete the table given below.

Non-metal oxide	Reaction with water	Acid produced	Acid Diagram
CO	$\text{CO}_{(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{HCOOH}_{(aq)}$	Methanoic acid	
CO <sub>2</sub>		Carbonic acid	
SO <sub>2</sub>		Sulphurous acid	
SO <sub>3</sub>		Sulphuric acid	
NO <sub>2</sub>		Nitric acid	
P <sub>4</sub> O <sub>10</sub>	$\text{P}_4\text{O}_{10(s)} + 6\text{H}_2\text{O}_{(l)} \rightarrow 4\text{H}_3\text{PO}_{4(aq)}$	Phosphoric acid	

### Sources of carbon dioxide

Carbon dioxide gas (CO<sub>2</sub>) is produced from:

1. Naturally by decomposition of dead plants and animals, respiration in all living organisms and volcanic eruption.
2. Combustion or burning of fossil fuels in cooking, vehicles and other machines, forest fire and any other fire, fermentation from microorganisms and brewery factories, and industrial processes like cement making.

### Effect of CO<sub>2</sub> gas on the Environment

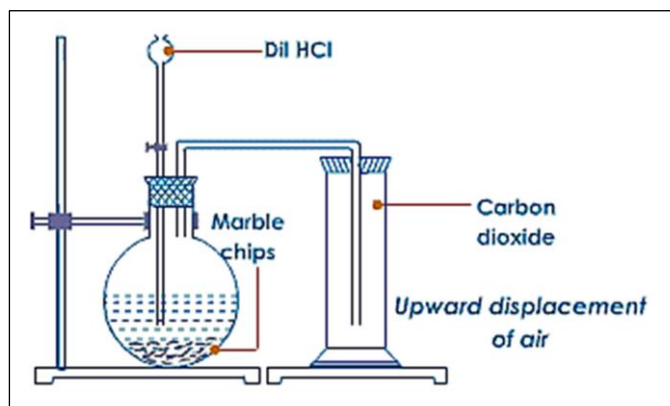
Carbon dioxide gas is used up by plants for photosynthesis. Too much of it is not good for the environment as it accumulates in the atmosphere and forms a layer. This shields the atmosphere and it reflects the heat back to earth. This increases the earth's temperature and causes the ice in the polar regions to melt. The sea level rise threatens the livelihood of coastal island people. This is called the Green House Effect.

## Preparation of carbon dioxide

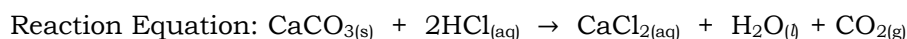
Carbon dioxide gas can be prepared by adding dilute hydrochloric acid to calcium carbonate or marble chips.

Note: Instead of calcium carbonate, other metal carbonates can also be used.

### Laboratory preparation of carbon dioxide gas



Source: <http://lucychemistry.blogspot.com>



## Sources of sulphur dioxide

Sulphur dioxide gas ( $\text{SO}_2$ ) is produced from:

1. Industrial activity that processes materials that contain sulphur. For example the generation of electricity from coal, oil or gas that contains sulphur.
2. Mineral ores also contain sulphur such as pyrite ( $\text{Fe}_2\text{S}$ ) and galena ( $\text{PbS}$ ) and sulphur dioxide is released when they are processed.
3. Industrial activities that burn fossil fuels containing sulphur can be important sources of sulphur dioxide.
4. Motor vehicle emissions contain sulphur dioxide as the result of fuel combustion.

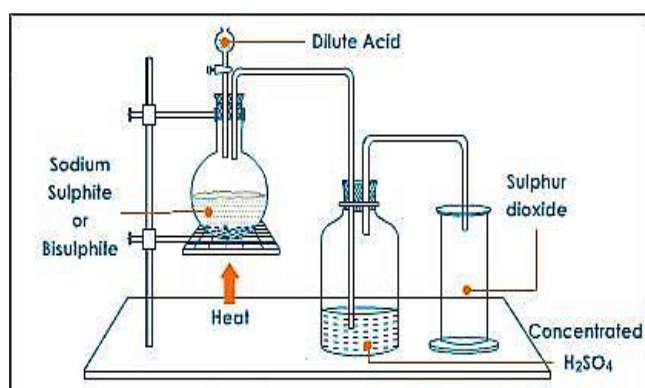
## Effect of sulphur dioxide gas on the environment

Sulphur dioxide gas reacts with water vapour in the atmosphere to form sulphurous acid. This acid falls with rain as acid rain and is harmful to the plant growth and animal survival.

## Preparation of Sulphur dioxide

Sulphur dioxide can be prepared by adding a dilute acid to sodium sulphite or Bisulphite.

### Laboratory preparation of sulphur dioxide



Source: <http://chemistryhomeworkhelp.blog.com>



## Comparison of Carbon dioxide and Sulphur dioxide

Properties	Carbon dioxide	Sulphur dioxide
Colour	Colourless	Colourless
State	Gas	Gas
Density	1.97 kgm <sup>-3</sup> Heavier than air ( $\rho = 1.29 \text{ kgm}^{-3}$ ) so can be collected by upward displacement of air.	1.43 kgm <sup>-3</sup> Heavier than air so can be collected by upward displacement of air.
Reaction with water	$\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$	$\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$
Reaction with sodium hydroxide	$2\text{NaOH} + \text{CO}_2 \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O}$	$2\text{NaOH} + \text{SO}_2 \rightarrow \text{Na}_2\text{SO}_3 + \text{H}_2\text{O}$
Effects on wet litmus paper	Turns damp blue litmus paper red showing acidic property.	Turns damp blue litmus paper red showing acidic property.
Test	Bubble $\text{CO}_2$ gas in limewater, it turns milky.	Freshly made orange potassium dichromate solution turns green.
Uses	In carbonated beverages. Fire Extinguishers	Sulphur dioxide is a reducing agent and is used for bleaching and as a fumigant and food preservative.

### Carbon monoxide (CO)

Carbon monoxide is produced by the incomplete combustion of fuel. It is harmful to the body as it causes blood poisoning. This happens when the blood carries carbon monoxide gas instead of the oxygen needed by the body cells for the energy releasing process (respiration).

#### Poisonous nature of CO

CO binds very strongly to the iron in haemoglobin (red blood pigment that carries oxygen). Once carbon monoxide attaches, it is very difficult to be released. If one breathes in carbon monoxide, it sticks to the haemoglobin and takes up all of the oxygen binding sites. As a result, the blood loses all of its ability to transport oxygen, hence a person suffocates. Prolonged exposure to fresh air (or pure oxygen) is required for the CO-tainted haemoglobin (carboxy-haemoglobin) to clear.

Since CO binds to haemoglobin so strongly, a person can be poisoned by carbon monoxide even at very low concentrations if exposed for a long period of time. Carbon monoxide is produced by cars, gas appliances, wood stoves and cigarettes.

#### Exercise 4.2.6

- Write an equation to show how these gases can be produced in the laboratory using the reagents given. Also give one way in which you can test for the presence of these gases.
  - Carbon dioxide - from  $\text{CaCO}_3$  and  $\text{HCl}$
  - Sulphur dioxide - from  $\text{Na}_2\text{SO}_3$  and  $\text{HCl}$
- Describe the effect of carbon monoxide on human beings.
- Draw a concept map on the sub-strand 'Chemistry of Air' and present it to the class.

## 4.3 Organic Substances

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Examine and outline the sources and formation of oil, natural gas and biogas.
- ✓ Investigate the impact of using fossil fuels on the environment and present measures of reducing these impacts.
- ✓ List and describe the different classes of hydrocarbons.
- ✓ Identify and discuss the properties of hydrocarbons.
- ✓ Explore and list some useful products that are made from hydrocarbons.

All living things, plants and animals alike, contain carbon compounds which are vital to their life. In addition, natural fuels such as coal and petroleum which are known in enormous quantities and they are called *organic compounds*.

### Oil – Source and Formation

The most important liquid fuels are gasoline (petrol), kerosene (paraffin), diesel oil and fuel oil. All of these are mixtures of hydrocarbons obtained from crude oil.

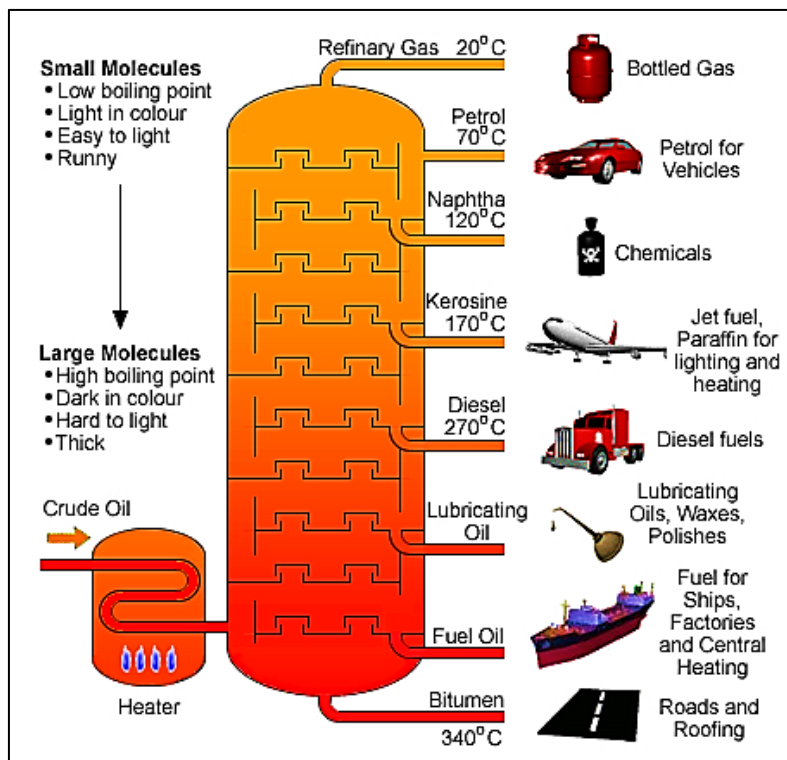
### Crude oil

Crude oil is a fossil fuel found in rocks. It has been formed at the bottom of oceans over millions of years. Scientists believe that the remains of sea plants and animals become trapped in the sediment on the sea bed where they decomposed without air to form oil.

### Refining crude oil

The fractional distillation of alkanes is performed in a refinery. Fuel and many other equally important products are obtained from the crude oil. Alkanes can be separated because large alkanes have high boiling points and small alkanes have low boiling points. The apparatus that is used to separate alkanes is called a *fractionating column*.

The figure below shows a fractionating column.



Source: <http://www.science-resources.co.uk/>

## Natural gas and Biogas

The most important gaseous fuel is natural gas. It is formed when plants decay in the absence of air. Natural gas is an inexpensive fuel because it requires very little treatment before use. Biogas is made by bacteria from plant and animal waste. It is used for cooking or for powering automobiles.

Advantages of fossil fuels	Disadvantages of fossil fuels
Fossil fuel is less expensive. Alternative energy sources such as solar, wind and water are more expensive but cleaner and replaceable.	Fossil fuels are non-renewable and burning those causes air pollution.

## Hydrocarbons

Organic chemicals that contain only carbon and hydrogen are called hydrocarbons. There are three classes of hydrocarbons: alkanes, alkenes and alkynes.



## Comparison amongst the three hydrocarbons

	Alkanes	Alkenes	Alkynes
<b>Description</b>	Alkanes are hydrocarbons that contain only single bonds between the carbon atoms.	Alkenes are hydrocarbons that have at least one double bond between two carbon atoms.	Alkynes are hydrocarbons that have at least one triple bond between two carbon atoms.
<b>General Formula</b>	$C_nH_{2n+2}$	$C_nH_{2n}$	$C_nH_{2n-2}$
<b>Naming (follows a system)</b>	Starts with the number of carbon atom and finishes with -ane.  <b>Example</b> C <sub>1</sub> -methane, C <sub>2</sub> -ethane, C <sub>3</sub> -propane, C <sub>4</sub> -butane, C <sub>5</sub> -pentane, C <sub>6</sub> -hexane	Starts with the number of carbon atom and finishes with -ene.  <b>Example</b> C <sub>2</sub> -ethene, C <sub>3</sub> -propene, C <sub>4</sub> -butene, C <sub>5</sub> -pentene, C <sub>6</sub> -hexene	Starts with the number of carbon atom and finishes with -yne.  <b>Example</b> C <sub>2</sub> -ethyne, C <sub>3</sub> -propyne, C <sub>4</sub> -butyne, C <sub>5</sub> -pentyne, C <sub>6</sub> -hexyne
<b>Properties</b>	<ol style="list-style-type: none"> <li>Colourless, insoluble in water and neutral. They will dissolve in oils and waxes.</li> <li>Undergo combustion reactions producing carbon dioxide, water and heat.</li> <li>Boiling and melting points increase as the number of carbon atom increases.</li> </ol>	Similar to those of alkanes but are more reactive.	Similar to those of alkanes but are more reactive
<b>Uses</b>	<ol style="list-style-type: none"> <li>Fuels e.g. methane, ethane</li> <li>Starting materials for making other chemicals, including alcohols, plastics and fibres.</li> </ol>	Manufacture of plastics, alcohols, artificial ripening of fruits, polymers for PVC and Teflon.	Used in welding gas, rocket fuel and medicine.

## Name and Formula of Alkanes

Name	No. of carbon atoms	Molecular Formula	Structural Formula
<b>Methane</b>	1	CH <sub>4</sub>	<pre>       H         H — C — H               H           </pre>
<b>Ethane</b>	2	C <sub>2</sub> H <sub>6</sub>	<pre>       H   H             H — C — C — H                   H   H           </pre>
<b>Propane</b>	3	C <sub>3</sub> H <sub>8</sub>	<pre>       H   H   H                 H — C — C — C — H                       H   H   H           </pre>

**Exercise: Complete the table below.**

Name	No. of carbon atoms	Molecular Formula	Structural Formula
<b>Butane</b>			
		$C_5H_{12}$	
			<pre>       H   H   H   H   H   H                         H — C — C — C — C — C — C — H                               H   H   H   H   H   H           </pre>

### Names and Formula of Alkenes

Name	No. of carbon atoms	Molecular Formula	Structural Formula
<b>Ethene</b>	2	$C_2H_4$	<pre>       H   H             H — C = C — H           </pre>
<b>Propene</b>	3	$C_3H_6$	<pre>       H   H   H                 H — C — C = C — H               H           </pre>

**Exercise: Complete the table below.**

Name	No. of carbon atoms	Molecular Formula	Structural Formula
<b>Butene</b>		$C_4H_8$	
	5		<pre>       H   H   H   H   H                         H — C — C — C — C = C — H                       H   H   H           </pre>
<b>Hexene</b>	6		

### Name and Formula of Alkynes

Name	No. of carbon atoms	Molecular Formula	Structural Formula
<b>Ethyne</b>	2	$C_2H_2$	$H-C \equiv C-H$
<b>Propyne</b>	3	$C_3H_4$	<pre>       H         H — C ≡ C — C — H                               H           </pre>

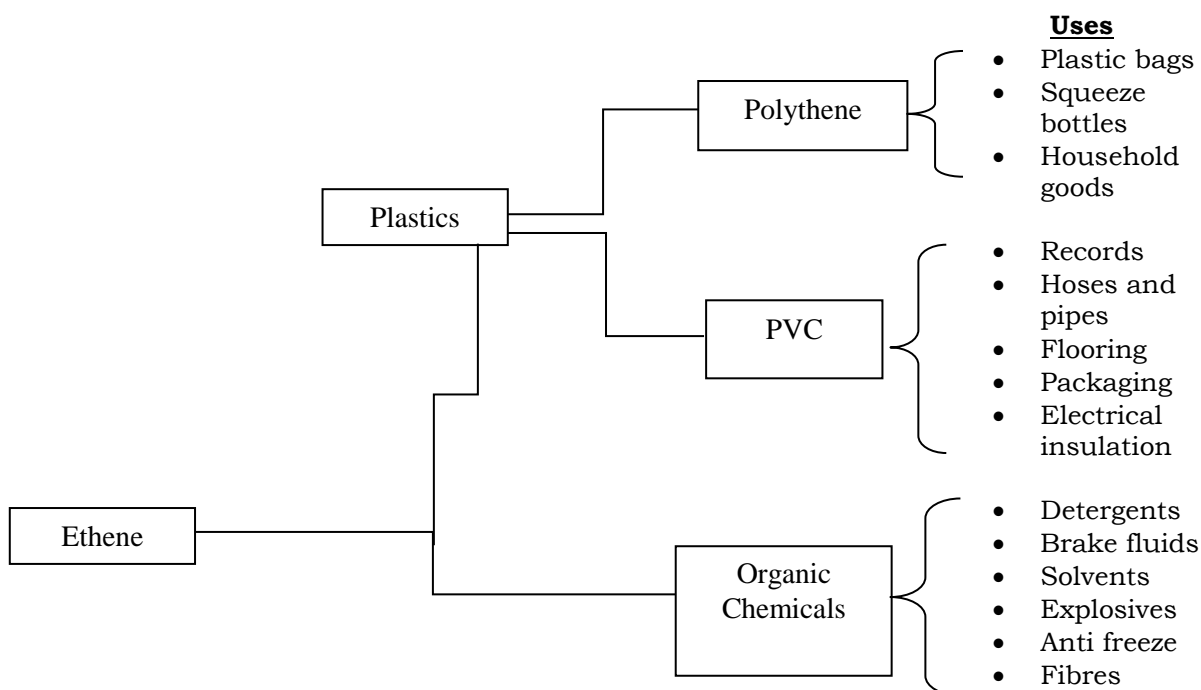
**Exercise: Complete the table below.**

Name	No. of carbon atoms	Molecular Formula	Structural Formula
	4	$C_4H_6$	
	5		<pre>       H   H   H                 H — C ≡ C — C — C — C — H                                       H   H   H           </pre>
<b>Hexyne</b>	6		

**Note:**

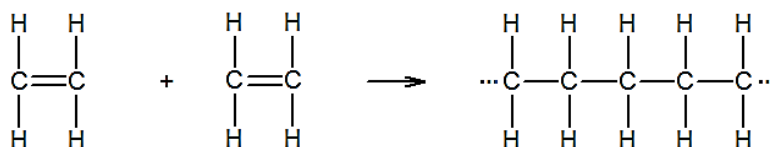
Compounds that contain only carbon-carbon single bonds are called saturated compounds. An example is alkanes.

Compounds which contain carbon-carbon double or triple bonds are called unsaturated compounds. Examples include alkenes and alkynes.

**Uses of Ethene****Polythene**

Polyethylene (polythene) is a polymer of ethene. It is a soft, chemically resistant thermoplastic. This means that it softens on heating, allowing it to be moulded and extruded into thin sheets.

Polythene is made by addition polymerisation, using the ethene molecules.

**Note:**

- A polymer is a very large molecule made from joining together thousands of small molecules.
- A monomer is a small molecule that is added together to make a polymer.
- Polymerisation is the process of making a polymer.

**Teflon**

Teflon is a compound of fluorine and carbon called poly tetrafluoroethylene (PTFE). It is a plastic material that is highly resistant against chemicals, corrosion and oxidation.

It is used in the manufacture of chemical processing equipment and for coating cooking utensils.

## Terylene

Terylene is a synthetic fibre made by a process of polymerising ethylene glycol and terephthalic acid. It is very extensively used in textile industries to make hard wear clothes like sarees and dress material.

## Polypropylene

Polypropylene, also known as polypropene is a thermoplastic polymer made from the monomer, propylene by addition polymerisation. It is used to make ropes, thermal underwear, carpets and stationaries.

## Rubber

Synthetic rubber is obtained by polymerisation of unsaturated monomers. Synthetic rubbers are marketed as compressed bales and square blocks. It is used in the tyre industry, medical equipment and molded parts. Sulphur is the oldest vulcanisation agent for unsaturated rubber.

### Note:

- Condensation polymerisation – when monomers join together and a water molecule is lost.
- Addition polymerisation – when monomers are added to each other to form long chain molecules (polymer) without by-products.

### Exercise 4.3.1

1. Write the general formula for:
  - i. Alkanes
  - ii. Alkenes
  - iii. Alkynes
2. Define the following terms:
  - i. Hydrocarbon
  - ii. Saturated hydrocarbons
  - iii. Unsaturated hydrocarbons
  - iv. Natural gas
  - v. Bio gas
  - vi. Monomer
  - vii. Polymer
  - viii. Polymerisation
3. What happens to the melting and boiling point of alkanes (*does it increase, decrease or stays the same*) as the size of alkanes increases.
4. Explain why alkanes can be separated using fractional distillation.
5. Write an equation which shows the formation of polythene from ethene.
6. Draw a concept map on the sub-strand 'Organic Substances' and present it to the class.

# STRAND 5 ENVIRONMENTAL CHEMISTRY

**Strand Outcome:** Recognise the chemical substances in the environment, both natural and introduced and their effects on living material.

## Sub-strands

- 5.1 Air and Air Pollution
- 5.2 Water and Water Pollution
- 5.3 Fertilisers and Soil Pollution
- 5.4 Climate Change

## 5.1 Air and Air Pollution

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Investigate and show how oxygen, carbon dioxide and water vapour are detected.
- ✓ Describe the extraction, preparation and uses of some gases.
- ✓ Prepare oxygen gas and carry out its confirmatory test, identify its properties and uses.
- ✓ Identify some air pollutants and explain how each pollutant affects the environment.
- ✓ Outline and present means and ways of reducing these effects on the environment.

### Air as a Mixture

Air is a mixture of gases. Most of it – 78% is nitrogen.

The table below shows the percentage composition of gases in the air.

Gas	Percentage (%)
Nitrogen	78
Oxygen	21
Argon	1
Carbon dioxide	0.04
Neon, Krypton, Xenon, Water Vapour	Very small amounts

### Detection of Gases

#### 1. Oxygen

Oxygen is detected by lighting a wooden splint, blowing it out and then placing it into the gas you think might be oxygen. If oxygen is present, the splint will burst into flames again.

## 2. Carbon dioxide

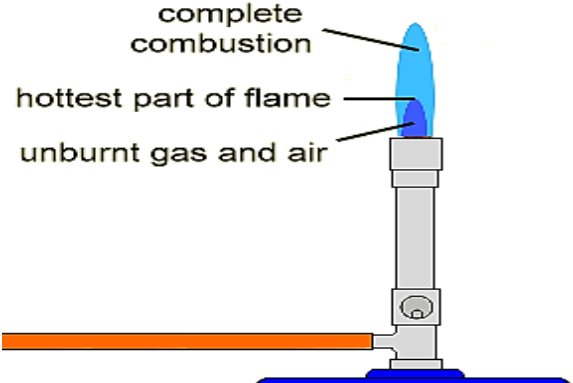
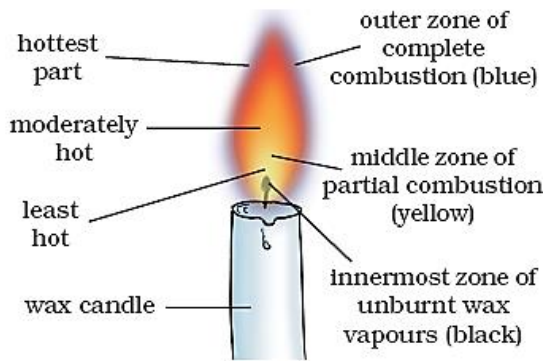
You can test a gas to see if it is carbon dioxide by using limewater. Limewater is a solution of calcium hydroxide in water. When carbon dioxide is mixed with lime water, calcium carbonate (chalk) is formed. This is not soluble in water, so it makes the limewater look cloudy (milky).

## 3. Water Vapour

A cold trap is used to detect water vapour in the air. A cold trap is a U-tube that is immersed in a container of ice and salt. The salt makes the ice melt at a temperature lower than 0°C. When air passes through the U-tube it freezes the water vapour in the air and this can be seen.

### Bunsen and Candle Flame

The table below shows the differences between bunsen and candle flame.

Bunsen Flame	Candle Flame
<ul style="list-style-type: none"><li>Burns with complete combustion.</li><li>Produces heat but not much light.</li><li>Blue flame.</li></ul>	<ul style="list-style-type: none"><li>Burns with incomplete combustion.</li><li>Produces light but not much heat.</li><li>Yellow flame.</li></ul>
 <p>Source: <a href="http://www.bbc.co.uk/">http://www.bbc.co.uk/</a></p>	 <p>Source: <a href="http://www.studyrankers.com">http://www.studyrankers.com</a></p>

## Air in Industry

Vast quantities of air are separated to provide us with pure nitrogen, oxygen, argon etc. The air is first liquefied and then fractionally distilled.

### Note:

- Liquefaction** – Is when air is changed to a liquid by compressing it and making it very cold.
- Fractional distillation** – Separates liquids on the basis of their boiling points.

The table below shows the boiling points and uses of gases

Name	Formula	Boiling point (°C)	Uses
Oxygen	O <sub>2</sub>	-183	Hospitals, welding, steel industry
Nitrogen	N <sub>2</sub>	-196	Packaging food, refrigerants
Helium	He	-269	Inflating balloons
Neon	Ne	-246	Filling coloured tube lights
Argon	Ar	-186	In candescent light globes, welding, radio valves, electrical rectifiers

## Air Pollution

In densely populated areas and particularly in large cities, often air contains many harmful substances. It is polluted. Air pollution may be due to smoke and/or poisonous gases.

The table below shows the effect of air pollutants

Pollutant	Symbol	Source	Effects
<b>Sulphur dioxide</b>	SO <sub>2</sub>	Incineration of garbage, motor vehicles, power plants, metal smelters, petroleum refineries.	Irritant to the eye and respiratory tract, causes acid rain, retards the formation of chlorophyll in plants.
<b>Nitrogen oxides</b>	NO, NO <sub>2</sub> , N <sub>2</sub> O	Motor vehicles, planes	Produces smog, harmful to the respiratory system.
<b>Carbon dioxide</b>	CO <sub>2</sub>	Motor vehicles, industries, deforestation	Contributes to enhanced greenhouse effect
<b>Carbon monoxide</b>	CO	Motor vehicles, wood fires	Causes dizziness and weakness, stops blood from carrying oxygen
<b>Chlorofluoro – carbons</b>	CFCs	Aerosols, air conditioners, and refrigeration systems, fast – food cartons	Destroys ozone layer
<b>Ozone</b>	O <sub>3</sub>	Secondary pollutant from automobiles	Harms lungs, destroys trees
<b>Dust</b>		Industries, motor vehicles	Irritates eyes and lungs

### Note:

- **Primary pollutant** – Is a pollutant that is directly released into the air e.g. from motor vehicles.
- **Secondary pollutant** – Is a pollutant that forms in the air from chemical reaction between primary pollutants.

## Oxygen

21% of air is made up of oxygen. Oxygen is slightly soluble in water hence able to sustain aquatic life. It also helps things to burn (combustion). Oxygen is a colourless and odourless gas.

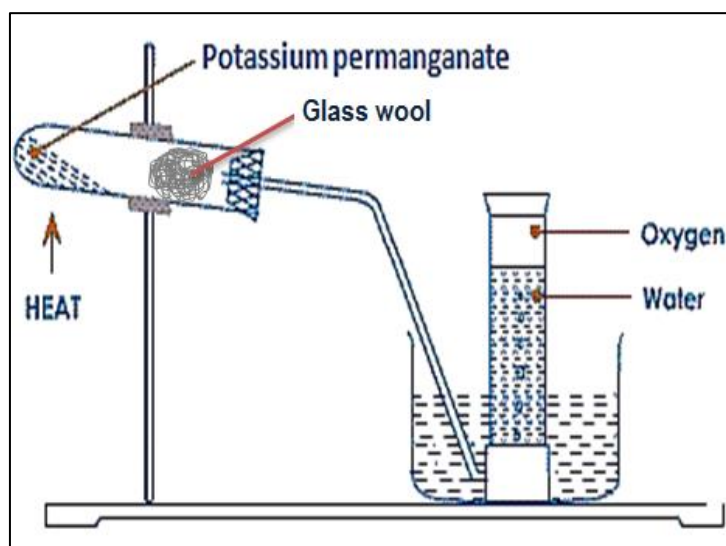
### Test for oxygen

Oxygen re-lights a glowing splint.

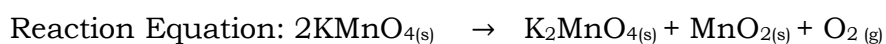
### Laboratory preparation of oxygen

Oxygen can be prepared in the laboratory by heating potassium permanganate.

#### Experimental set-up for the laboratory preparation of oxygen



Source: <http://chemistry.tutorcircle.com>

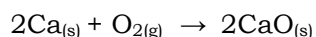


#### Safety Measures

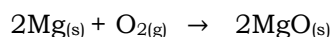
1. Glass wool prevents hot  $\text{KMnO}_4$  from splattering into the tubing.
2. A section of rubber tubing between the test-tube and water is pinched to prevent water from “sucking back” into the hot test-tube.

### Combustion of selected elements

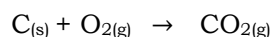
1. **Calcium** – burns with a bright flame. A white powder, calcium oxide is produced.



2. **Magnesium** – burns with a bright white flame. A grey powder, magnesium oxide is produced.



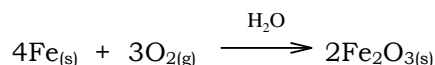
3. **Carbon** – glow red and produces a colourless gas, carbon dioxide.



4. **Wood and Paper** – burns with a yellow flame. A greyish white powder, ash is produced.

## Rusting

The red-brown deposits on steel and iron is called rust. Rust is iron oxide ( $\text{Fe}_2\text{O}_3$ ). Oxidation of iron causes rusting; a process where iron combines with oxygen to form rust.



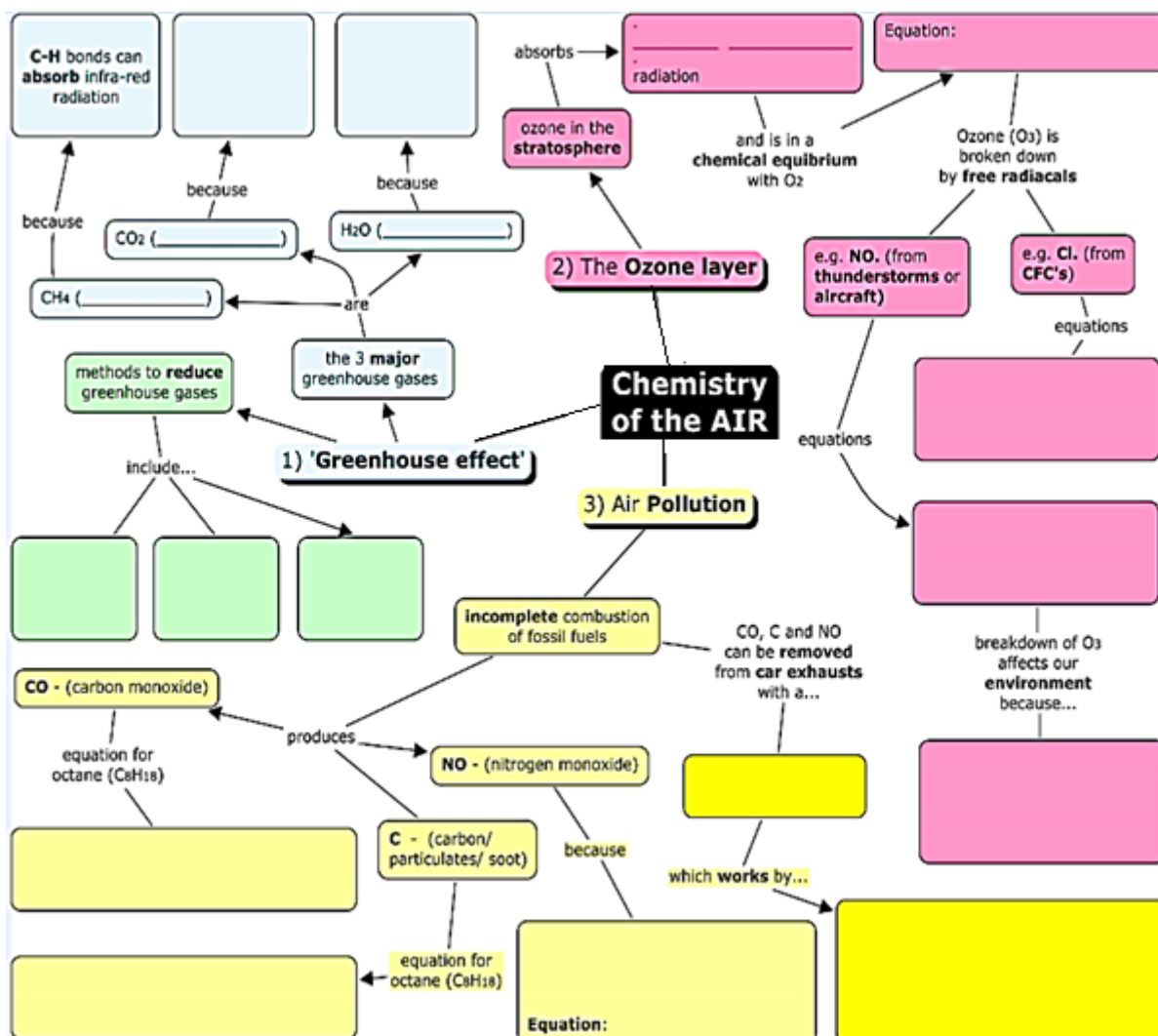
Rust forms when iron is exposed to both water and oxygen and can be prevented by painting, oiling and electroplating.

## Ozone Layer

The layer of ozone ( $\text{O}_3$ ) in the stratosphere is very useful. It protects the earth from the harmful ultra-violet light of the sun. Ozone layer is damaged by the substances known as CFC's which are used in refrigerators, air conditioners, foam packaging, spray cans as well as cleaning solvents. Discharge of nitrogen oxide ( $\text{NO}$ ) from planes also destroys the ozone layer. Depletion of ozone layer can lead to cell damage causing mutations and skin cancer and slow growth of crops.

### Exercise 5.1.1

Complete the concept map given below



## 5.2 Water and Water Pollution

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Describe and show some properties of water.
- ✓ Collect information and present on types of water pollution in Fiji, their sources, effects and ways to minimize these effects.
- ✓ Describe the stages of water treatment processes and its importance to human health.
- ✓ Collect information and talk about some mineral water extraction in Fiji.
- ✓ Study and describe stages of sewage treatment as applied locally.
- ✓ Explore and discuss the effect of discharging raw sewage on the environment.

### Importance of water

Water is essential for life and for maintaining our lifestyle, which depends extensively on industries. Water makes up a very large proportion of all living things. It is essential for many of the chemical reactions in living organisms. Water is also essential for many of our industrial processes. It is used to dissolve reactants, to clean equipment and to cool products of reactions.

### Density of Water and Ice

Water expands when frozen; therefore, the density of ice is less than that of water. For this reason, ice floats on top of water. Water molecules in ice arrange themselves in a regular hexagonal pattern. This arrangement occupies more space than liquid water. What this means is that water remains underneath the ice, enabling aquatic life to survive.

### Surface Tension of Water

Water molecules are attracted to one another. They are attracted equally from all directions. The forces balance out. This force keeps the surface together. The surface acts like a 'skin', with the molecules sticking together. This 'skin' is under tension and the effect is known as surface tension.

### Solubility and Water

The solubility of a substance is measured by the amount of that substance that dissolves at a particular temperature. The solubility of a gas in water depends on the pressure of the gas. If a gas is put under pressure, more dissolves. Gases also become less soluble with increasing temperature.

Oxygen is also present in all natural waters and is very important for the plants and animals which live in them. Green water plants perform photosynthesis and keep oxygen and carbon dioxide balance just as they do on land. However, when water is polluted e.g. with raw sewage, bacteria in the water may feed on the pollutant and multiply very rapidly. These bacteria need oxygen and use it up faster than it can be replaced. The balance is thus upset and animal life can no longer survive in the water.

### Water Pollution

Water can be polluted by toxic chemicals which mainly comes from pesticides and poisonous chemical or by heavy metals, such as mercury and lead. Organic materials from household garbage and human waste can also pollute water. Pollutants such as heavy metals accumulate inside fish and people can be affected by eating fish which are contaminated. Toxic chemicals from land fill sites can also be washed into the waters which can contaminate surrounding soil and water and harm fish and other aquatic creature.

Organic pollutants which have high biochemical oxygen demand (BOD) use up oxygen in the water and kill aquatic life.

## Hard and Soft Water

Water which will not form lather with soap is called hard water. The opposite is true for soft water. Hard water contains dissolved salts which react with soap and detergent and prevent it from forming lather. Usually the soap causes an insoluble scum to form in hard water.

Dissolved minerals that cause hardness in water are mainly calcium ions ( $\text{Ca}^{2+}$ ), magnesium ions ( $\text{Mg}^{2+}$ ) and very occasionally iron ions ( $\text{Fe}^{3+}$ ).

## Types of Hardness

There are two types of hardness of water; temporary hardness and permanent hardness. Temporary hardness is caused by calcium hydrogen carbonate and can be removed by boiling.



Permanent hardness is caused by calcium sulphate and cannot be removed by boiling. Washing soda and ion exchange is used to remove permanent hardness as well as temporary hardness. Water softening is removal of calcium ions from hard water.

## Differences between Soap and Synthetic Detergent

Soap is made from natural products like vegetable oil while synthetic detergent is made from synthetic material mainly petroleum products. Soap forms scum when used in hard water whereas synthetic detergents do not form scum in hard water. Soap is biodegradable while some detergents are not biodegradable.

## Solvent cleaners and bleaches

Cleaning solvent is used to remove oil, grease and other contaminants. They are usually liquids and their purpose includes removing offensive odour and prevention of contamination. Some solvent cleaners can kill bacteria and clean at the same time.

Bleaches are used as household chemicals to whiten clothes and remove stains and as disinfectants primarily in the bathroom and kitchen.

## Components of soap

Soap is made up of vegetable oil and caustic soda. A soap molecule has two ends: one end that is soluble in grease and the other end that is soluble in water. Soap removes oil and grease by helping it to dissolve in water.

## Stain removal

Solvent cleaners remove stains by dissolving the stains in the solvent such as water or alcohol. Detergents decrease the surface tension of water, essentially making it 'wetter' so that it is less likely to stick to itself and more likely to interact with oil and grease.

It is important to know which types of detergent are good for washing keeping in mind their effect on skin, clothes and our environment.

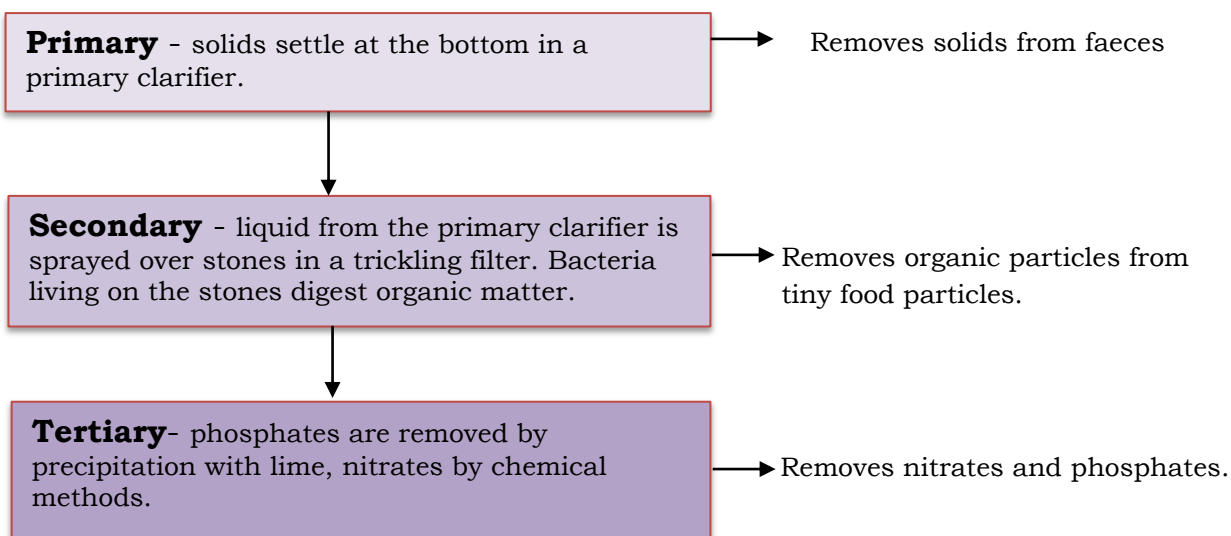
## Sodium hypochlorite

Sodium hypochlorite is the most common bleaching agent. It is used in many households to whiten laundry, disinfect hard surfaces in kitchens and bathrooms and keep swimming pools free of infectious agents.

## Sewage Treatment

A sewage treatment plant like the one in Kinoya separates sewage into effluent and sludge.

### Stages of sewage treatment



#### Note:

**Effluent** – is the liquid that is released from the sewage treatment plant.

**Sludge** – is the solid from the primary and secondary clarifier.

## Water Treatment

Drinking water comes from rainwater, surface water and ground water. Water needs to be treated before it can be used for drinking.

There are three steps in the treatment of drinking water: Coagulation, filtration and chlorination.

Coagulation	Filtration	Chlorination
Aluminium sulphate ( $\text{Al}_2(\text{SO}_4)_3$ ) helps coagulate dirt particles, copper sulphate ( $\text{CuSO}_4$ ) prevents algal growth, sodium carbonate neutralises acidity and softens water.	Layers of sand and gravel remove floc.	Chlorine kills bacteria, fluoride reduces tooth decay and lime neutralizes acidity

### Effect of Fluoride on Health

Fluoride is added to reduce tooth decay. However, excessive consumption of fluoride (above the required amount) may lead to increased likelihood of bone fractures in adults and may result in effects on bone leading to pain and tenderness.

### Lead Poisoning

Lead is a heavy metal found naturally in the environment as well as in manufactured products. The major sources of lead emissions are motor vehicles, lead smelters, waste incinerators and lead-acid battery manufactures. People, animals and fish are mainly exposed to lead by breathing and ingesting it in food, water, soil, or dust. Lead accumulates in the blood, bones, muscles and fats. It causes damage to the kidneys, liver, brain and nerves and other organs. Excessive exposure to lead causes seizures, mental retardation, and behavioural disorder and memory problems.

### Exercise 5.2.1

1. Briefly explain why the density of ice is less than that of water.
2. Briefly explain what you understand by the term 'surface tension' of water.
3. There are two types of hardness of water, temporary and permanent hardness.
  - i. Which type of hardness can be removed by boiling? Given a suitable reaction equation to support your answer.
  - ii. List way(s) to remove permanent hardness in water.
4. What is the difference between hard water and soft water?
5. Briefly explain why soap does not lather in hard water?
6. List some effects of lead poisoning.
7. List some practices which lead to water pollution.
8. What does the primary sewage treatment remove?
  - A. Effluent
  - B. Organic particles
  - C. Phosphates and nitrates
  - D. Solids from faeces
9. In sewage treatment, what does *effluent* refer to?
  - A. The liquid that comes into the plant.
  - B. The liquid that is released from the plant.
  - C. The solid particles that comes into the plant.
  - D. The solids that settle at the bottom of the clarifiers.
10. The steps in treating drinking water is
  - A. chlorination, filtration, coagulation.
  - B. chlorination, coagulation, filtration.
  - C. filtration, chlorination, coagulation.
  - D. coagulation, filtration, chlorination.
11. How are nitrates and phosphates removed during sewage treatment?
12. Why are aluminium sulphate and chlorine added during water treatment?
13. What material(s) is used for filtering the water during water treatment?
14. Draw a concept map on the sub-strand 'Water and Water Pollution' and present it to the class.

## 5.3 Fertilisers and Soil Pollution

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Identify and list some properties and uses of nitrogenous fertilisers.
- ✓ Discuss the impact of using fertilisers on the environment and suggest ways of minimising this effect.
- ✓ Compare the nitrogen content in various nitrogenous fertilisers.

### Fertilisers

Fertilisers are added to the soil to replace the nutrients that are missing. Fertilisers can either be organic or chemical. The most common chemical fertiliser added to soils is nitrogen fertiliser. Organic fertilisers include methods such as mulching, making compost and adding vegetable peelings.

### Common nitrogen fertilisers

Fertiliser	Description / Name
<b>NPK</b>	Percentage of nitrogen, phosphorus and potassium vary according to the need in the soil.
<b>(NH<sub>4</sub>)<sub>2</sub>SO<sub>4</sub></b>	Ammonium sulphate
<b>(NH<sub>4</sub>)<sub>3</sub>PO<sub>4</sub></b>	Ammonium phosphate
<b>NH<sub>4</sub>NO<sub>3</sub></b>	Ammonium nitrate
<b>(NH<sub>2</sub>)<sub>2</sub>CO</b>	Urea
<b>NH<sub>3</sub></b>	Ammonia

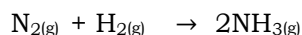
In the fertilisers above, nitrogen is present in all of them. This means that nitrogen is an important nutrient and must be present in large amounts. In fact, high nitrogen content of fertilisers will mean the healthy growth of a plant.

### Examples of fertilisers used locally.



## Manufacturing Fertilisers

An important ingredient in the making of nitrogen fertiliser is ammonia. Ammonia is made by a process called the Haber Process. This process involves taking nitrogen from the air and allowing it to react with hydrogen obtained from natural gas.



To make the various fertilisers, ammonia is then reacted with other chemicals to produce the common nitrogen fertilisers listed in the table below.

Chemical added to Ammonia	Fertiliser produced	Chemical Equation
Sulphuric acid	Ammonium sulphate	$2\text{NH}_{3(\text{g})} + \text{H}_2\text{SO}_{4(\text{aq})} \rightarrow (\text{NH}_4)_2\text{SO}_{4(\text{aq})}$
Nitric acid	Ammonium nitrate	$\text{NH}_{3(\text{g})} + \text{HNO}_{3(\text{aq})} \rightarrow \text{NH}_4\text{NO}_{3(\text{aq})}$
Phosphoric acid	Ammonium phosphate	$3\text{NH}_{3(\text{g})} + \text{H}_3\text{PO}_{4(\text{aq})} \rightarrow (\text{NH}_4)_3\text{PO}_{4(\text{aq})}$
Carbon dioxide	Urea	$\text{NH}_{3(\text{g})} + \text{CO}_{2(\text{g})} \rightarrow 2\text{CO}(\text{NH}_2)_{2(\text{aq})}$

In the table above, it is evident that ammonia is basic in character and that it will react easily with acids. At the end, the product formed is a salt. Therefore, these nitrogen fertilisers are salts and have characteristics of salts in where they are highly soluble in water.

## Properties of Fertilisers

When manufacturing fertilisers, it is important that they have the following properties which are typical of all fertilisers:

1. Must be highly soluble in water because plants can only absorb nutrients that are dissolved in water.
2. Must have high nitrogen content because the higher the nitrogen content, the less mass of fertiliser that is needed, therefore the better the growth of the plant.
3. Must be cheap so that it is affordable to all farmers.
4. The raw materials to make them are locally available and that they contain nutrients that the locality needs.

## Plant Growth

In order for plants to grow well and crop yield is of the desired state, all nutrients must be present. If some nutrients are not present, then the following changes can be noticed on a plant:

- (i) White or yellow patches on leaves.
- (ii) Plant may appear very small.
- (iii) Plant may not be fruiting despite leaves and rest of plant growing well.
- (iv) Leaves may appear curled inwards and not open.
- (v) Slow growth; stunted plants

### Exercise 5.3.1

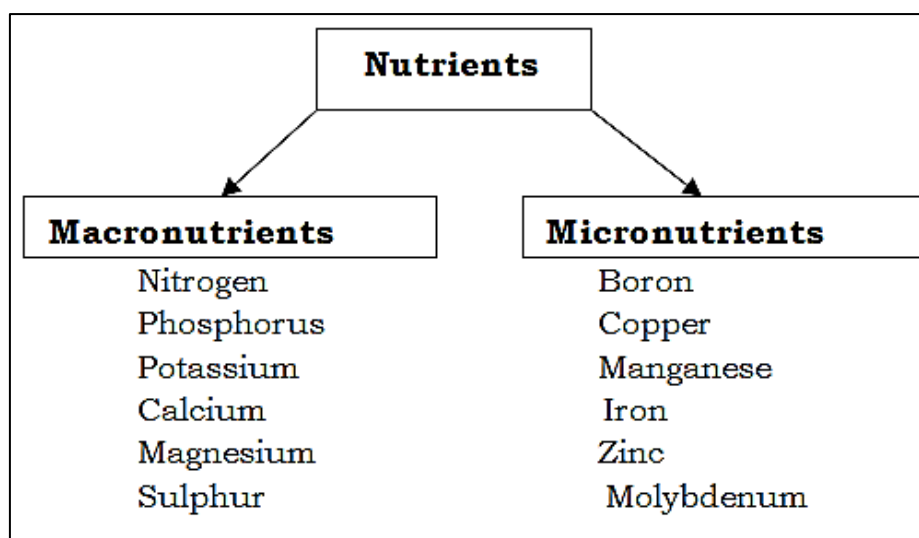
What conclusions can you make from the illustration given below?

The illustration shows a bag of 5-10-5 fertilizer. Callouts explain the benefits of each component:

- 5% NITROGEN** MEANS BETTER LEAF GROWTH: GRASS, LEAFY VEGETABLES, LEAFY ORNAMENTALS.
- 10% PHOSPHORIC ACID** MEANS MORE FRUITS AND FLOWERS: TOMATOES, STRAWBERRIES, ROSES.
- 5% POTASH** BUILDS STRONG ROOTS: POTATOES, BEETS, CARROTS.

All plants need carbon dioxide, water and minerals from the soil in order to grow. The minerals that plants need are called *nutrients* and are classified as:

- (i) Macronutrients – Those nutrients needed in large amounts by plants.
- (ii) Micronutrients - Those nutrients needed in small amounts by plants.



#### Exercise 5.3.2

To help you remember the macronutrients and micronutrients make up two sentences – one for macronutrients and one for micronutrients, consisting of the symbols or first letters of the names of the elements considered as nutrients.

A good supply of all the nutrients above is necessary for plants to grow well. If a plant is well supplied with potassium and iron and growing in enough sunlight but lacks nitrogen, it will still not grow well. Due to this reason, nitrogen is known as the Limiting Factor. A limiting factor refers to something (can be nutrient or water or sunlight) that a plant is lacking that is slowing its growth. Therefore, when the soil lacks a certain nutrient, chemical fertilisers can be used. Fertilisers help replenish the soil with nutrients that it is lacking.

#### Effect of Nutrients on Plant Growth:

Nutrient	Symbol of Nutrient	Effect
<b>Carbon</b>	C	Carbon, hydrogen and oxygen are needed by plants to manufacture carbohydrates.
<b>Hydrogen</b>	H	
<b>Oxygen</b>	O	
<b>Nitrogen</b>	N	Needed for good growth of leaves and is good for building up proteins.
<b>Phosphorus</b>	P	- Important for healthy root development. - Aids in nitrogen fixation in leguminous plants. - Essential for cell division.
<b>Potassium</b>	K	- Required for healthy skin and fruit formation. - Aids in the formation of proteins and carbohydrates - Regulates water in the plants.
<b>Calcium</b>	Ca	Essential for the development of growth tissues e.g. root tips.
<b>Magnesium</b>	Mg	Forms a necessary part of the chlorophyll.
<b>Sulphur</b>	S	Is part of many proteins and some oils.

#### Note:

Plants absorb water from the soil and along with this, nutrients are taken in. Animals eat these plants as a source of protein and their excretions containing urea releases nitrates back into the soil. Leaching of nutrients from farms can lead to eutrophication. Eutrophication is the growth of microscopic plants in water ways, therefore depriving marine organisms their space to live.

Fertilisers are highly soluble salts and readily leach with high rainfall. Different fertilisers differ in their impact on soil pH. When fertilisers dissolve in water, they break up to form ions which can therefore increase or decrease soil acidity or alkalinity.

Ions/Nutrients from fertilisers	Effects on Soil pH
Ammonium ions ( $\text{NH}_4^+$ )	Causes a decrease in soil pH over time, therefore making the soil acidic.
Nitrate ions ( $\text{NO}_3^-$ )	Causes an increase in soil pH, therefore making soil more alkaline.
Ca, Mg, K and Na	May slightly increase or cause no change in soil pH
Sulphur, ammonium sulphate and compounds such as iron and aluminium sulphates.	Can reduce soil pH, therefore making soil more acidic.

**Note:**

Negative ions like  $\text{NO}_3^-$  are basic since they can be referred to as proton acceptors while positive ions like  $\text{NH}_4^+$  are likely to be acidic since they can donate a proton,  $\text{H}^+$ .

Fertilisers are important in replenishing the soil with the necessary nutrients and their addition to the soil is crucial for healthy growth of plants. However, their additions will cause the soil pH to vary.

If soil becomes too acidic, then soil acidity can be corrected and brought back to the desired pH by adding lime. Lime is calcium carbonate and is basic in character due to the presence of carbonate ions when lime dissolves in water. The carbonate ion,  $\text{CO}_3^{2-}$  may act as a proton acceptor and therefore acts as a base.

However, if pH is too high, then elemental sulphur, concentrated sulphuric acid or aluminium sulphate can be added to the soil to reduce alkalinity.

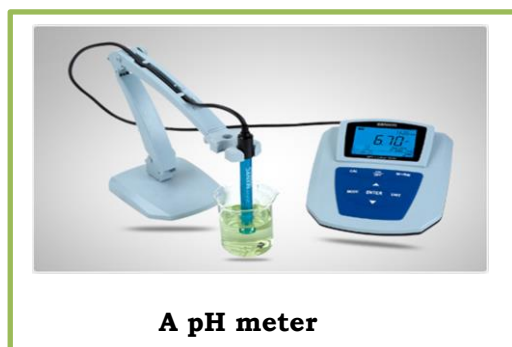
Soil pH affects the relative abundance of micro-organisms in the soil. This is why pH of soil has to be in the most desired value. Bacteria prevail more in alkaline soils while fungi dominate in acidic soils. These are important because they are responsible in the cycling of nutrients in the soil. Adjusting soil pH may also help farmers manage plant diseases.

### Measuring Soil pH

Measuring the pH of soils is important in determining various properties of soil.

Two ways of measuring the pH of soil are

1. **A pH meter or pH conductor** - is the most accurate method of measuring the pH of a soil sample.



Source: <http://www.shsan-xin.com/>

2. **Universal Indicator** - Universal indicators can be in either paper or liquid form. It is easy to use but less accurate than the use pH meter.



Source: [www.rtg.wa.edu.au](http://www.rtg.wa.edu.au)

### **Effect of Solubility of Fertilisers on Acidic and Basic Soils**

Fertilisers can be added to the soil through various methods. However, one property that it must have is that it has to be highly soluble in water. Nutrients and minerals present in the fertilisers are more soluble in acidic soils than in alkaline or neutral soils.

When rainfall is low, solubility of fertilisers is low. Therefore, availability of nutrients and minerals to plants is also low. Dry soils have been found to have a pH above neutral, indicating alkaline condition. When this happens, plants can be seen to be in a state of malnutrition because they are not able to maintain a balanced amount of nutrients from the soil.

In soils that are strongly acidic, the availability of phosphorus, potassium, sulphur, calcium and magnesium to the plants is reduced. This allows the concentration of soluble aluminium and manganese to increase in the soil, hence causing harm to plants.

A pH range of 6 to 7 is ideal for the ready availability of plant nutrients.

### **Uses of Phosphate and Nitrogen Fertilisers & Its Environmental Impact**

Phosphorus from phosphate fertilisers and nitrogen from ammonium fertilisers are two essential nutrients needed for growth and survival of both plants and animals. These two nutrients are added to the soil to increase crop yield to meet the demand of the increasing human population.

During heavy rainfall, nitrogen and phosphorus in the form of nitrates and phosphates leach into groundwater and occur as land runoff into streams, rivers, estuaries and other water sources. Their significant accumulation over time may therefore lead to the growth of plants and algae. This is known as **eutrophication**. When the intensity of algae growth over the water increases, light available to plants deeply rooted in the water is reduced. When these plants die and decompose, they used up the oxygen in the water. Hence, due to reduced oxygen content of the water, fish is not only greatly affected but there occurs a loss of important habitat to fish and other marine organisms.

The algal growth over the water intensifying over time because of the nutrient enriched aquatic system. This may occur as algal blooms.



*The figures above show the excessive growth of water weeds in some creeks in Nasinu and in the Rewa River.*

## Percentage Nitrogen of Common Nitrogen Fertilisers

Nitrogen is the only nutrient needed in the largest amount in fertilisers. The higher the nitrogen percentage in fertilisers, the faster plants will grow. If plants do not have enough nitrogen, they will grow slowly and have yellow leaves.

Name of fertiliser	Formula of fertiliser	Percentage of nitrogen	Advantage	Disadvantage
Ammonium sulphate	$(\text{NH}_4)_2\text{SO}_4$	21%	Easy to use; also supplies sulphur	Low nitrogen content; most acidifying fertiliser
Ammonium nitrate	$\text{NH}_4\text{NO}_3$	35%	High nitrogen content	Absorbs water from the air; not suitable for soils that undergo heavy leaching and denitrification
Urea	$\text{CO}(\text{NH}_2)_2$	47%	High nitrogen percentage	More expensive; denitrification on wet soils can be serious; leaching can be a problem in coarse soils
Ammonia	$\text{NH}_3$	82%	Inexpensive	Injected into the soil by special equipment; hazardous to handle
NPK	NPK	Variable	Combines nitrogen, phosphorus and potassium in a single fertiliser	Must purchase proper ratio.

### Calculating Percentage Composition of Nitrogen in Fertilisers

It is important to know the percentage of nitrogen in all fertilisers because the presence of nitrogen in all fertilisers is essential for growth of plants. To calculate the percentage nitrogen, the following formula can be used:

$$\% \text{ N} = \frac{\text{Atomic mass (N)} \times \text{number of nitrogen atoms in the compound}}{\text{Molecular mass of compound}} \times \frac{100}{1}$$

#### Example

Determine the percentage of nitrogen in Urea

#### Solution

$$M [\text{CO} (\text{NH}_2)_2] = 60$$

$$\% \text{ N} = \frac{14 \times 2}{60} \times \frac{100}{1} = 47\%$$

#### Exercise 5.3.3

Calculate the percentage composition of nitrogen in the following fertilisers.

- (i)  $(\text{NH}_4)_2\text{SO}_4$     (ii)  $\text{NH}_4\text{NO}_3$     (iii)  $\text{NH}_3$     (iv) NPK

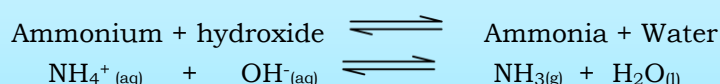
## Methods of Quantitative & Qualitative Analysis of Nitrogen in Common Nitrogen Fertilisers

### 1. Qualitative Analysis

This analysis involves the separation of desired constituents from a given sample. A series of chemical tests is conducted to identify the cations or anions present in the solution of the fertiliser. Nitrogen is present in fertilisers or in soils in the form of nitrates and ammonium ions. Therefore, a simple test for ions can easily be carried out to determine their presence in solutions of nitrogen fertilisers.

#### Test for Nitrates

Add sodium hydroxide solution and aluminium powder to the solution containing the nitrate ions. The aluminium reduces the nitrate ion ( $\text{NO}_3^-$ ) to an ammonium ion ( $\text{NH}_4^+$ ). The ammonium ion reacts with hydroxide to produce ammonia gas and water.

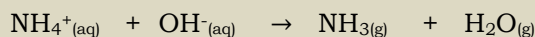


#### **Identifying the ammonium ion, $\text{NH}_4^+$**

**Test:** Add aqueous sodium hydroxide to the solid or solution under test and warm the mixture.

**Result:** If ammonium ions are present then a pungent-smelling gas is produced. The gas produced turns damp red litmus paper blue. It is ammonia,  $\text{NH}_3$ .

**Equation:**



### 2. Quantitative Analysis

This analysis involves determining the amount or percentage of any constituent in any compound; in this case, nitrogen. Several methods can be used:

- Gravimetric analysis** – Involves the determination of a component of the fertiliser by changing one compound containing that desired component into another compound and measuring the amount in percentage.

#### **Example**

- The gravimetric analysis of phosphorus is based on the precipitation of phosphorus as  $\text{MgNH}_4\text{PO}_4 \cdot 6\text{H}_2\text{O}$  from a solution that contains monohydrogen phosphate ions ( $\text{HPO}_4^{2-}$ ) ammonium ions ( $\text{NH}_4^+$ ) and magnesium ions ( $\text{Mg}^{2+}$ ). The slow neutralisation of these ions with aqueous ammonia produces the precipitate. This precipitate can be obtained dry through vacuum filtration and then weighed to help obtain its percentage yield.
- The amount of sulphate in a solution of barium sulphate can be determined quantitatively by gravimetric analysis also. This involves the slow addition of a dilute solution of barium chloride to a hot sample of the sulphate solution that is slightly acidified by adding a little hydrochloric acid. At the end, a white precipitate of barium sulphate will form and this can be filtered off and oven dried. The percentage of sulphate in the product- barium sulphate can then be worked.

- ii. **Volumetric or titrimetric analysis** – This involves measuring the volume of a reagent containing the desired component. Often involves a titration experiment.

In all quantitative analysis, the basic tool used is the analytical balance which is used for accurate weighing of samples and precipitates.

### Example

1. Weigh 1.3 g of fertiliser and dilute with water to 250 ml.
2. Prepare 20 ml aliquots and place this into a conical flask.
3. Place 20 ml of 0.100 M sodium hydroxide solution into a separate flask.
4. Add about 50 ml of distilled water and boil until all ammonia is gone.
5. Let the flask cool down then fill the burette with 0.100 M HCl.
6. Add methyl red indicator and titrate with the HCl until endpoint is reached.
7. Use the results from above to calculate the % of nitrogen in the fertiliser sample.

### Exercise 5.3.4

1. List some uses of nitrogenous fertilisers.
2. List three properties of nitrogenous fertilisers which make them suitable for agricultural use.
3. List some impacts of using fertilisers on the environment. Suggest ways of minimising these effects.
4. What is Eutrophication?
5. Which of the following fertilisers has the highest nitrogen content?
  - A. Urea
  - B. Ammonia
  - C. Ammonium nitrate
  - D. Ammonium sulphate
6. The desirable soil pH range for optimum plant growth is between
  - A. 1-3.
  - B. 3-5.
  - C. 6-7.
  - D. 11-12.
7. State the following concerning the two type of analysis - "Quantitative" and "Qualitative"
  - (a) Major difference
  - (b) Similarity
8. Draw a Concept Map on the Sub-Strand 'Fertilisers and Soil Pollution'.

## 5.4 Climate Change

### Achievement Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Study and present on climate change, greenhouse effect and global warming
- ✓ Examine and explain the impacts of global warming on the local resources and human livelihood.
- ✓ Explore and present means and ways of combatting global warming.

It has been noticed that the words climate change and global warming are often used interchangeably. However, they have different definitions:

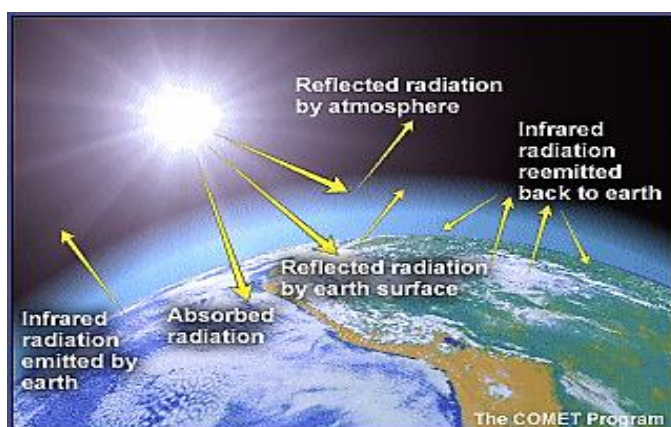
#### Global warming

- This is a more specific term.
- It refers to an increase in the earth's average temperature because of greenhouse gases build up in the atmosphere.
- Global warming is a cause of climate change.

#### Climate Change

- This is a broader term.
- It refers to changes (increase/decrease) to long-term weather patterns, such as temperature, rainfall or snowfall.
- A warmer Earth leads to changes in rainfall patterns, a rise in sea level, and a wide range of impacts on plants, wildlife, and humans.

#### A look at the earth's atmosphere



Source: [www.windows2universe.org](http://www.windows2universe.org)

**Note:** There is a 'blanket' around the earth. We live in a 'greenhouse'

A layer of greenhouse gases – mostly water vapor and much smaller amounts of carbon dioxide, methane and nitrous oxide act as a thermal blanket for the earth. About half the light reaching the Earth's atmosphere passes through the air and clouds to the surface of the earth. This light is absorbed and radiated upward in the form of infrared heat. About 90 percent of this heat is absorbed by the greenhouse gases and radiated back toward the surface of the earth. This is called the 'greenhouse effect'. It is a normal phenomenon. The surface of the earth is thus warmed to a life-supporting average temperature of 15°C.

### **The Enhanced Greenhouse Effect**

A lot of scientists agree that man's activities are making the natural greenhouse effect stronger. In other words, man is producing more greenhouse gases which keep increasing in the atmosphere. This is called the 'Enhanced Greenhouse Effect'.

The following are some causes of the increase in the greenhouse gases:

- ✓ Burning fossil fuels – e.g coal, oil and natural gas. These release carbon dioxide.
- ✓ Cutting down and burning trees also produce carbon dioxide.
- ✓ The use of chlorofluorocarbons (CFCs) in aerosols such as hairspray cans, fridges and in making plastics.
- ✓ Use of commercial and organic fertilisers on farms.
- ✓ Nitric acid production.

Due to the increase in the greenhouse gases in the atmosphere, more heat gets trapped and the earth gets warmer. This is known as Global Warming.

### **The role of carbon dioxide in global warming**

Most climate scientists agree that the main cause of the current global warming trend is the human expansion of the "greenhouse effect". This means the warming that results when the atmosphere traps heat radiating from Earth toward space.

Carbon dioxide is a minor but very important component of the atmosphere. It accounts for approximately 70 per cent of the enhanced greenhouse effect. The following are ways in which it is released into the atmosphere:

- Through natural processes such as respiration and volcano eruptions.
- Decomposition of wastes in landfills, agriculture and especially rice cultivation as well as ruminant digestion and manure management associated with domestic livestock.
- Through human activities such as deforestation, land use changes and burning fossil fuels.

Humans have increased the atmospheric CO<sub>2</sub> concentration by a third since the Industrial Revolution began. This is the most important long-lived "forcing" of climate change.

### **Role of carbon dioxide on climate and ocean composition change**

In the past 200 years, the oceans have absorbed approximately half of the carbon dioxide produced by fossil fuel burning and cement production.

- Increase in carbon dioxide concentration lowers the ocean pH and carbonate ion (CO<sub>3</sub><sup>2-</sup>) concentration. This makes the ocean acidic (ocean acidification).
- Ocean acidification is irreversible.
- Calcifying organisms will find it more difficult to produce and maintain their shells and hard structures.
- There is weakening of coral skeletons and reef structures.



Source: [www.teachoceanscience.net](http://www.teachoceanscience.net)

*Healthy coral polyps take  $\text{CaCO}_3$  from the ocean water to build their skeleton.*



Source: [www.teachoceanscience.net](http://www.teachoceanscience.net)

*Due to the increase in  $\text{CO}_2$  in the water, the coral's skeleton is damaged by ocean acidification.*

Reducing carbon dioxide emissions to the atmosphere appears to be the only practical way to minimize the risk of large-scale and long-term changes to the oceans.

### **Impacts of nitrogen fertilisers on global warming**

- Increased use of synthetic fertilisers leads to a rise in atmospheric nitrous oxide ( $\text{N}_2\text{O}$ ), which is a major greenhouse gas.
- For example, in the United States, agricultural soil management is the largest source of  $\text{N}_2\text{O}$  resulting in about 69% of the total  $\text{N}_2\text{O}$  emissions in 2011.
- Breakdown of nitrogen in livestock manure and urine also emit  $\text{N}_2\text{O}$ .

### **Effect of global warming on the water cycle**

Global warming is having an effect on the water cycle thus affecting:

- ✓ The amount of water available.
- ✓ The distribution of water.
- ✓ The quality of water.
- ✓ The timing of rainfall.

The water cycle is expected to undergo significant changes.

- ✓ More water would evaporate from the land and the ocean.
- ✓ There would be an increase in precipitation and runoff in some parts of the world – mostly in cold months in some parts of US, especially the Northeast and Midwest, thus leading to flooding.
- ✓ Other areas – notably the Southwest of US – can expect less precipitation, especially in the warm months, and longer, more severe droughts.
- ✓ There would be less snow and more rain in many parts of the cold countries.
- ✓ Wet areas are expected to get wetter and dry areas drier.
- ✓ Water temperature, for example, will generally rise in streams, lakes, and reservoirs as air temperature rises. This would lead to lower levels of dissolved oxygen in water, hence more stress on the fish, insects, crustaceans and other aquatic animals that rely on oxygen.
- ✓ Rise in sea level worldwide due to - the expansion of the ocean as it warms, and the increased melt from ice sheets, ice caps and glaciers.

Note: The effects of climate change on groundwater are not fully understood yet.

### **Effect of global warming on volcanic islands and atolls**

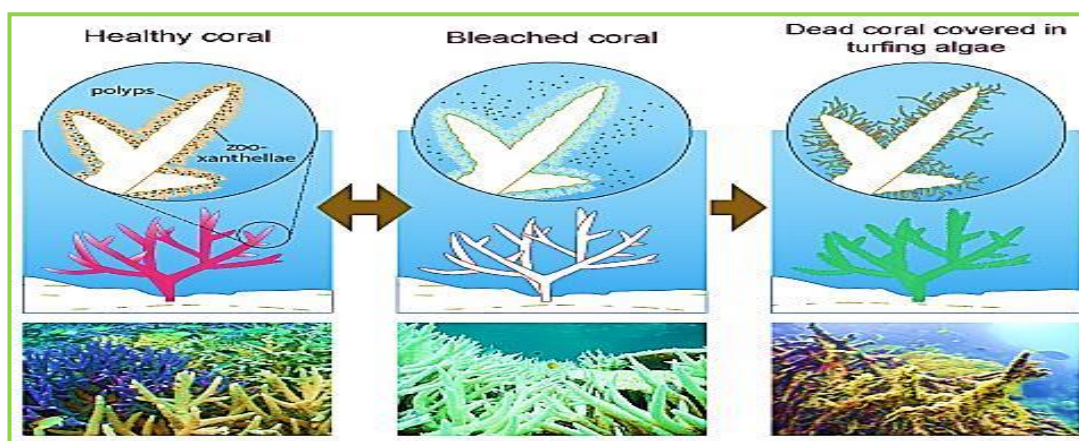
- ✓ Serious degradation of the coastal environment and natural resources.
- ✓ Higher rates of erosion and coastal land loss are expected in many Pacific Islands due to the increase in sea level. For example, for Majuro atoll in the Marshall Islands and

Kiribati, it is estimated that for a 1 m rise in sea level as much as 80% and 12.5%, respectively, of total land would be vulnerable.

- ✓ Pacific Islands are mainly vulnerable to coastal flooding.
- ✓ Influence on marine and terrestrial pathogens, such as coral diseases and oyster pathogens.
- ✓ Low-lying island states and atolls are likely to experience increased sea flooding, inundation (outpouring) and salinisation (salty) as a direct consequence of sea level rise.

### Effect of global warming on corals

- Warmer water temperatures result in coral bleaching.
- When water is too warm, corals expel the algae (zooxanthellae) living in their tissues (relationship between the coral and zooxanthellae, which give coral much of their colour, breaks down).
- Without the zooxanthellae, the tissue of the coral animal appears transparent and the coral's bright white skeleton is revealed. This is called coral bleaching.
- Other stressors can also cause bleaching, including freshwater inundation (low salinity) and poor water quality from sediment or pollutant run-off.



Source:<http://www.gbrmpa.gov.au>

### Effect of bleaching on the coral

- Corals begin to starve once they bleach. If the corals are stressed for long then the bleached corals die.
- Some corals are able to feed themselves and survive.
- If the water temperatures return to normal, corals can regain their zooxanthellae, hence return to their normal colour and survive.

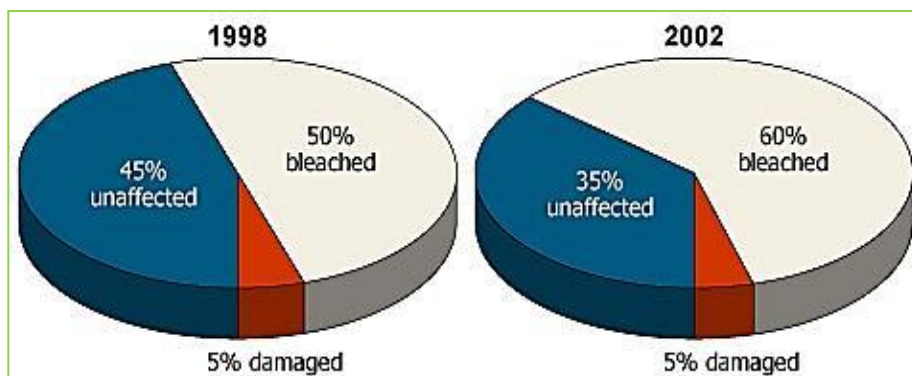
#### Note:

1. Coral reefs that have high rates of coral death following bleaching can take many years or decades to recover.
2. When a coral bleaches, it is not dead. Corals can survive a bleaching event but they are under more stress and are subject to mortality.
3. Cold water may also lead to coral bleaching. For e.g. In January 2010, cold water temperatures in the Florida Keys caused a coral bleaching event that resulted in some coral death.

## Coral bleaching on the Great Barrier Reef

The Great Barrier Reef has experienced mass coral bleaching events in the past.

The data given below shows the percentage of coral bleaching and coral damage in 1998 and 2002 due to an increase in the sea temperature.



Source: <http://www.gbrmpa.gov.au>

### Consequences of global warming

- Higher temperatures could lead to increased droughts and wildfires, heavier rainfall and a greater number of more powerful and dangerous hurricanes (Category 4 and 5 hurricanes).
- More flooding from storm surges, flash flooding and river flooding.
- Rising temperatures damage coral reefs.
- Melting of the habitats of polar bears and Antarctic penguins.
- It is predicted that the Arctic summers could be ice-free by 2040 due to the melting of glaciers and ice caps.
- Sea levels could rise as much as 23 inches by 2100 if current warming patterns continue. Global sea level has already risen by 4 to 8 inches in the past century.
- More deadly heat waves.
- Affect the types of crops grown in different parts of the world.
- Hotter weather also enables deadly mosquitoes (e.g dengue fever) to travel greater distances.
- Carbon dioxide in the air increases the risk of asthma and allergies.
- Increase in number of infectious disease like malaria and dengue fever.
- More food and waterborne illness outbreaks. During heavy rainfall pathogens from contaminated soils, farms, and streets are washed into drinking water sources.

**UNLESS WE ACT NOW, OUR CHILDREN WILL INHERIT A HOTTER WORLD, DIRTIER AIR AND WATER, MORE SEVERE FLOODS AND DROUGHTS AND MORE WILDFIRES.**

Shown below are photos taken at Daku Village in Tailevu showing the effects of global warming.



Effect of Global Warming at Nasau Village in Ra.



#### **Exercise 5.4.1**

1. Climate change is considered to be one of the most important environmental issues of our time.
  - i. Differentiate between global warming and climate change.
  - ii. State ways in which climate change can be mitigated.
2. Kiribati and Marshall Islands have low lying atolls. Describe one impact that global warming can have on these island nations.
3. Describe one impact that ocean acidification can have on coral reef systems.
4. What is the difference between natural and enhanced greenhouse effect?
5. List some greenhouse gases.
6. Describe the impact of global warming on corals.
7. Make a presentation outlining what you can do to help reduce the effect of global warming in Fiji:
  - i) At home
  - ii) In school
  - iii) In town or city
  - iv) In your community

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# GLOSSARY

**Acid** - any chemical that gives off protons or  $H^+$  in water. Acids have a pH less than 7. They turn the pH indicator phenolphthalein colourless and turn litmus paper red.

**Acid-base reaction** - a reaction involving a procedure which is used to determine the concentration of an acid or base. A measured volume of an acid or base of known concentration is reacted with a sample to the equivalence point.

**Activity series** - is a list of metals ranked in order of decreasing reactivity to displace hydrogen gas from water and acid solutions. It can also be used to predict which metals will displace other metals in aqueous solutions.

**Allotrope** - elements that have different structures (and therefore different forms), such as Carbon (diamonds and graphite)

**Alkali metals** - the metals of Group 1 on the periodic table

**Alkane** - an organic molecule that only contains single carbon-carbon bonds with general formula  $C_nH_{2n+2}$ .

**Alkene** - an organic molecule that contains at least one  $C=C$  or carbon-carbon double bond with general formula  $C_nH_{2n}$

**Alkyne** - A compound that consists of only carbon and hydrogen that contains at least one carbon-carbon triple bond with general formula  $C_nH_{2n-2}$ .

**Ammonia** - is a colourless gas with a sharp, characteristic odour. It is easily liquefied by pressure, and is very soluble in water. Ammonia acts as a weak base.

**Ammonia fountain** - is a type of chemical demonstration which consists of introducing water through an inlet to a container filled with ammonia gas. Ammonia dissolves into the water and the pressure in the container drops. As a result more water is forced into the container from another inlet creating a fountain effect.

**Anhydrous** - A compound with all water removed, especially water of hydration. For example, strongly heating copper (II) sulphate pentahydrate ( $CuSO_4 \cdot 5H_2O$ ) produces anhydrous copper (II) sulphate ( $CuSO_4$ ).

**Anion** - negatively charge ions

**Anode** - the positive side of a dry cell battery or a cell

**Aqueous**-a substance dissolved in water.

**Atom** - a chemical element in its smallest form, and is made up of neutrons and protons within the nucleus and electrons circling the nucleus

**Atomic number** - the number representing an element which corresponds with the number of protons within the nucleus

**Base** - a substance that accepts a proton and has a high pH; a common example is sodium hydroxide (NaOH)

**Binary compound** - A compound that contains two different elements. NaCl is a binary compound; NaClO is not.

**Blast Furnace** - is a type of metallurgical furnace used for smelting to produce industrial metals, generally iron, but also others such as lead or copper.

**Boiling point**- the temperature in which the substance starts to boil

**Bond** - the attraction and repulsion between atoms and molecules

**Boyle's law** - states that the pressure of an ideal gas is inversely proportional to its volume, if the temperature and amount of gas is held constant. Doubling gas pressure halves gas volume, if temperature and amount of gas don't change.

**Burette** - glassware used to dispense specific amounts of liquid when precision is necessary (e.g. titration and resource dependent reactions)

**Calibration** - is correcting a measuring instrument by measuring values whose true values are known. Calibration minimizes systematic error<sup>+</sup>.

**Cathode** - is the electrode which gains electrons or is reduced. In other words, it is where reduction occurs in an electrochemical cell.

**Cation** - positively charged ion

**Centrifuge** - equipment used to separate substances based on density by rotating the tubes around a centred axis

**Charles's law** - is an experimental gas law which describes how gases tend to expand when heated. A modern statement of Charles's law is: When the pressure on a sample of a dry gas is held constant, the Kelvin temperature and the volume will be directly related

**Chemical change** - occur when a substance combines with another to form a new substance

**Chemical equation** - is the symbolic representation of a chemical reaction wherein the reactant entities are given on the left-hand side and the product entities on the right-hand side

**Chemical property** - is any of a material's properties that becomes evident during a chemical reaction; that is, any quality that can be established only by changing a substance's chemical identity.

**Chemical reaction**-the change of one or more substances into another or multiple substances

**Chromatography**-the separation of a mixture of substances using filter paper and a solvent

**Climate change**-the change of weather patterns that is often linked to increase in global warming

**Coke**-is a fuel with few impurities and a high carbon content, usually made from coal

**Combustion**-or burning is a high-temperature exothermic chemical reaction between a fuel and an oxidant, usually atmospheric oxygen that produces oxidized, often gaseous products, in a mixture termed smoke.

**Compound**- a substance that is made up of two or more chemically bonded elements

**Concentrated solution**- A solution that has a large ratio of the amounts of solute to solvent.

**Concentration**-the amount of one type of substance dissolved in another. It is measured in g/L or mol/L

**Concept map**-is a diagram that depicts suggested relationships between concepts. It is a graphical tool that designers, engineers, technical writers, and others use to organize and structure knowledge.

A concept map typically represents ideas and information as boxes or circles, which it connects with labelled arrows in a downward-branching hierarchical structure. The relationship between concepts can be articulated in linking phrases such as *causes*, *requires*, or *contributes to*.

**Control**-the part of an experiment that act as a standard by which to compare experimental observations.

**Cooling curve**-is a line graph that represents the change of phase of matter, typically from a gas to a solid or a liquid to a solid. The independent variable (X-axis) is time and the dependent variable (Y-axis) is temperature

**Corrosive**- a substance that will destroy and damage other substances with which it comes into contact. It may attack a great variety of materials, including metals and various organic compounds, living tissue: it causes chemical burns on contact.

**Covalent bond**- a chemical bond formed when two atoms share two electrons

**Cross-over method**-is a shortcut that can help to determine the formula of an ionic compound

**Crystallization**-the formation of crystals when a saturated solution is left to cool

**Decantation**- is a process for the separation of mixtures, by removing a layer of liquid, generally one from which a precipitate has settled.

**Decomposition**- the breakdown of a substance into two or more products

**Density**-the mass of a substance divided by its volume

**Diffusion**- When particles move from areas of high concentration to areas of low concentration. For example, if you open a bottle of ammonia on one end of the room, the concentration of ammonia molecules in the air is very high on that side of the room. As a result, they tend to migrate across the room, which explains why you can smell it after a little while.

**Dilute solution**- A solution in which the ratio of the quantities of solute to solvent is small.

**Displacement**-a reaction in which one atom or group of atoms replaces another in a compound

**Double bond**- sharing of two pairs of electrons

**Double displacement**- When the cations of two ionic compounds switch places

**Element**-an atom that is defined by its atomic number

**Electrical conductivity**-the flow of electric current through a substance. Conduction is due to moving electrons in metals and moving ions in ionic solutions or molten form

**Electron**-a subatomic particle with a net charge that is negative

**Electron configuration**-is the pattern in which electrons are arranged in shells around the nucleus of an atom

**Electron shell**-an orbital around the atom's nucleus that has fixed number electrons (usually two or eight)

**Electrolysis** -the breakdown of compound using electricity

**Electrolyte**-a solution that conducts a certain amount of current and can be split categorically as weak and strong electrolytes

**Energy level**- A possible level of energy that an electron can have in an atom.

**Enhanced Greenhouse Effect**- the addition of more heat trapping gases to the atmosphere. The extra greenhouse gases are not 'natural' in origin, and are primarily due to the burning of fossil fuels (coal, oil and natural gas), with smaller contributions from land use changes, such as land clearing. The increased concentrations of greenhouse gases in the atmosphere are contributing to a warming of the earth's surface

**Enthalpy**-measure of the total energy of a thermodynamic system (usually symbolized as  $H$ )

**Enthalpy of fusion**- is the change in enthalpy resulting from heating a given quantity of a substance to change its state from a solid to a liquid

**Enthalpy of sublimation**- is the change in enthalpy resulting from heating a given quantity of a substance to change its state from solid to gas

**Enthalpy of vaporization**- is the enthalpy change required to transform a given quantity of a substance from a liquid into a gas

**Experiment**- is an orderly procedure carried out with the goal of verifying, disproving, or establishing the validity of a hypothesis

**Formula**- is a concise way of expressing information symbolically as in a mathematical or chemical formula

**Formula mass**- the sum of the atomic weights of the atoms in the empirical formula of the compound

**Global warming**- is the increases in the Earth's average temperature because of greenhouse gas build up in the atmosphere

**Graphite**- form of carbon that conducts electricity and is used for inert electrodes

**Greenhouse Effect**- is a process by which thermal radiation from a planetary surface is absorbed by atmospheric greenhouse gases, and is re-radiated in all directions.

**Greenhouse gas**-a gas such as methane or carbon dioxide that absorbs infrared radiation from the Earth's surface

**Group**- A column (the things up and down) in the periodic table. Elements in the same group tend to have the same properties

**Haematite**- is the iron ore, a mineral form of iron (III) oxide ( $\text{Fe}_2\text{O}_3$ ), one of several iron oxides

**Halogen**-Group 7 on the Periodic Table and are all non-metals

**Heating curve**- is a plot of the temperature versus time when a substance is heated

**Hypothesis**- is a proposed explanation for a phenomenon which requires a scientific method to test it. Scientists generally base scientific hypotheses on previous observations that cannot satisfactorily be explained with the available scientific theories.

**Hydrocarbon**- A molecule containing carbon and hydrogen

**Hydrate**- is a substance that contains water or its elements

**Immiscible**- When two substances don't dissolve in each other. Think of oil and water. They're immiscible. Organic compounds and water are frequently immiscible

**Indicator**-a special compound added to solution that changes colour depending on the acidity of the solution; different indicators have different colours and effective pH ranges

**Inert**-does not react

**Inference**- is the act or process of deriving logical conclusions from premises known or assumed to be true

**Ion**-a molecule that has gained or lost one or more electrons

**Ionic bond**-electrostatic attraction between oppositely charged ions

**Ionic equation**- describe the chemical reaction while clearly indicating which of the reactants and/or products exist primarily as ions

**Insoluble**-a substance that will not dissolve

**Isotopes**- When an element has more than one possibility for the number of neutrons, these are called isotopes. All known elements possess isotopes

**Kinetic theory**-the idea that the arrangement and motion of the atoms can explain states of matter, diffusion and rates of reaction

**Law of Conservation of mass**- is a relation stating that in a chemical reaction , the mass of the products equals the mass of the reactants because matter is neither created nor destroyed during a chemical reaction

**Law of Definite Composition**- states that the elements in a given compound are always combined in the same proportion by mass. This law form the basis for the definition of a chemical compound

**Law of Definite Proportion**- states that a chemical compound always contains exactly the same proportion of elements by mass

**Metal**-Chemical element that is a good conductor of both electricity and heat and forms cations and ionic bonds with non-metals

**Metalloid**-A substance possessing both the properties of metals and non-metals

**Melting point**- is the temperature at which a substance changes state from solid to liquid at atmospheric pressure

**Molecule**-a chemically bonded number of atoms that are electrically neutral

**Molecular mass**- It is calculated as the sum of the mass of each constituent atom multiplied by the number of atoms of that element in the molecular formula

**Mixture**-two or more substances mixed together but not chemically combined

**Miscible**-relating to two or more substances, such as water and alcohol, that can be mixed together or can dissolve into one another in any proportion without separating

**Mole**-the relative formula mass of a substance in grams

**Neutralisation**- The reaction of an acid with a base to form water and a salt

**Neutron**-a neutral unit or subatomic particle that has no net charge

**Noble gas**-group 18 elements, those whose outer electron shell is filled

**Non-metal**-an element which is not metallic

**Observation**-is the active acquisition of information from a primary source by observation which employs the senses. It can also involve the recording of data via the use of instruments.

**Ore**-is a type of rock that contains sufficient minerals with important elements including metals that can be economically extracted from the rock. The ores are extracted through mining; these are then refined (often via smelting) to extract the valuable element(s).

**Oxidation**- When a substance loses electrons.

**Percentage composition**-the percent composition of a component in a compound is the percent of the total mass of the compound that is due to that component.

**Period**- A row (left to right) in the periodic table.

**Periodic table**-is a tabular arrangement of the chemical elements, organized on the basis of their atomic numbers (numbers of protons in the nucleus), electron configurations, and recurring chemical properties. Elements are presented in order of increasing atomic number, which is typically listed with the chemical symbol in each box.

**pH**-the measure of acidity (or basicity) of a solution

**pH scale**-a logarithmic scale for expressing the acidity or alkalinity of a solution

**Physical change**-is a change affecting the form or state of a chemical substance, but not its chemical composition.

**Physical properties**- A property which can be determined without changing something chemically.

**Pig iron**-is the intermediate product of smelting iron ore with a high-carbon fuel such as coke, usually with limestone as a flux. It is the molten iron from the blast furnace, which is a large and cylinder-shaped furnace charged with iron ore, coke, and limestone

**Pipette**-a pipette used to measure out solutions accurately

**Polypropylene**- is a thermoplastic polymer used in a wide variety of applications including packaging and labelling, textiles (e.g., ropes, thermal underwear and carpets), stationery, plastic parts and reusable containers of various types, laboratory equipment, loudspeakers, automotive components, and polymer banknotes. An addition polymer made from the monomer propylene, it is rugged and unusually resistant to many chemical solvents, bases and acids.

**Precipitate**-formation of a solid in a solution or inside another solid during a chemical reaction or by diffusion in a solid

**Proton**-a positive unit or subatomic particle that has a positive charge

**Pure chemicals**- there is only one substance present

**Qualitative analysis** way of identifying substances by carrying out chemical tests based on observations

**Quantitative analysis**

**Redox**- A reaction that has both an oxidation and reduction.

**Reduction**-the removal of oxygen from or the addition of addition of electrons to a substance

**Saturated solution**- When the maximum amount of solute is dissolved in a liquid

**Sedimentation**- is the tendency for particles in suspension to settle out of the fluid in which they are entrained, and come to rest against a barrier

**Separating funnel**- is a piece of laboratory equipment made mostly of glass and is used to separate two liquids that usually are very hard to separate, for example, Oil and Water.

**Significant figures**- are the digits in a number which are known precisely, plus one estimated digit.

**Silica**-is a chemical compound that is an oxide of silicon with the chemical formula  $\text{SiO}_2$

**Single bond**- is a chemical bond between two chemical elements involving two bonding electrons. In other words, one pair of electrons is shared.

**Slag**-the waste material formed between metals are extracted in a furnace

**Solubility**- A measurement of how much of a solute can dissolve in a liquid.

**Solubility curve**- is a graph of the solubility of a compound (grams/100 grams water on the Y axis) at various temperatures (Celsius on X Axis). **Spectator ion**-ion that does not take part in a reaction

**Standard solution**- is a solution containing a precisely known concentration of an element or a substance

**Strong acid**- An acid that fully dissociates in water

**Supersaturated**- When more solute is dissolved in a liquid than is theoretically possible.

**Synthesis**- When you make a big molecule from two or more smaller ones

**Teflon**- a "non-stick" chemical used on cookware.

**Terylene**- a synthetic polyester fibre or fabric based on terephthalic acid, characterized by lightness and crease resistance and used for clothing, sheets, ropes, sails, etc. US name (trademark): **Dacron**

**Titration**-the process of titrating one solution with another, also called volumetric analysis

**Titrimetric analysis**- involves measuring the volume of a reagent containing the desired component.

**Triple bond**-the sharing of three pairs of electrons within a covalent bond (example N<sub>2</sub>)

**Universal indicator**-is a pH indicator composed of a solution of several compounds that exhibits several smooth colour changes over a pH value range from 1-14 to indicate the acidity or alkalinity of solutions

**Unsaturated**- is a chemical compound that has a chain of carbon atoms linked together by single bonds and has hydrogen atoms filling all of the other bonding orbitals of the carbon atoms

**Valence electron**-the outermost electrons of an atom, which are located in electron shells

**Volumetric analysis**-any method of quantitative chemical analysis in which the amount of a substance is determined by measuring the volume that it occupies or, in broader usage, the volume of a second substance that combines with the first in known proportions, more correctly called titrimetric analysis

**Water of crystallisation**-the water combined with a compound in a crystal

**Weak acid**-acid that is only slightly ionised when dissolved in water

**Wrought iron**- is an iron alloy with a very low carbon (0.04 to 0.08%) content in contrast to cast iron (2.1% to 4%), and has fibrous inclusions known as slag up to 2% by weight

# APPENDIX

## PERIODIC TABLE OF THE ELEMENTS

Group I	II	III	IV	V	VI	VII	VIII
Row 1	1 H						2 He
2	3 Li	5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	21 Sc	23 V	24 Cr	25 Mn	26 Fe	27 Co
5	37 Rb	39 Y	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh
6	55 Cs	57 La	73 Ta	74 W	75 Re	76 Os	77 Ir
7	87 Fr	89 Ac	104 Ku	104 Ku	104 Ku	104 Ku	104 Ku

+ Lanthanide Series

+ Actinide Series

58- Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
140	141	144	(145)	150	152	157	159	162	165	167	169	173	175
80	91	92	93	94	95	96	97	98	99	100	101	102	103
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
232	231	238	237	(244)	(243)	(247)	(247)	(252)	(254)	(257)	(258)	(259)	(260)

( ) mass number of most stable isotope.

Key:

Z	Atomic number
X	symbol of element
A	Relative atomic mass