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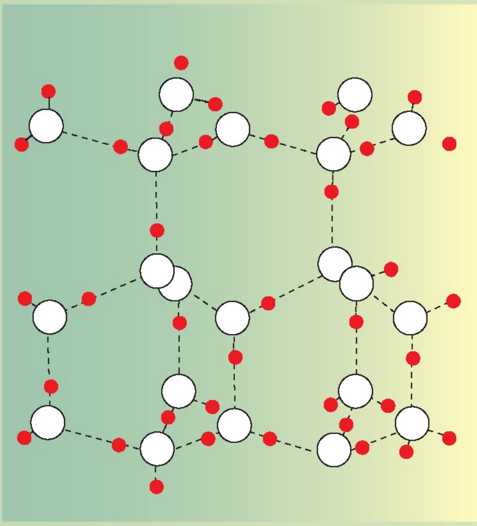
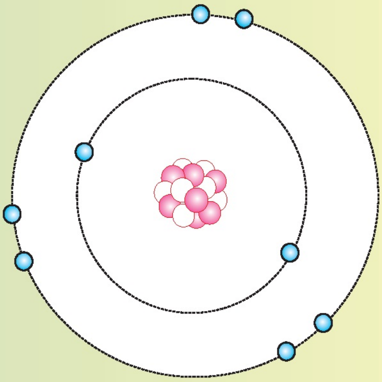
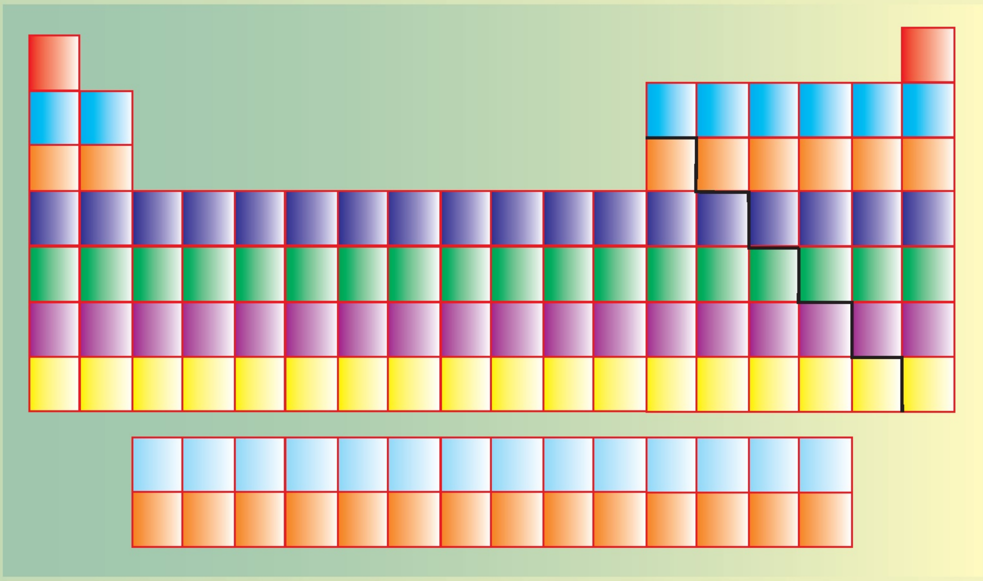
Longhorn Secondary
Chemistry
Form 2







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Chemistry

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TABLE OF CONTENTS

UNIT 1: Structure of the atom and the periodic table

1.1 The structure of the atom

Names and symbols of the first twenty elements of the periodic table

In Form I, we learnt that there are about 119 known elements. We have also learnt names and symbols of some common elements that can be arranged according to their behaviour under different conditions.

In this class, we shall learn how to arrange elements according to their behaviour. Continue to practise the use of chemical symbols particularly those of the first 20 elements of the periodic table. These are listed in Table 1.0 below.

Table 1.0: Names and symbols of the first twenty elements of the periodic Table

	Name of element	Chemical symbol
1	Hydrogen	H
2	Helium	He
3	Lithium	Li
4	Beryllium	Be
5	Boron	B
6	Carbon	C
7	Nitrogen	N
8	Oxygen	O
9	Fluorine	F
10	Neon	Ne
11	Sodium	Na
12	Magnesium	Mg
13	Aluminium	Al
14	Silicon	Si
15	Phosphorus	P
16	Sulphur	S
17	Chlorine	Cl
18	Argon	Ar
19	Potassium	K
20	Calcium	Ca

Simple structure of the atom

Atoms are very small particles of an element, but they contain even smaller particles called *sub-atomic particles*. There are three kinds of sub-atomic particles. These are the *protons*, *electrons* and *neutrons*. The protons and neutrons are found in the *nucleus*. The nucleus is the middle part of the atom. It is a very dense and extremely small part of the atom. Outside the nucleus is a much larger region of the atom where electrons occur in energy levels at different distances from the nucleus of the atom. The electrons can be imagined as circulating the nucleus, at great speeds in those energy levels. We can represent the structure of an atom simply as follows:

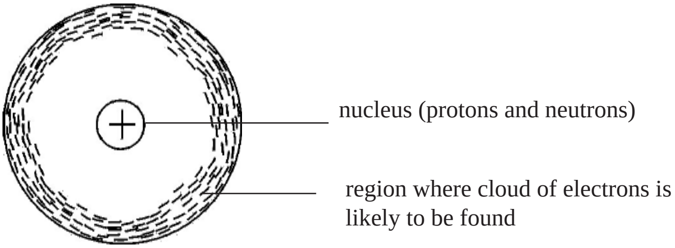


Fig. 1.1: Atom showing nucleus and electrons

Electrons are negatively charged (–ve), protons are positively charged (+ve) and neutrons have no charge. Therefore the nucleus has a positive charge because of the protons as illustrated in Fig. 1.1.
If we magnify the nucleus, we can show the sub-atomic particles (protons and neutrons) as in Fig. 1.2.

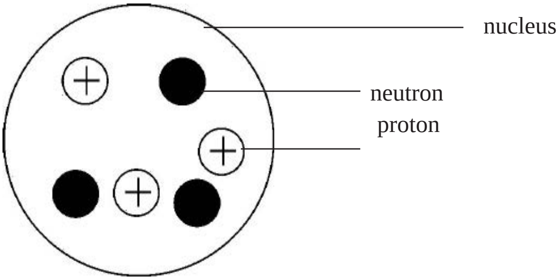
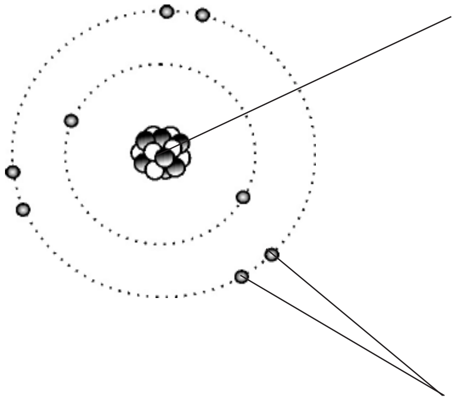


Fig.1.2: Nucleus containing protons and neutrons

Since protons have similar charge (+ve) they repel each other. However, the repulsion is minimised by the presence of the neutrons as illustrated in Fig. 1.3.



The nucleus

- It is at the **centre** of the atom.
- It contains **protons** and **neutrons**.
- It has a **positive charge** because of the **protons**.
- **The whole mass** of the atom is **concentrated in the nucleus**.
- However, it is **tiny** compared to the size of the atom.

The electrons

- They move **around** the nucleus.
- They are **negatively charged**.
- They are **tiny,** but cover a lot of space.
- They occupy the energy levels.

Fig. 1.3: Atomic structure showing protons and neutrons in the nucleus and electrons in the energy level

In a neutral atom:

Number of protons = Number of electrons

This means that each proton (+ve) has an electron (–ve) as a partner.

Table 1.1 Particles present in an atom

Sub-atomic particle	Mass	Charge	Where found in atom
Proton	1	+ (positive)	inside nucleus
Neutron	1	0 (neutral)	inside nucleus
Electron	$\frac{1}{1840}$	– (negative)	outside nucleus

The structure of the atom

In a well planned town or city, houses are constructed according to the plan. The house rent also varies. Those people who live in some neighbourhoods of the city are wealthy people and therefore pay a lot more than those who live in other areas.

Atomic "city"

These are found in a country where people live according to the following city plan.

1. There are three suggested types of houses that are allowed in each neighbourhood.
   Only 2 people are allowed to occupy any house.

1st neighbourhood	-	has one low cost stone house only
2nd neighbourhood	-	has four middle cost stone houses only
3rd neighbourhood	-	has one high cost stone house only

2. All houses must be built away from city/town centre.

How many people are in:

(a) Low cost houses in the 1st neighbourhood?
(b) Middle cost houses in the 2nd neighbourhood?
(c) What is the maximum number of high cost houses expected to be in the 3rd neighbourhood?

We notice that the further the houses are from the city centre, the more expensive they are. The cheapest houses are in the first neighbourhood. So the people who live in the 3rd neighbourhood are very wealthy. From our atomic city vocabulary list, we mentioned that the houses are settled according to wealth and ability to pay higher. This implies that the 3rd neighbourhood has wealthy people. The 2nd has poor people and the 1st has the poorer people. In other words the wealth <u>increases</u> as we move from the city centre.

To understand the atomic structure, let us look at a suggested analogy of atomic 'city'.

The city plan (vocabulary list) is as follows:

City = Atom
People = Electrons
Neighbourhoods = Energy levels
Wealth = Energy
City centre = Nucleus

Look at Fig. 1.4. What would you get if you completed the dotted lines?

The neighbourhoods are represented by energy levels where an energy level is the path of an electron around a nucleus in an atom.

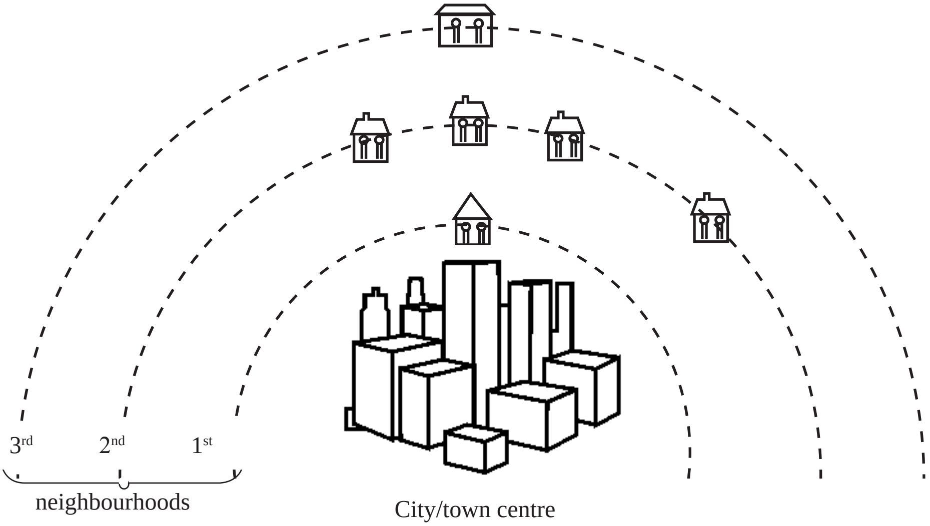


Fig. 1.4: Atomic 'city' showing the neighbourhoods.

Activity 1.0

Complete the dotted lines in Fig. 1.4 to form circles. Represent each number of people in each type of house by dots (•) or crosses (x), and distribute them in each neighbourhood. Compare your diagram with Fig. 1.5 which represents a magnesium atom.

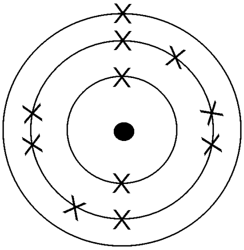


Fig. 1.5: Structure of a magnesium atom

The various energy levels in an atom are represented by a series of circles sharing the same centre (nucleus), separated from each other by roughly equal distances. The nucleus of the atom is at the centre of the circles. The electrons in the energy levels are represented by dots (•) or crosses (x). The energy levels are labelled 1st, 2nd, 3rd, 4th and so on starting from the one nearest to the nucleus as shown in Fig. 1.6.

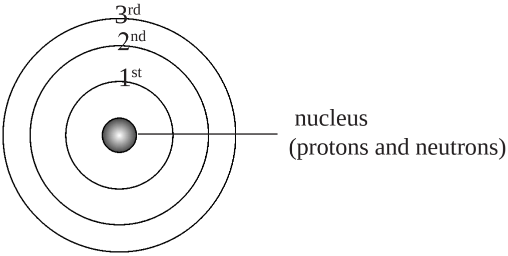


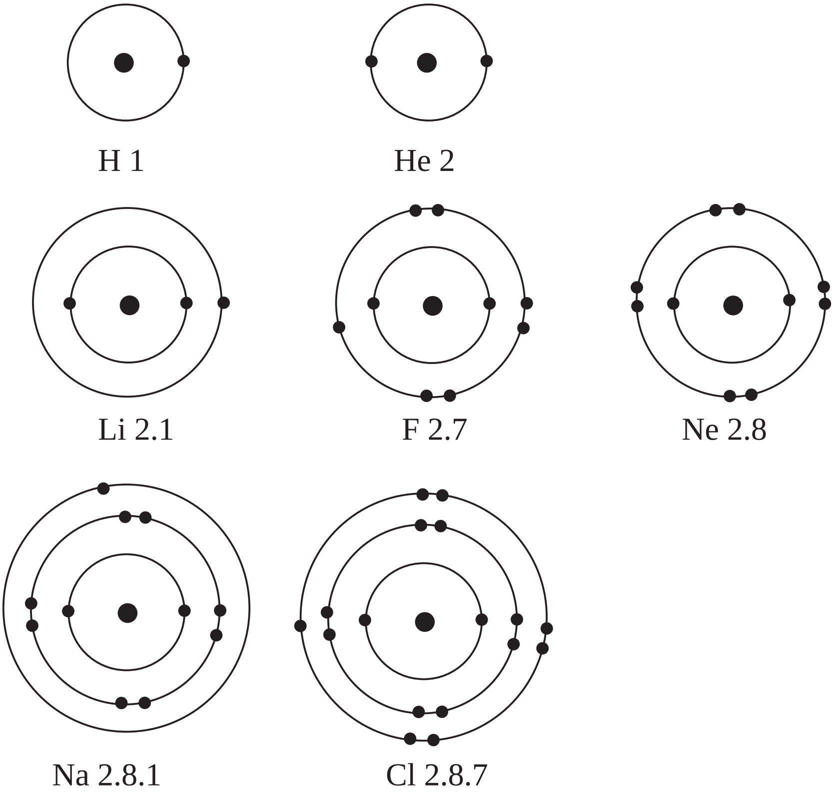
Fig. 1.6: Energy levels

The electrons that occupy the 1st energy level have lower energy than those in the 2nd energy level. Subsequently those in the 2nd energy level have lower energy than those in the 3rd energy level and so on. The 1st energy level usually has maximum of two electrons while the 2nd energy level has a maximum of eight (8) electrons. The maximum number of electrons that an energy level can hold are given by $2n^2$. Where n = number of energy level. 1st, 2nd, 3rd, etc. This means that when we draw the atomic structure, we should never put more than two electrons in the 1st energy level or more than 8 in the 2nd level. The third energy level is usually not full but can accommodate one to eight electrons. Third energy level holds a maximum of 8 electrons for the first 20 elements. The first twenty elements of the period table have 1 to 20 protons and 1 to 20 electrons. We can summarise the electrons in each energy level as shown in Table 1.2.

Table 1.2 The electron arrangements of the first twenty elements

Element	Symbol	Number of protons	1st Energy level	2nd Energy level	3rd Energy level	4th Energy level	Electron arrangement
Hydrogen	H	1	•				1
Helium	He	2	••				2
Lithium	Li	3	••	•			2.1
Beryllium	Be	4	••	••			2.2
Boron	B	5	••	•••			2.3
Carbon	C	6	••	••••			2.4
Nitrogen	N	7	••	•••••			2.5
Oxygen	O	8	••	••••••			2.6
Fluorine	F	9	••	•••••••			2.7
Neon	Ne	10	••	••••••••			2.8
Sodium	Na	11	••	••••••••	•		2.8.1
Magnesium	Mg	12	••	••••••••	••		2.8.2
Aluminium	Al	13	••	••••••••	•••		2.8.3
Silicon	Si	14	••	••••••••	••••		2.8.4
Phosphorous	P	15	••	••••••••	•••••		2.8.5
Sulphur	S	16	••	••••••••	••••••		2.8.6
Chlorine	Cl	17	••	••••••••	•••••••		2.8.7
Argon	Ar	18	••	••••••••	••••••••		2.8.8
Potassium	K	19	••	••••••••	••••••••	•	2.8.8.1
Calcium	Ca	20	••	••••••••	••••••••	••	2.8.8.2

Arrangement of electrons in an atom helps us to explain the patterns in properties of the elements. These properties are the basis of the periodic table which we will discuss later. Now let us illustrate the electron arrangement of the following atoms: Hydrogen, helium, lithium, fluorine, neon, sodium and chlorine respectively.



H 1
He 2
Li 2.1
F 2.7
Ne 2.8
Na 2.8.1
Cl 2.8.7

Fig. 1.7: Atomic structure of H, He, Li, F, Ne, Na and Cl

- Draw diagrams to show the electron arrangement of the following atoms: K, Ar, Ca and S.

1.2 Atomic characteristics

Atomic number and mass number

We have learnt how to draw atomic structure showing the nucleus and arrangement of electrons in various energy levels. It is important at this stage to know how to calculate the number of protons and neutrons inside the nucleus. This will enable us to show the number of these sub-atomic particles when we draw atomic structure diagrams. Atomic number and mass number are two simple numbers that tell us something about an atom.

Note the following points:

- The *atomic number* tells us how many protons are there in a nucleus. It is denoted by letter Z.

- It also tells us the number of electrons in an atom.
- The atomic number Z of an element = The number of protons in the nucleus of its atom = The number of electrons in an atom
- The *mass number* is the sum of protons and neutrons, therefore it is always bigger than the atomic number.
- Mass number is denoted by letter A.
- Roughly the mass number is double the atomic number.
- To get the number of neutrons, we just subtract the atomic number from the mass number, that is A – Z. Study Table 1.3 which shows the relationship between atomic number, number of neutrons and mass number for some elements.

Table 1.3 Relationship between the number of protons, neutrons and mass number of some elements

Atom	Symbol	Number of protons: Atomic number, Z	Neutrons	Mass number, A
Hydrogen	H	1	0	1
Carbon	C	6	6	12
Nitrogen	N	7	7	14
Sodium	Na	11	12	23
Chlorine	CI	17	18	35

From Table 1.3, we notice that mass number, A, is a sum of protons and neutrons i.e

Mass number A = number of protons + number of neutrons.

If we represent neutrons by N, we see that $A = Z + N$

Using letters A, Z and N, how can we get the number of neutrons N, in the nucleus of an atom?

The answer to this question is $N = A - Z$

Example

Calculate the number of neutrons in a chlorine atom given that the atomic number, $Z = 17$ and mass number, $A = 35$

The number of neutrons $= A - Z$

$= 35 - 17 = 18$

Therefore, a chlorine atom has 18 neutrons. Remember that since the number of protons and electrons are equal in an atom, once you know the atomic number which represents the number of protons, you automatically know the number of electrons in the atom. Suppose you are required to draw the atomic structure of an atom showing the sub-atomic particles as in the following example:

A certain element X has atomic number 6 and mass number 14.
Draw the atomic structure of element X showing all the sub-atomic particles.
The sub-atomic particles to show are protons, neutrons and electrons.
Atomic number = number of protons = number of electrons = 6
Number of neutrons = mass number – atomic number or A - Z
= 14 – 6 = 8 neutrons

This means that we will show 6 protons and 8 neutrons in the nucleus. The 6 electrons will be indicated in the energy levels as shown in Fig. 1.8.

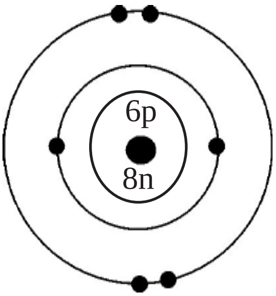


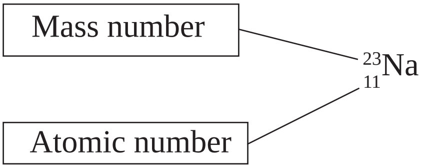
Fig. 1.8: Atomic structure of X

Exercise
Now draw the atomic structure of an atom Y showing the sub-atomic particles given the following information:
Atomic number, Z = 8
Mass number, A = 17

Usually, the atomic number, Z, and mass number A, of an atom of an element X can be written alongside the symbol of that element, one as a superscript and the other a subscript as shown below:

Mass number, → A
Element, → X
Atomic number, → Z

So the symbol for an atom of sodium would be written as



Note: The top number is referred to as the superscript and bottom number as the subscript.

We can also represent the sub-atomic particles using symbols with the mass of the particle as a superscript and the charge as a subscript as follows:

Proton – $^{1}_{1}p$

Neutron – $^{1}_{0}n$

Electron – $^{0}_{-1}e$

An electron has zero mass and charge –1. Can you tell what the letters p and n stand for?

Isotopes

We are all familiar with charcoal. May be some of us know how charcoal is formed. Burning of charcoal should be discouraged because cutting down trees to make charcoal interferes with our water catchment areas. What is the colour of charcoal? What is charcoal in chemistry?

Charcoal is black in colour. In chemistry charcoal is simply carbon. A piece of charcoal has millions and millions of carbon atoms. Let us represent a piece of charcoal as in Fig. 1.9.

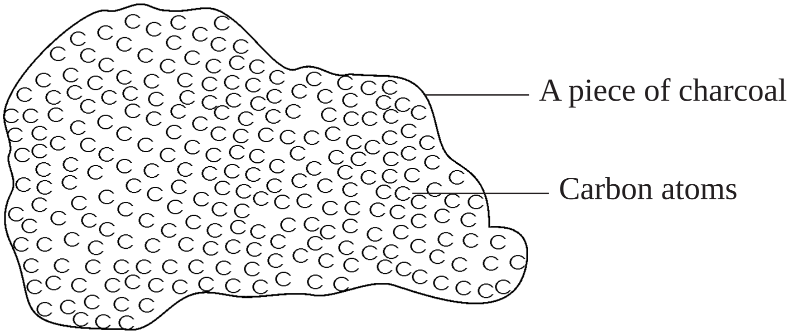
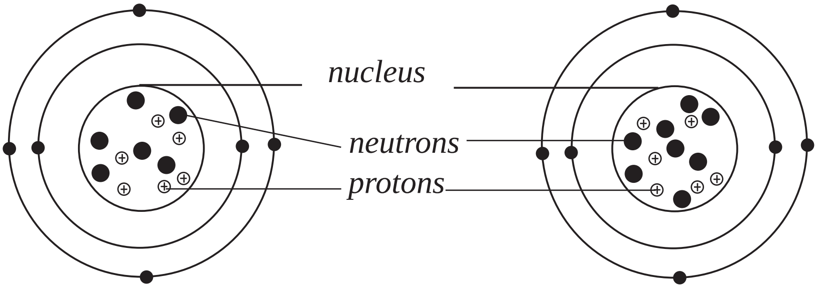


Fig. 1.9: Millions of carbon atoms that make charcoal

Carbon atoms exist in two forms represented as $^{12}_{6}C$ and $^{14}_{6}C$

The two atoms can be represented using diagrams as shown in Fig. 1.10.



Carbon-12 *Carbon-14*

Fig. 1.10: Carbon isotopes

What difference do you notice between the two carbon atoms?
The bottom numbers are the same but the top numbers are different. This means that they have the **same atomic number** but **different mass numbers**.
What causes the difference in mass number?
Calculate the sub-atomic particles for each of the carbon atoms.
You will notice that the different masses are caused by 2 extra neutrons in the nucleus of carbon-14. Such atoms are called **isotopes**.

Isotopes are atoms of the same element, which have the same number of protons in the nucleus (atomic number) but different mass numbers.
Some elements are made up of just one type of atom, while others exist as mixtures of isotopes. Some of these elements are listed in Table 1.4.

Table 1.4 Examples of isotopes

(a) Hydrogen

HYDROGEN	**Hydrogen (99.99%)**	**Deuterium (0.01%) Heavy hydrogen**	**Tritium (Trace)**
Symbols of isotopes of hydrogen	${}^{1}_{1}H$	${}^{2}_{1}H$	${}^{3}_{1}H$
No. of protons	1	1	1
No. of neutrons	0	1	2
No. of electrons	1	1	1
Mass number (p + n)	1	2	3

(b) Carbon

CARBON	Carbon-12 (98.9%)	Carbon -13 (1.1 %)	Carbon -14 (trace)
Symbols of carbon isotopes	${}^{12}_{6}C$	${}^{13}_{6}C$	${}^{14}_{6}C$
Protons	6	6	6
Neutrons	6	7	8
Electrons	6	6	6
Mass number (p + n)	12	13	14

(c) Chlorine

CHLORINE	Chlorine -35 (75 %)	Chlorine -37 (25%)
Symbols of chlorine isotopes	${}^{35}_{17}Cl$	${}^{37}_{17}Cl$
Protons	17	17
Neutrons	18	20
Electrons	17	17
Mass number (p + n)	35	37

(d) Oxygen

OXYGEN	Oxygen -16	Oxygen -17	Oxygen -18
Symbols of oxygen isotopes	${}^{16}_{8}O$	${}^{17}_{8}O$	${}^{18}_{8}O$
Protons	8	8	8
Neutrons	8	9	10
Electrons	8	8	8
Mass number	16	17	18

Relative Atomic Mass

Atoms are so tiny that their masses cannot be measured on a balance or scale directly. We cannot weigh an atom! Hydrogen is the lightest known element.

A hydrogen atom is therefore the lightest. Originally, the mass of a hydrogen atom was taken as the standard reference atom. Its atomic mass was arbitrarily fixed as one unit i.e H = 1.

The mass of any other element was found by comparing its mass with that of hydrogen. The idea was to find out how many times the atom of another element is as heavy as one atom of hydrogen hence relative atomic mass (R.A.M) with a symbol (A_r). This can be expressed mathematically as follows.

$$\text{Relative Atomic Mass} = \frac{\text{mass of 1 atom of the element}}{\text{mass of 1 atom of hydrogen}}$$

For instance an oxygen atom (O), has a mass of 16. This means one oxygen atom is 16 times heavier than one atom of hydrogen.

Sometimes the symbol A_r is used for R.A.M with the symbol of the atom in parenthesis after the symbol. For example, A_r (O) means relative atomic mass of oxygen.

Many changes have occurred since the *hydrogen scale* was introduced. In the twentieth century, the hydrogen scale was replaced by oxygen as the standard since oxygen combines with most elements. But later the oxygen scale was found to be unsatisfactory because oxygen has several isotopes which would have different masses depending on which oxygen isotope was being considered (see Table 1.4) (d).

In 1961 the International Union of Pure and Applied Chemistry (IUPAC) recommended as the standard the most abundant of the carbon isotopes, carbon - 12 (${}^{12}_{6}\text{C}$). It has an abundance of 98.9%.

Nowadays, the atomic mass of an element is measured by comparing it with $\frac{1}{12}$ of the mass of one atom of a carbon-12 (${}^{12}_{6}\text{C}$).

One atom of ${}^{12}_{6}\text{C}$ isotope is taken to have a mass of 12.00 units.

$$\frac{1}{12} \text{ of the mass of one atom of carbon} - 12 = 1.00 \text{ units}$$

$$A_r \text{ of an element} = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12} \times \text{mass of one atom of carbon-12}}$$

The relative atomic mass (A_r) of an element, is defined as *the average mass of one atom of the element compared with* $\frac{1}{12}$ *of the mass of one atom of carbon - 12.*

Note: Relative atomic mass has no units. It is a ratio of two masses. The relative atomic masses are not whole numbers like mass numbers. This is because the abundance of isotopes of an element is different.

For example a sample of chlorine is a mixture of two isotopes, ^{35}Cl and ^{37}Cl in the ratio of 3 : 1 respectively, i.e 75% is ^{35}Cl and only 25% is ^{37}Cl.

Calculation of the Relative Atomic Mass

As mentioned earlier many elements naturally consist of a mixture of isotopes. The abundance in which the isotopes occur in an element differs in different elements. This is why the term ‘average’ mass of one atom is used in the definition above. The proportions (abundance) in which the isotopes occur in an element may be stated as:

- a ratio
- percentage of the total
- a fraction of the total.

Example 1: You are given ratio abundance

Chlorine consists of two isotopes, chlorine – 35 and chlorine – 37 in the ratio 3 : 1. Calculate the relative atomic mass (R.A.M) of chlorine.

Solution

Suppose the sample contains 4 atoms of chlorine, in the ratio 3 : 1, respectively, then 3 atoms will each have a mass of 35 and 1 atom will have a mass of 37.

the total mass of $^{35}Cl = 35 \times 3$

While the total mass of $^{37}Cl = 37 \times 1$

Therefore, the average mass of chlorine atoms will be

$$\frac{\text{Total mass of all atoms}}{\text{Total }\textbf{number}\text{ of atoms}} = \frac{(35 \times 3) + (37 \times 1)}{4}$$

$$\text{The R.A.M} = 35.5$$

Note: R.A.M has no units. It is not the mass number since chlorine has 2 mass numbers; 35 and 37. The R.A.M is also nearer to the mass number of the more abundant isotope, i.e, ^{35}Cl.

Example 2: You are given percentage abundance

Silicon (Z = 14) consists of three isotopes: silicon – 28, 92.2%, silicon – 29, 4.7% and silicon - 30, 3.1%.
Find the relative atomic mass of silicon.

Solution

Percentage abundance simply means that if we have 100 atoms of an element called silicon, 92.2 atoms will each have a mass of 28

Therefore, the total mass of these $= \frac{28 \times 92.2}{100}$

4.7 atoms will be silicon – 29

Therefore, the total mass of these $= \frac{29 \times 4.7}{100}$

3.1 atoms will be silicon – 30

Therefore, the total mass of these $= \frac{3.1 \times 30}{100}$

The average mass of a silicon atom is

total mass of all atoms $= \frac{28 \times 92.2}{100} + \frac{4.7 \times 29}{100} + \frac{3.1 \times 30}{100}$

Therefore, R.A.M of silicon = 28.1

Example 3: Abundance is given as a fraction of the total

A sample of an element X consists of $\frac{9}{10}$ of ${}^{16}_{8}X$ and $\frac{1}{10}$ of ${}^{18}_{8}X$

Show that the relative atomic mass of X would be 16.2

Solution

$$\text{R.A.M} = (16 \times \frac{9}{10}) + (18 \times \frac{1}{10}) = 16.2$$

1.3 The periodic table

We have learnt about the structure of the atom and how the electrons are arranged in various energy levels. We are now going to use that knowledge to build up the periodic table.

Activity 1.1 Build up of the periodic table

Materials required

- 2 graph papers.
- Pens or pencils
- Rulers
- Pair of scissors or razor blades
- Glue stick

Procedure

1. Draw twenty squares on the graph papers (4 cm × 4 cm).
2. Cut out the squares.
3. On the top left hand corner of each square, label A on one side and B on the other side as illustrated in Tables 1.5 and 1.6.
4. On side A, write the electron arrangement of the element only. For example if it is magnesium, just write 2.8.2. Table 1.5.
5. On side B, write the actual symbol of the element at the centre of the paper. Include the atomic and mass numbers of the elements e.g. $^{24}_{12}Mg$ as in Table 1.6. Write the name of each element below the symbol.
7. With side A facing up, shuffle or mix the squares or rectangles randomly.
8. Take a piece of manila paper and draw a table of eight columns (up-down) and four rows (left to right) as shown in Table 1.7. The squares should be 5 cm × 5 cm.
9. Place the pieces of paper, with side A facing up on the squares on the manila paper as follows.
    - (i) Those with one energy level only, place them in the first row. If it has 1 electron in the outermost energy level place it in column I. If it has 2 electrons in the outermost energy level place it in column II.

6.2 The periodic table

Table 1.5 Side A

A 1	A 2	A 2 .1	A 2 .2	A 2 .3
A 2 .4	A 2 .5	A 2 .6	A 2 .7	A 2 .8
A 2 .8. 1	A 2 .8. 2	A 2 .8. 3	A 2 .8. 4	A 2 .8. 5
A 2 .8. 6	A 2 .8. 7	A 2 .8. 8	A 2 .8. 8. 1	A 2 .8. 8. 2

Table 1.6 Side B

B $^{1}_{1}H$ Hydrogen	B $^{4}_{2}He$ Helium	B $^{7}_{3}Li$ Lithium	B $^{9}_{4}Be$ Beryllium	B $^{11}_{3}B$ Boron
B $^{12}_{6}C$ Carbon	B $^{14}_{7}N$ Nitrogen	B $^{16}_{8}O$ Oxygen	B $^{19}_{9}F$ Fluorine	B $^{20}_{10}Ne$ Neon
B $^{23}_{11}Na$ Sodium	B $^{24}_{12}Mg$ Magnessium	B $^{27}_{13}Al$ Aluminium	B $^{28}_{14}Si$ Silicon	B $^{31}_{15}P$ Phosphorus
B $^{32}_{16}s$ Sulphur	B $^{35}_{17}Cl$ Chlorine	B $^{40}_{18}Ar$ Argon	B $^{39}_{19}K$ Potassium	B $^{40}_{20}Ca$ Calcium

(ii) If it has two energy levels, place it in row two. If there is only 1 electron in the outermost energy level, place it in column I; if it has 2 electrons in the outermost energy level, place it in column II and so on.

(iii) Place those with three energy levels in row three, and those with four energy levels in row number four.

Always place them under the correct columns according to the number of electrons in the outermost energy level.

(iv) Once you have arranged your papers neatly do not disturb them.
(v) Turn each piece without changing its position to side B. Arrange them neatly.
(vi) Now compare the pattern with an actual periodic table of elements. see appendix 1.
(vii) Paste your pieces of papers on the manila paper.

Table 1.7

	Column I	Column II	Column III	Column IV	Column V	Column VI	Column VII	Column VIII
Row one								
Row two								
Row three								
Row four								

The periodic table

The periodic table is organised as a big grid. In a grid, there are rows and columns. The periodic table has rows and columns too and they each mean something different. The elements are placed in specific columns because of the way they behave. They have similar chemical properties because they have the same number of electrons in their outermost energy level.
In a periodic table, each of the rows is called a *period*.

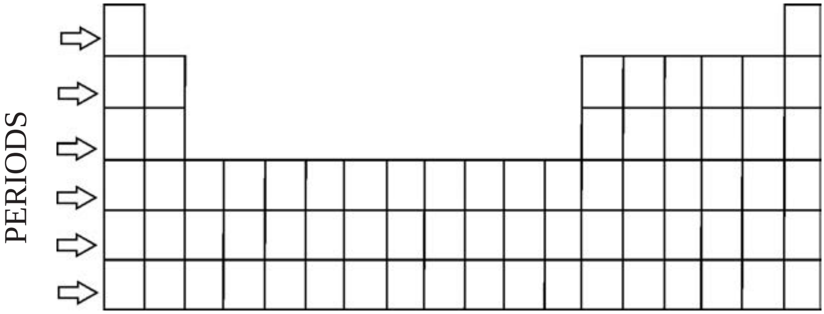


Fig. 1.11 Periods in the periodic table

Elements which are located in the same row or period have something in common. All the elements in a period have the same number of energy levels. If we want to know in which period to find an element we count the number of energy levels. If it has 4 energy levels; it is in the 4th period. Every element in the top row (the first period) has one energy level. All of the elements in the second row (the second period) have two energy levels. The maximum number of energy levels is seven representing period seven.

The periodic table has a special name for its columns too. A column in the periodic table is called a **group**. The elements in a group have the same number of electrons in their outermost energy level.

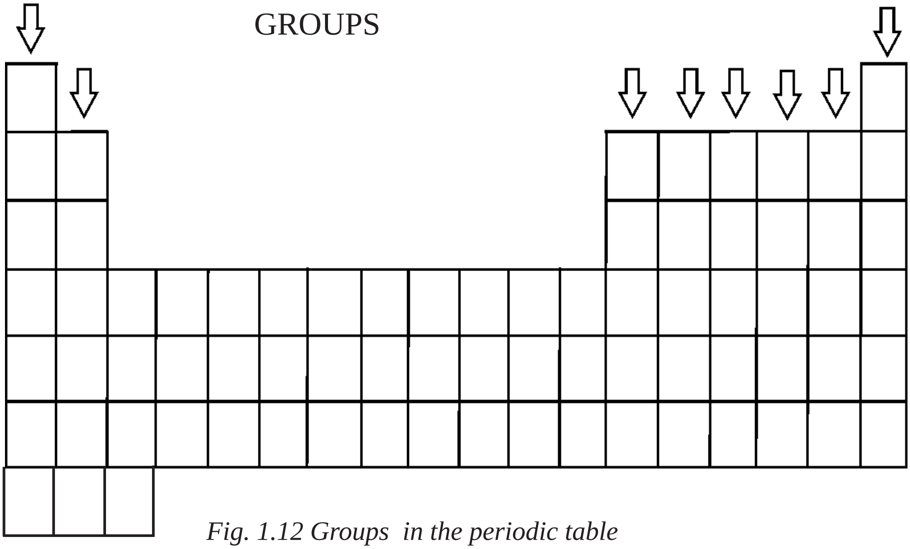


Fig. 1.12 Groups in the periodic table

An element in the first column belongs to group I. It has *one* electron in its outermost energy level. An element in the second column is in group II and has two electrons in the outermost energy level. As we keep counting the groups, we know how many electrons are in the outermost energy level. Therefore, the group number also is the number of electrons in the outermost energy level.

Electron arrangement and the periodic table

Electron arrangement short hand

Given the atomic number of an element, we can write the electron arrangement in shorthand. We shall illustrate this in the following exercise. The atomic numbers of X, Y and Z are 12, 13 and 17 respectively. Write the electron arrangement of each.

Solution.

X	2.8.2
Y	2.8.3
Z	2.8.7

All these three elements are in period 3 because each has three energy levels. X is in group two, Y in group 3 and Z in group 7, following the number of electrons in their outermost energy levels.

From activity 1.1, the order of the elements according to their electron arrangement should be as in Table 1.8 below.

Table 1.8 Electron arrangement of the first 20 elements of the periodic table

	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII
Period 1	**A** 1							**A** 2
Period 2	**A** 2.1	**A** 2.2	**A** 2.3	**A** 2.4	**A** 2.5	**A** 2.6	**A** 2.7	**A** 2.8
Period 3	**A** 2.8.1	**A** 2.8.2	**A** 2.8.3	**A** 2.8.4	**A** 2.8.5	**A** 2.8.6	**A** 2.8.7	**A** 2.8.8
Period 4	**A** 2.8.8.1	**A** 2.8.8.2						

Check in the periodic table of elements in Appendix 1. Where did you place hydrogen and helium in your build up of periodic table? Can you suggest why there is a difference in the one you built up and the ones in the textbook?

Hydrogen is a unique element. It is placed in Group I as its atoms have one electron in the outermost energy level. It also fits in Group VII because its atoms have one electron short of a fully filled first energy level.

Helium is different from all other elements. It has two electrons in its single energy level which can accommodate a maximum of 2 electrons. Therefore, its outermost energy level is full of the maximum electrons it can possibly accommodate. Later in this unit we will learn more about these Group VIII elements whose outermost energy levels are full with the maximum number of electrons.

1.4 Ion formation

Group VIII elements are described as being *stable*. This means that their outermost energy levels have the maximum number of electrons that they can accomodate as illustrated in Table 1.9.

Table 1.9 Electron arrangement of some Group VIII elements

Element	Symbol	Atomic Number	Electron arrangement
Helium	He	2	2
Neon	Ne	10	2.8
Argon	Ar	18	2.8.8

Octet rule

The process of achieving a *stable* electron arrangement of 2 or 8 electrons in the outermost energy level is known as *duplet* e.g. in helium or *octet rule* respectively.
In order to achieve a more stable outermost electron arrangement of 2 or 8 electrons, electrons must be gained or lost in the outermost energy level. When an atom gains or loses an electron or electrons, the particle formed is charged either negatively (–ve) or positively (+ve). Such a particle is called an **ion**.

Formation of lithium and magnesium ions

Let us consider the formation of lithium ion.
A lithium atom has an electron arrangement of 2.1. It has only one electron in the outermost energy level. For it to be stable, it can either lose one electron to have an electron arrangement of 2 or gain seven electrons to attain an arrangement of 2.8.
More energy will be required for the atom to gain 7 electrons than to lose 1 electron Fig. 1.13. Therefore, it is easier for a lithium atom to lose 1 electron.

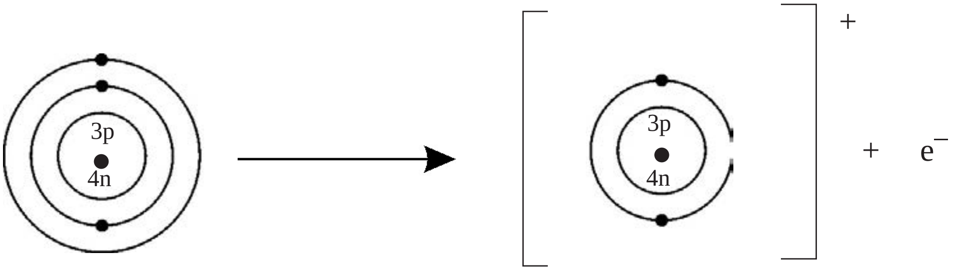


Fig. 1.13: Formation of lithium ion

Lithium atom	Lithium ion
Electron arrangement 2.1	*Electron arrangement* 2
Nucleus has - 3 protons (positively charged) - 4 neutrons (neutral)	*Nucleus has* - 3 protons (positively charged) - 4 neutrons (neutral)
Outside nucleus - 3 electrons (negatively charged)	*Outside nucleus* - 2 electrons (negatively charged)
Net charge +3-3=0	*Net charge* +3-2=+1
Symbol of the atom Li	*Symbol of the ion* Li^{+1} but represented as Li^{+}

Remember lithium has 3 protons in the nucleus which are positively charged and 4 neutrons which are neutral. The charge in the nucleus is +3. The electrons in the two energy levels are three and are negatively charged. This gives charge of -3 outside the nucleus. The net charge of the atom is
+3 – 3 = 0. That is why a lithium atom is neutral.

When a lithium atom loses one electron, the charge of the nucleus remains as +3 whereas outside the nucleus the charge reduces to –2. The net charge then becomes +3 –2 = +1. Lithium ion is then represented as Li^{+} but not Li^{+1}.

- Let us try again with magnesium

For magnesium, it is easier to lose two electrons than gain six electrons. Therefore a magnesium ion is formed by loss of two electrons.

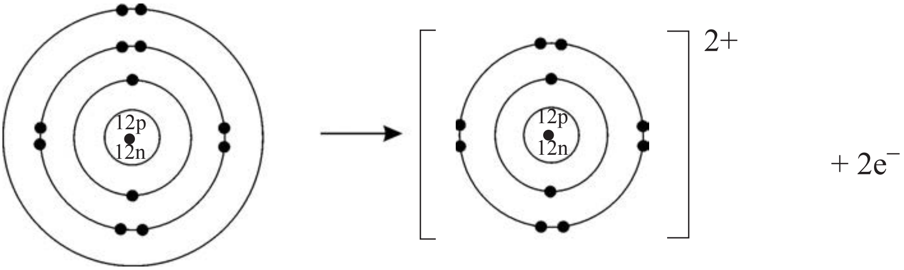


Fig. 1.14: Formation of magnesium ion, Mg^{2+}

Magnesium atom	Magnesium ion
Electron arrangement 2.8.2	*Electron arrangement* 2.8
Nucleus – 12 protons (positively charged) – 12 neutrons (neutral)	*Nucleus* – 12 protons (positively charged) – 12 neutrons (neutral)
Outside the nucleus – 12 electrons (negatively charged)	*Outside the nucleus* – 10 electrons (negatively charged)
Net charge: +12 - 12 = 0	*Net charge* +12 - 10 = +2
Symbol of the atom: Mg	*Symbol of the ion:* Mg^{2+} (but not Mg^{+2} or Mg^{++})

Li^+ and Mg^{2+} are called *ions.* Ions are electrically charged atoms. Li^+ and Mg^{2+} are positively charged ions. They are also called *cations.* Cations are formed when atoms lose one or more electrons to form positively charged ions.

- How are sodium and aluminium ions formed?
  Follow the same procedure as for lithium and magnesium to work this out.

Formation of fluoride and sulphide ions

A fluorine atom has the electron arrangement of 2.7. For fluorine atom to attain a stable electron arrangement of 2.8, it will either lose seven electrons or gain one electron respectively. Much more energy will be required for fluorine to lose seven electrons than to gain one electron. For this reason it is easier for fluorine to gain one electron than to lose the seven electrons. Therefore, a fluorine atom forms a **fluoride** ion by gaining one electron.

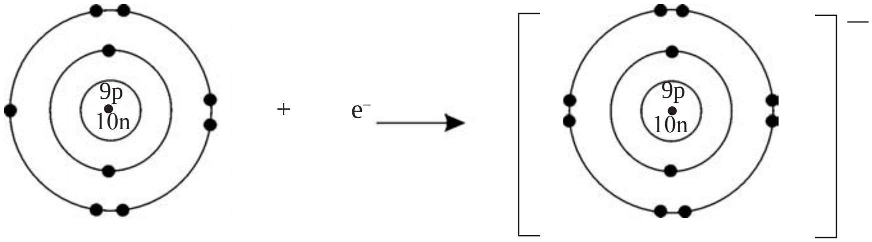


Fig. 1.15: Formation of fluoride ion, F^-

Flourine atom	**Flouride ion**
Electron arrangement 2.7	*Electron arrangement* 2.8
Nucleus – 9 protons (positively charged) – 10 neutrons (neutral)	*Nucleus* – 9 protons (positively charged) – 10 neutrons (neutral)
Outside the nucleus – 9 electrons (negatively charged)	*Outside the nucleus* – 10 electrons (negatively charged)
Net charge + 9 – 9 = 0	*Net charge* –10 + 9 = -1
Symbol of atom: F	*Symbol of the ion:* F^-

- Let us try with sulphur.

Sulphur atom has the electron arrangement of 2.8.6.
For sulphur, it is easier to gain two electrons than to lose six electrons.

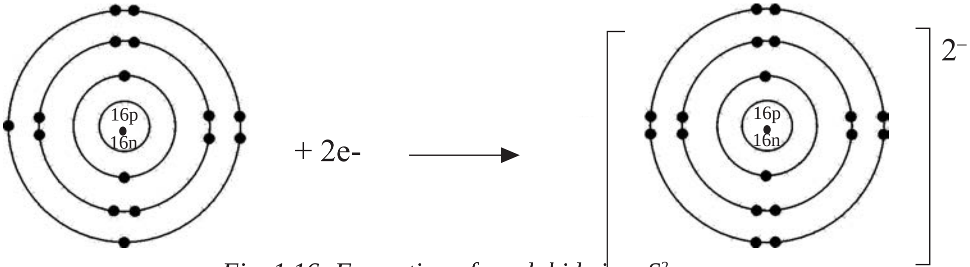


Fig. 1.16: Formation of a sulphide ion, S^{2-}

Sulphur atom	**Sulphide ion**
Electron arrangement 2.8.6	*Electron arrangement* 2.8.8
Nucleus –16 protons (positively charged) –16 neutrons (neutral)	*Nucleus* –16 protons (positively charged) –16 neutrons (neutral)
Outside the nucleus –16 electrons (negatively charged)	*Outside the nucleus* – 18 electrons (negatively charged)
Net charge +16-16 = 0	*Net charge* -18 +16 = $^{-}2$
Symbol of atom: S	*Symbol of the ion:* S^{2-}

F^{-} and S^{2-} are ions. They are electrically charged. F^{-} and S^{2-} are negatively charged ions. Negatively charged ions are called *anions*. They are formed when a neutral atom gains one or more electrons.

- How is a chloride ion formed? Follow the same procedure as for fluorine and sulphur to show the formation of a chloride ion.

Table 1.10 Electron arrangement of some ions

Atom	Symbol of atom	Electron arrangement of the atom	Symbol of ion	Electron arrangement of the ion
Lithium	Li	2.1	Li^{+}	2
Sodium	Na	2.8.1	Na^{+}	2.8
Fluorine	F	2.7	F^{-}	2.8
Chlorine	Cl	2.8.7	Cl^{-}	2.8.8
Aluminium	Al	2.8.3	Al^{3+}	2.8
Magnesium	Mg	2.8.2	Mg^{2+}	2.8
Sulphur	S	2.8.6	S^{2-}	2.8.8

Ionisation energy

We have seen from the structure of atoms that protons in the nucleus which are positively charged (+ve) attract electrons which are negatively charged (–ve) and located in the energy levels. Therefore, in order to remove electron, we must overcome this force of attraction. In other words we must supply energy to pull off the electron(s). The energy supplied is called *ionisation energy*. The unit for this energy is called *joule*(J) which can be converted to *kilojoules*(kJ).

Ionisation energy is defined as the energy required to remove electron(s) from an atom in gaseous state to produce an ion. The equation for the loss of an electron is represented as:

$$M(g) \rightarrow M^+(g) + e^- \quad \text{M-represents a metal atom}$$

If an atom loses 2 electrons the equation is as follows:

$$M(g) \rightarrow M^{2+}(g) + 2e^-$$

Note the charge on the ion is the same as the number of electrons lost. If an atom loses three electrons, then the charges will be 3+

$$M(g) \rightarrow M^{3+}(g) + 3e^-$$

Electron affinity

We have seen that non-metals gain electron(s) to gain stability as illustrated in Table 1.10. Remember that electrons are negatively charged. Therefore when an electron tries to enter the outermost energy level, it will be repelled by the electrons which are already there. So "force" is required to put the electrons into the energy level. This "force" is a form of energy. This energy is known as *electron affinity*. The ions formed are always negatively charged. For example:

$$Cl + e^- \rightarrow Cl^-$$
$$S + 2e^- \rightarrow S^{2-}$$

Oxidation numbers

We mentioned earlier, an atom will gain or lose electrons according to the number of electrons in the outermost energy level. In the process the atom becomes an ion. The *oxidation number* of an ion is simply the charge on the ion. + or – sign is written before the number unlike the charge. See Table 1.11

Table 1.11 Oxidation numbers

Ion	Oxidation number
H^+	+1
Na^+	+1
K^+	+1
Ca^{2+}	+2
Al^{3+}	+2
Cl^-	–1
F^-	–1
O^{2-}	–2

Why is the oxidation number of H, +1, yet it is a non-metal? Sometimes a hydrogen atom loses an electron like a metal. During other reactions it gains an electron to achieve the helium electron arrangement.

Table 1.12 Periodic table with some common oxidation numbers

I	II		III	IV	V	VI	VII	VIII
H +1								He 0
Li +1	Be +2	▨				O –2	F –1	Ne 0
Na +1	Mg +2	▨	Al +3			S –2	Cl –1	Ar 0
K +1	Ca +2	▨						

From the periodic table above there are similarities of the oxidation number of elements that belong to the same group. For example, Group I elements have oxidation of +1 whereas group seven elements have oxidation numbers of -1 etc.

Valencies

The number of electrons an atom requires to attain the stable **noble gas** electron arrangement is known as *valency* or combining power of the atom or group of atoms. A group of atoms which occur in compounds but cannot exist on its own is called a *radical*, for example sulphate, SO_4^{2-}, hydroxide, OH^-, etc.

The oxidation number has a negative or positive sign unlike valency which does not have. In metals the valency is just the number of electrons in the outermost energy level. For non-metals it is the difference between group 8 and group number of the elements, e.g. the valency of oxygen is 8 – 6 = 2, valency of phosphorous is 8 – 5 = 3.

Table 1.13 Valencies of the first twenty elements in the periodic table

Atomic Number	Element	Symbol	Electron arrangement	To gain stability of a noble gas atoms gain or lose electrons	Valency
1	Hydrogen	H	1	Not stable, can lose $1e^-$	1
2	Helium	He	2	Stable, cannot lose or gain	0
3	Lithium	Li	2.1	Not stable, loses $1e^-$	1
4	Beryllium	Be	2.2	Not stable, loses $2e^-$	2
5	Boron	B	2.3	Not stable, loses $3e^-$	3
6	Carbon	C	2.4	Not stable, loses or gains $4e^-$	4
7	Nitrogen	N	2.5	Not stable, easy to gain $3e^-$	3
8	Oxygen	O	2.6	Not stable, easy to gain $2e^-$	2
9	Fluorine	F	2.7	Not stable, easy to gain $1e^-$	1
10	Neon	Ne	2.8	Stable, cannot gain or lose electron	0
11	Sodium	Na	2.8.1	Not stable, loses $1e^-$	1
12	Magnesium	Mg	2.8.2	Not stable, loses $2e^-$	2
13	Aluminium	Al	2.8.3	Not stable, loses $3e^-$	3
14	Silicon	Si	2.8.4	Not stable, loses 4 or gains $4e^-$	4
15	Phosphorous	P	2.8.5	Not stable, gains $3e^-$	3 or 5
16	Sulphur	S	2.8.6	Not stable, gains $2e^-$	2
17	Chlorine	Cl	2.8.7	Not stable, loses $1e^-$	1
18	Argon	Ar	2.8.8	Stable (octet) cannot gain or lose electrons	0
19	Potassium	K	2.8.8.1	Not stable, loses $1e^-$	1
20	Calcium	Ca	2.8.8.2	Not stable, loses $2e^-$	2

Table 1.14 Valencies of the first twenty elements in the periodic table

I	II		III	IV	V	VI	VII	VIII
H 1								He 0
Li 1	Be 2	▨	B 3	C 4	N 3	O 2	F 1	Ne 0
Na 1	Mg 2	▨	Al 3	Si 4	P 3 or 5	S 2	Cl 1	Ar 0
K 1	Ca 2							

Note that valency corresponds to the group number for metals. For non-metals subtract group number from 8.

Table 1.15 Valencies of other common elements

Name of metal	Symbol	Valency
Zinc	Zn	2
Iron	Fe	2 or 3
Tin	Sn	2 or 4
Lead	Pb	2 or 4
Copper	Cu	1 or 2
Silver	Ag	1
Barium	Ba	2

Table 1.16 Valencies of common radicals

Valency 1		Valency 2		Valency 3	
Radical	**Formula**	**Radical**	**Formula**	**Radical**	**Formula**
Ammonium	NH_4^+	Carbonate	CO_3^{2-}	Phosphate	PO_4^{3-}
Hydroxide	OH^-	Sulphate	SO_4^{2-}		
Nitrate	NO_3^-	Sulphite	SO_3^{2-}		
Chloride	Cl^-				
Hydrogencarbonate	HCO_3^-				
Hydrogensulphate	HSO_4^-				

We already mentioned that valency is combining power of an element or radical. However some elements have varying valencies, for example, copper can have valency of 1 or 2. The valencies of these elements are indicated in Roman numbers in brackets after the name of the element when naming compounds. This is illustrated in Table 1.17.

Table 1.17 Valencies of elements in compounds

Compound	Element	Valency
Copper(I) oxide	Copper	1
Copper(II) oxide	Copper	2
Iron(II) sulphate	Iron	2
Iron(III) chloride	Iron	3
Sulphur(IV) oxide	Sulphur	4
Sulphur(VI) oxide	Sulphur	6
Carbon(IV) oxide	Carbon	4
Carbon(II) oxide	Carbon	2

Writing chemical formulae

Every subject has its own special language as you might have realized by now. For example, in Chemistry we have substances with names like sodium chloride. However, we have a short notation of writing names called *formula.* The formula of a compound shows the atoms present in the compound in their simplest ratio. The formula of sodium chloride is NaCl. The formula for water is H_2O. Why then do we write the number 2 in the formula for water whereas we do not in sodium chloride? These are some of the questions we shall answer in this section as we discuss how to write chemical formulae.

Activity 1.2 Writing chemical formulae using a game of cards

Procedure

1. Get pieces of plain paper.
2. For group I elements or cations, with valency 1, cut the pieces of paper into the following shape but make sure that they are of the same size. Indicate the symbol and the name of the element.

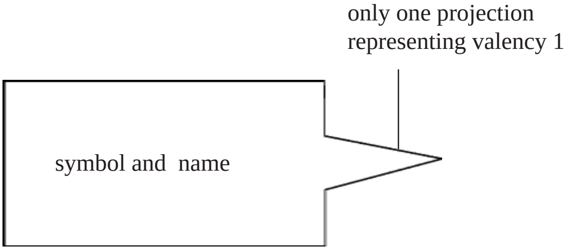


Fig. 1.17: Shape representing a cation with valency 1

3. For group II elements or cations with valency 2, cut the paper such that it has two projections as shown in the figure below.

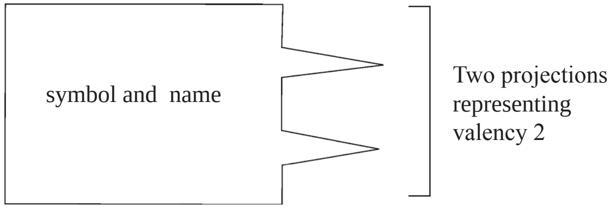


Fig. 1.18: Shape representing a cation with valency 2

4. For group III elements or cations with valency 3, your piece of paper will have three projections as shown in Fig. 1.19.

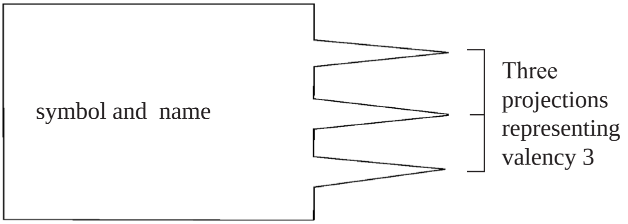


Fig. 1.19: Shape representing a cation with valency 3

5. For group VII elements or anions with valency 1, cut the piece of paper so that it has one groove as shown in Fig. 1.20

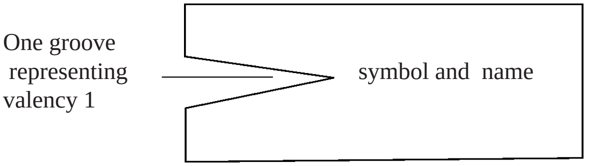


Fig. 1.20: Shape of anion with valency 1

6. Group VI elements or anions with valency 2, will be represented by pieces of paper with two grooves.

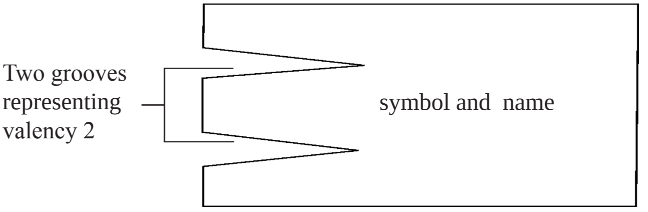


Fig. 1.21: Shape representing an anion with valency 2

7. For elements with valency 3, cut out three grooves.

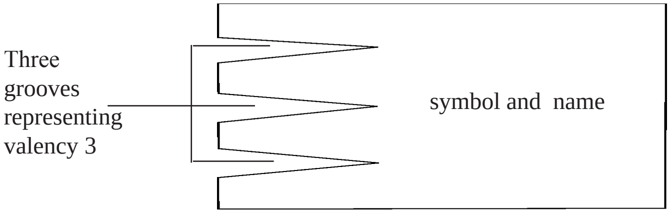


Fig. 1.22: Shape representing an anion with valency 3

Note: It is important to take care of the following point before you play this game.

- When writing the formula, metals are always written first. For example it is not correct to write ClNa or Br_2Ca even though it gives us the same information. The correct way is NaCl and $CaBr_2$ respectively.

Examples

(a) Write the chemical formula of sodium sulphate.

From the above activities, get the sodium and the sulphate cards as in step 1.

Step 1

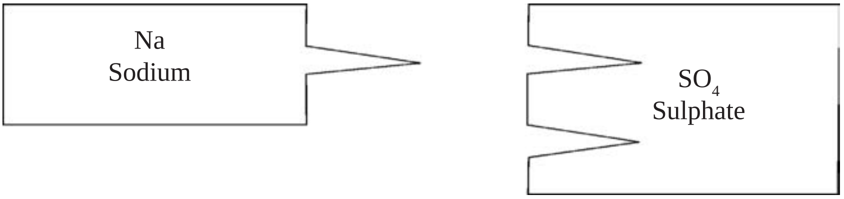


Fig. 1.23: Sodium and sulphate cards

Step 2

Fix the two cards together such that the protruding side of the sodium fits into the groove of the sulphate.

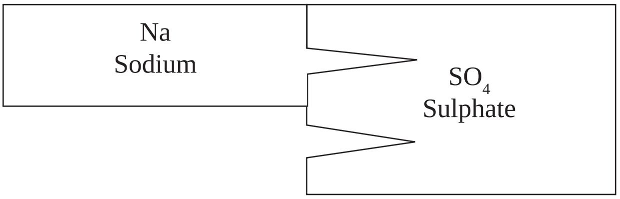


Fig. 1.24: Joined sodium and sulphate cards

Notice that one groove of the sulphate is still free. For it to be complete, you need another card of sodium.

Step 3

Fix one other card of sodium

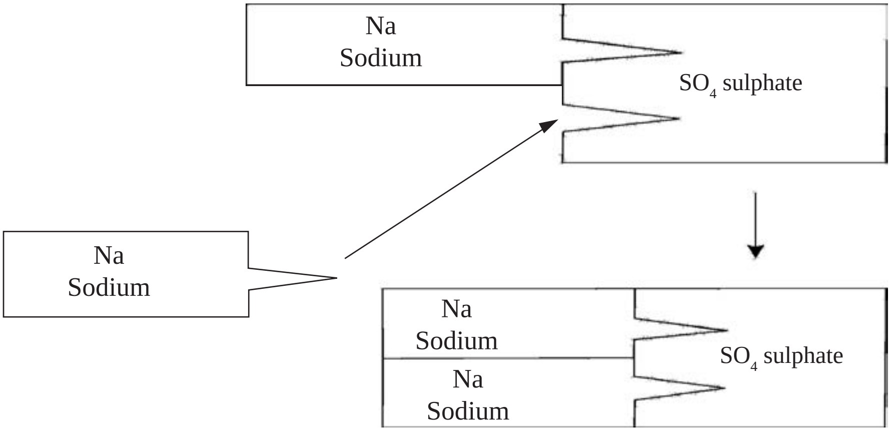


Fig. 1.25: Complete joined cards representing sodium sulphate

You can see that we need two cards of sodium and one card of sulphate. The chemical formula of sodium sulphate is therefore $Na_2SO_{4.}$

(b) Write the formula of potassium hydroxide.

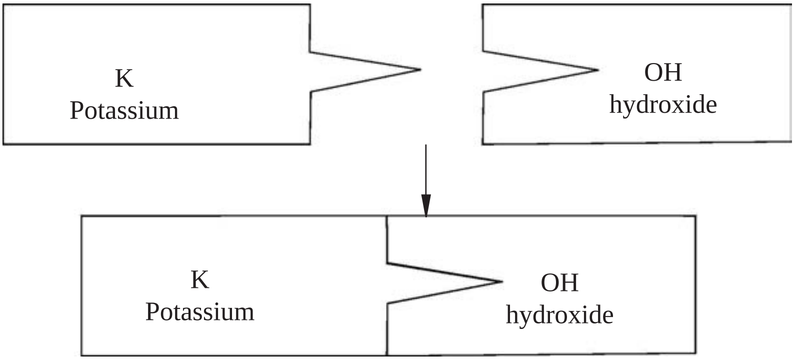


Fig. 1.26: Representation of potassium hydroxide with the cards

We need one potassium card and one hydroxide card. So the formula of potassium hydroxide is KOH.

Now let us use another simple method of writing chemical formulae. Always remember that to write a correct formula we must write down the;

(a) correct symbol of the element or radical

(b) correct valency of the symbol or radical.

Let us write the formula for sodium sulphate again following the steps below:

Step 1

Write the symbols of the elements and radical.

$Na \quad SO_4$

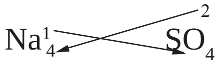
Step 2

Write the valencies of the element and radical above and to the right side of each.

$Na^1 \quad SO_4^2$

Step 3

Exchange the valencies by writing them below the symbols as shown by the arrows.



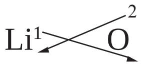
Step 4

Write the symbols close together

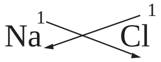
Na_2SO_4

Other examples

(a) Lithium oxide

Symbols	Li	O
Valencies	1	2
Formula		$\longrightarrow Li_2O$

(b) Sodium chloride

Symbols	Na	Cl
Valencies	1	1
Formula		$\longrightarrow NaCl$

(c) Calcium chloride

Symbols	Ca	Cl
Valencies	2	1
Formula	$Ca^{2} \times Cl^{1} \longrightarrow CaCl_2$	

(d) Ammonium sulphate

Symbols	NH_4	SO_4
Valencies	1	2
Formula	$NH_4^{1} \times SO^{2} \longrightarrow (NH_4)_2\ SO_4$	

(e) Magnesium hydroxide

Symbols	Mg	OH
Valencies	2	1
Formula	$Mg^{2} \times OH^{1} \longrightarrow Mg(OH)_2$	

Note: We write brackets for (d) and (e) because the radical consists of two different elements eg. OH contains oxygen atom and hydrogen atom; NH_4 consists of nitrogen and hydrogen atoms. But it is wrong to put bracket for (c) i.e. $Ca(Cl)_2$ as "C" and "l do not represent different elements.

(f) Write the formula of lead (IV) oxide.

Symbol	Pb	O
Valencies	4	2
Formula	$Pb^{4} \times O^{2}$	

Note: Before we bring the valencies down, we must divide the valencies by the common factor, 2.

PbO_2 is the formula

(g) Write the formula of iron(III) oxide

Here the valencies have no common number which can divide them and give us a whole number. Therefore, we just exchange the valencies as shown $\mathrm{Fe}^{3} \times \mathrm{O}^{2}$

The formula is $\mathrm{Fe}_2\mathrm{O}_3$.

- Use the card game for the above worked examples and to see whether you get the same formulae.

Exercises

Write the formulae for the following compounds

1. Calcium carbonate
2. Sodium phosphate
3. Ammonium phosphate
4. Lead(II) sulphate
5. Iron(II) sulphate
6. Iron(II) hydroxide
7. Lead nitrate
8. Manganese(IV) oxide
9. Copper(II) hydroxide
10. Iron(III) chloride

Writing simple balanced chemical equations

Chemical equations are short, clear and accurate descriptions of chemical reactions. A reaction process can be explained using an equation. For example, when oxygen reacts with magnesium ribbon, a white solid of magnesium oxide is formed. We have been using a word equation to describe such a reaction. Thus:

Magnesium + oxygen → Magnesium oxide

The (+) sign here is not used to mean addition, but in chemistry it is used to mean 'reacts with'. The → sign is used to indicate formation of a product. Using equal sign instead of an arrow, is wrong.

Names of starting substances like magnesium and oxygen in the above example are written on the left side of the arrow; these substances are called *reactants*. The new substances produced by the chemical reaction are called *products* and are written on the right side of the arrow.

Magnesium + oxygen → Magnesium oxide
(Reactants) **(Product)**

Chemical equations have various notations that indicate the physical states of the reactants and products. These notations are very important. Infact, whenever one writes an equation and misses to write them, the equation is not complete. These notations are as follows:

Table 1.20 State symbols

Physical state	**Representation of state**	**Description**
Solid	(s)	a solid can be a precipitate, suspension, etc
Liquid	(l)	a pure liquid like water, paraffin, etc
Aqueous	(aq)	a solute or liquid dissolved in water
Gas	(g)	a gas or vapour

Balancing chemical equations

In order for an equation to describe a reaction accurately, the equation must be balanced. Chemical reactions should always follow the law of conservation of mass so that the total mass of reactants must be equal to the total mass of products.

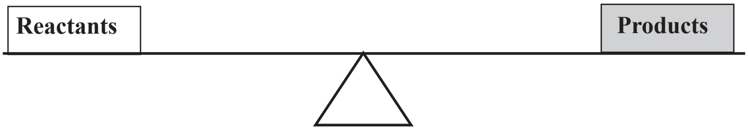


Fig. 1.27: Conservation of mass in a balanced chemical reaction

Let us consider our previous example of magnesium ribbon reacting with oxygen. We can write an initial equation containing the formulae of reactants and products.

$$Mg \quad + \quad O_2 \quad \rightarrow \quad MgO$$

This equation is not balanced. Count the number of atoms of the reactants and the products.

REACTANTS	PRODUCT
1 Magnesium atom 2 Oxygen atoms	1 Magnesium atom 1 Oxygen atom

The left side (reactants) has two oxygen atoms but the right side (product) has only one oxygen atom.

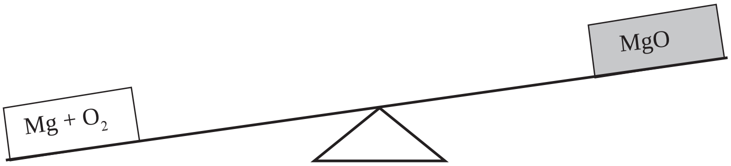


Fig. 1.28: Unbalanced equation, more oxygen atoms on the left

An equation can be balanced using a number of rules. Let us follow these rules to balance the above equation.

Rule number 1

Write the equation using correct formulae for the reactants and products

$$\text{Magnesium} + \text{oxygen} \rightarrow \text{Magnesium oxide}$$
$$Mg + O_2 \rightarrow MgO$$

Rule number 2

Count the number of atoms of each element in the reactants and in the products. Check whether they are equal as in Table 1.21.

Table 1.21 Balancing number of atoms for each element in the reactants and product

Atoms	**Reactants**	**Product**
MG	1	1
O	2	1

We notice that the oxygen atoms are not equal. We have 2 atoms on the left and only one on the right.

Rule number 3

To make oxygen atoms equal, balance the equation by writing numbers *in front of the formula*. Remember that 1 is assumed to be there already. Therefore start by inserting 2. If this does not balance, go to 3, 4 until the equation is balanced. Usually we do not go to very big numbers.

$$2Mg + O_2 \rightarrow 2MgO$$

The number 2 now balances the equation. When you have $2O_2$ it means two oxygen molecules. The number in front of a formula means, everything following is multiplied by that number.

For example: 2Mg means 2 × Mg that is why we have 2 Mg atoms on the left and 2 Mg on the right.

This equation $2Mg + O_2 \rightarrow 2MgO$ means 2 atoms of Mg react with 1 molecule of oxygen (containing 2 atoms) to form 2 molecules of magnesium oxide.

Step number 4

Count again the number of atoms of each element on the reactants and product sides. Note that all atoms are balanced as illustrated in Table 1.22.

Table 1.22: Balancing number of atoms for each element in the reactants and products

Atoms	Reactants	Product
Mg	2	2
O	2	2

Step number 5

Insert the correct state symbols for each substance.

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

Let us practice the above steps by writing equations for the following reactions.

Example

Zinc granules react with dilute hydrochloric acid to produce aqueous zinc chloride and hydrogen gas.

Step 1

Write an initial equation using correct chemical formulae of reactants and products.

$$\begin{array}{ccccccc} \text{Zinc} & + & \text{hydrochloric acid} & \rightarrow & \text{zinc chloride} & + & \text{hydrogen} \\ Zn & + & HCl & \rightarrow & ZnCl_2 & + & H_2 \end{array}$$

Step 2

Count atoms

Table 1.23 Balancing the number of atoms for each element in the reactants and products

Atoms	Reactants	Products
ZN	1	1
H	1	2
CL	1	2

Note: If you look at the reactants and products, zinc is already balanced but hydrogen and chlorine are not. We need to have 2 hydrogen atoms and 2 chlorine atoms on the reactants side (left side).

Step 3

Insert a number infront of the formula to balance the equation.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Note: A number of 2 has been put infront of HCl at the left side to make sure that the balance is achieved. Let us confirm in the next step.

Step 4

Count the atoms again.

Table 1.24 Balancing the number of atoms for each element in the reactants and products

Atoms	Reactants	Products
ZN	1	1
H	2	2
CL	2	2

Step 5

Insert correct state symbols for each substance.

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Solid zinc reacts with dilute hydrochloric acid to form aqueous zinc chloride and hydrogen gas.

Exercises

Write the correct balanced equations for the following reactions.

1. Reaction of copper(II) oxide with hydrogen to produce copper and water.
2. Reaction of solid calcium carbonate with dilute hydrochloric acid to produce aqueous calcium chloride, carbon(IV) oxide and water.
3. Reaction of solid copper(II) oxide with dilute sulphuric acid to produce aqueous copper(II) sulphate and water.
4. Reaction of aqueous aluminium chloride with aqueous sodium hydroxide to produce aluminum hydroxide and aqueous sodium chloride.
5. Reaction of solid copper carbonate with dilute hydrochloric acid to produce aqueous copper(II) chloride, carbon(IV) oxide and water.

Projects

- Construction of cards for chemical formulae.
- Construction of models of the periodic table.

Summary

1. Atoms are the smallest particles of an element.
2. The central part of the atom is the nucleus.
3. The nucleus contains smaller particles called protons and neutrons.
4. The number of protons = the number of electrons.
5. Protons are positively charged while electrons are negatively charged. Neutrons have no charge.
6. Electrons move around the nucleus along energy levels.
7. The atomic number (Z) refers to the number of protons in the nucleus of an atom.
8. The mass number (A) is the sum of protons and neutrons.
9. Number of neutrons = mass number – atomic number.
10. The energy levels have different energies and are at different distances from the nucleus of the atom.
11. The energy level next to the nucleus has the least amount of energy and can accomodate a maximum of two electrons.
12. The other energy levels if it is the outermost, can accomodate a maximum number of eight electrons.
13. The electrons in the outermost energy level are very important in determining the chemistry of the element.
14. Isotopes are atoms of the **same** element.
15. Isotopes have the same number of protons and electrons in each atom, but have different numbers of neutrons in their nuclei.
16. Since isotopes have same number of electrons, they have same or nearly same chemical properties.

17. Nearly all the mass of an atom is concentrated in the nucleus. The mass of a proton and neutron are approximately equal. The mass of an electron is negligible.
18. The relative atomic mass of an element is the number of times the average mass of one of its atoms is greater than $\frac{1}{12}$ of the mass of one atom of carbon – 12.
19. Some relative atomic masses are not whole numbers because the relative abundance of the various isotopes is not equal.
20. The elements in the periodic table are arranged in order of increasing atomic number.
21. Elements in the same group have the same number of electrons in their outermost energy level.
22. Elements in the same period have the electrons added to the same energy level.
23. The valency is the combining power of an atom or group of atoms (radicals).

Revision Exercise 1

1. Here is a symbol of lithium atom:

$$^{7}_{3}\mathrm{Li}$$

(a) What is the name given to the superscript (top number)?
(b) What is the name given to the subscript (bottom number)?
(c) In the lithium atom shown above, calculate the number of protons, neutrons and electrons in the atom.

2. Select the pair which represents two atoms with the same number of neutrons.

A. $^{12}_{6}\mathrm{C}$ and $^{24}_{12}\mathrm{Mg}$

B. $^{18}_{8}\mathrm{O}$ and $^{19}_{9}\mathrm{F}$

C. $^{23}_{11}\mathrm{Na}$ and $^{20}_{10}\mathrm{Ne}$

3. Explain briefly why some elements have relative atomic masses which are not whole numbers.

4. An atom is the smallest particle of an element. Name the sub-atomic particles found in an atom and state where they are found.

5. Explain the following terms:

(i) Atomic number
(ii) Mass number
(iii) Isotopes

6. The following symbols refer to isotopes of oxygen.
(i) $^{16}_{8}O$

(ii) $^{18}_{8}O$

What is the number of protons and neutrons in the nucleus of each isotope?

7. A class was asked to select from the list below the elements whose electron arrangements were impossible to write down.
Which elements did the class select?
Explain your answer.

$^{23}_{11}Na$; ^{10}B; $_{6}C$; $^{12}_{6}C$; ^{25}Mg; $^{26}_{12}Mg$

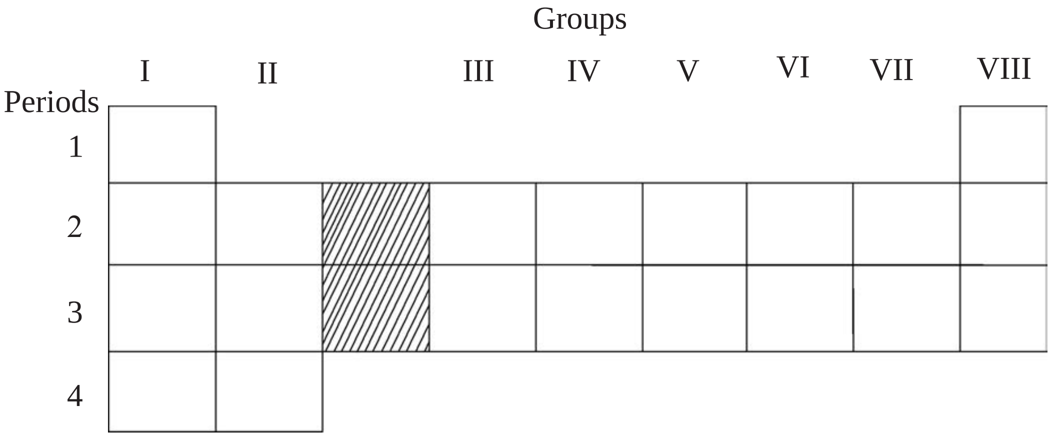
8. Write down and complete the following table.

Atom	**Mass number**	**Number of protons**	**Number of neutrons**	**Number of electrons**
F	19			9
Na			12	11
Ne	20	10		
C	12	6		

9. An atom of element X can be represented by the symbol $^{16}_{6}X$.
(a) State the:
(i) number of electrons of atom X.
(ii) mass number of X
(iii) atomic number of X.

(b) Element X contains 90% of $^{16}_{6}X$ and 10% of $^{18}_{6}X$. Calculate the relative atomic mass of X.

10. Put the following letters A to H in the correct places in the periodic table above to fit these descriptions.



A. An element with 7 protons.
B. It belongs to period three and has 4 electrons in the outermost energy level.
C. An element with oxidation number +3 and with 13 protons.
D. An element that forms its ion by loss of one electron and belongs to period four.
E. An element that can be placed in either group I or group VII.
F. An element with the highest number of energy levels.
G. An element with electronic arrangement of 2.8.8.
H. An element with valency of 3. It gains electrons to form ions and belongs to period three.

UNIT 2: Chemical families; patterns in properties

2.1 Introduction

The physical and chemical properties of elements and their compounds are highly dependent upon the number of electrons that are found in their outermost energy level. When elements are arranged in the periodic table in order of increasing atomic number, a regular change in the outermost electron arrangement is observed. In turn a regular variation of properties is also observed. Thus, as we move across a period certain repeated patterns can be observed. These lead us to make predictions and intelligent guesses about the unknown properties of elements and their compounds. In addition, elements that are in the same group have the same number of electrons in the outermost energy level and have similar chemical properties.

These facts lead to the observation of gradual changes in the physical and chemical properties of the elements. In order to illustrate these changes we will consider a number of trends and patterns observed in the periodic table.

2.2 Group I elements - Alkali metals

Group 1 elements are metals. They are called *alkali metals* and are the most reactive group of metals. They are found on the left hand side of the periodic table as illustrated in Fig 2.1.

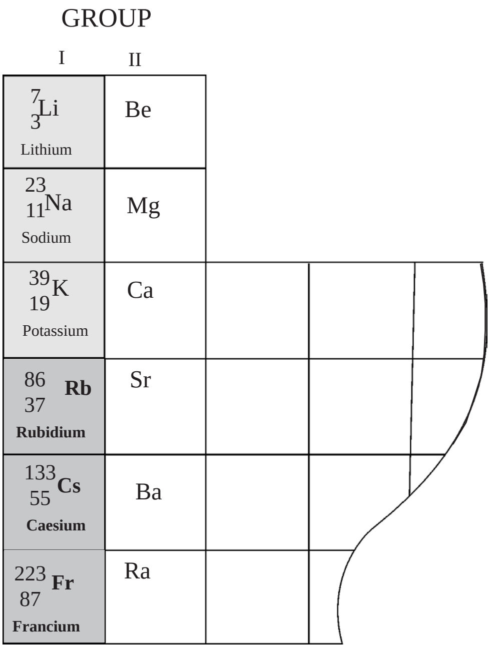


Fig 2.1 Alkali metals

Note: At this level we are going to study only lithium, sodium and potassium.

2.3 Trends in physical properties of alkali metals

Experiment 2.1

To investigate the physical appearance of alkali metals.

Apparatus and chemicals

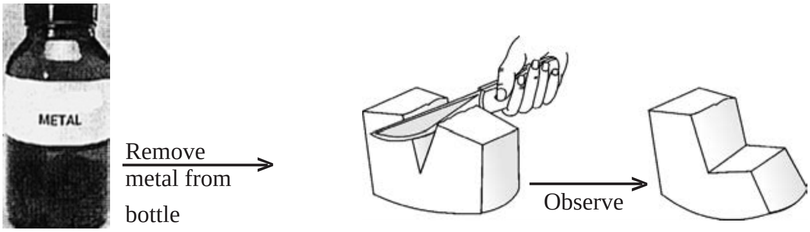
- tile (ceramic)
- filter paper
- knife
- pair of tongs
- lithium
- sodium
- potassium

Procedure

1. Using a pair of tongs, remove a small piece of lithium from the bottle.
2. Place it on the tile.
   - Describe the appearance of the lithium from the bottle?
3. Cut the piece of lithium into two with a knife to expose the inside.
   - What is the colour of the freshly cut part of lithium.
4. Repeat the above procedure with sodium and potassium and answer the same questions.
5. Record your observations in a table like the one shown below.
   - What do you conclude about the appearance of the group I element?
   - Why does the appearance of the metal change after a short while?

Table 2.1 Physical properties of alkali metals

	Appearance of metal from bottle	**Appearance of freshly cut metal**
Lithium		
Sodium		
Potassium		



(a) Metal in bottle (b) Cut metal (c) Exposed inner part of metal

Fig. 2.2 Physical appearance of alkali metals

Table 2.2 Physical properties of alkali metals

Element	Symbol	Physical state	Physical appearance	Electron arrangement	Melting point	Boiling point	Thermal conductivity	Electrical conductivity
Lithium	Li	Solid	Silvery metal	2.1	181°C	134°C	Good conductor of heat	Good conductor of electricity
Sodium	Na	Solid	Silvery metal	2.8.1	98°C	883°C	Good conductor of heat	Good conductor of electricity
Potassium	K	Solid	Silvery metal	2.8.8.1	63°C	759°C	Good conductor of heat	Good conductor of electricity

- State and explain the trends in physical properties of alkali metals.

We saw in the previous section that alkali metals have one electron in the outermost energy level. Because of this, alkali metals have very similar physical and chemical properties. They also show characteristic trends in these properties as one goes down the group.

Typical metal properties

Group I elements show typical metal properties such as:

- good conductivity of heat
- good conductivity of electricity
- high boiling points
- shiny surface when freshly cut.

Other physical properties of alkali metals

- Low melting points
- Very low density (they float on water)
- Very soft and can be cut with an ordinary knife easily.

Important trends down the group.

- The melting points and boiling points generally decrease.
- The density generally increases.

Atomic radius

The atomic radius is the distance between the nucleus and the outermost energy level in an atom. It is sometimes called atomic size.

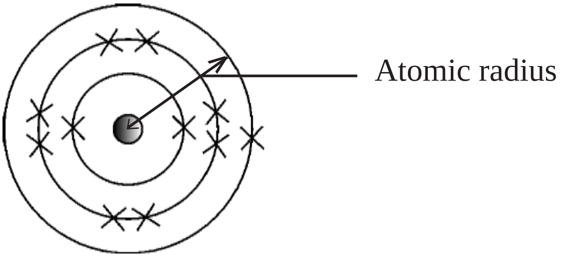


Fig. 2.3 Atomic radius of sodium

Of the three group 1 elements, namely, lithium, sodium and potassium, lithium has the smallest atomic radius followed by sodium. Potassium has the largest atomic radius. The atomic radius increases as we go down the group as illustrated in Fig. 2.4.

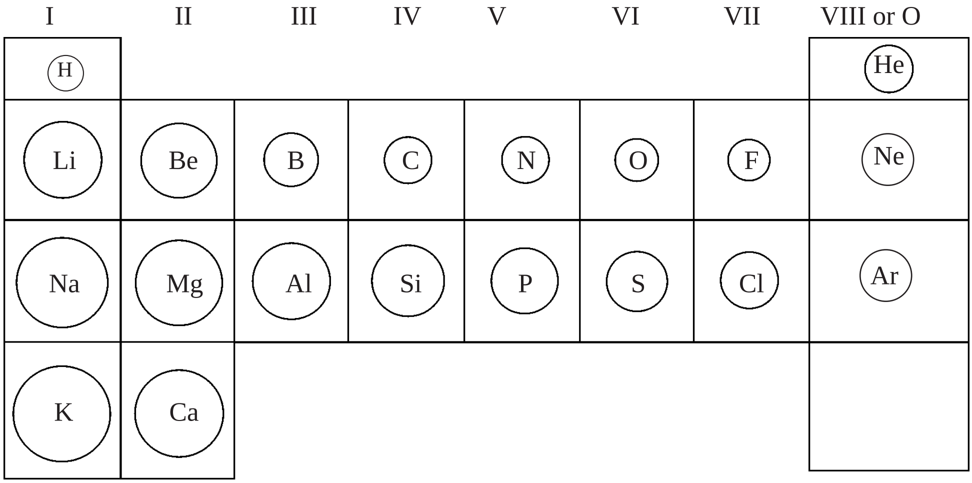


Fig. 2.4 Periodic variation of atomic radii of the 1st twenty elements of the periodic table.

From Fig 2.4, we notice that the atomic radius generally increases down a group. Why do you think this happens? Before we answer this question let us first look at Fig. 2.5.

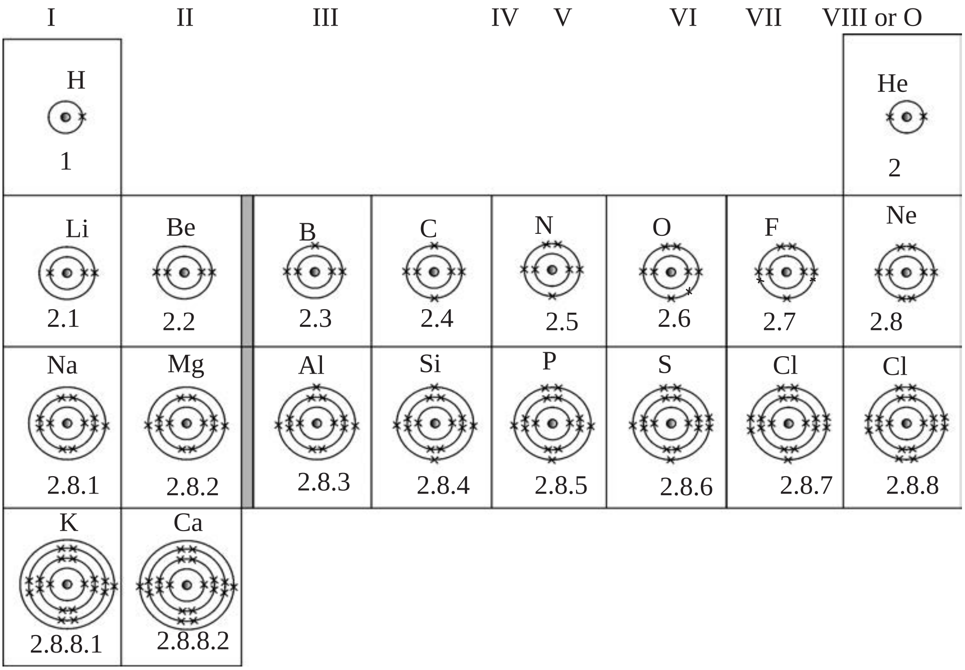


Fig. 2.5 Periodic variation of electron arrangement

The atomic radius increases down the group. This is because as we go down the group, the number of energy levels increase. Electrons in these energy levels are far from the nucleus. Although the nuclear charge increases due to the increase in the number of protons down the group, the electrons in the inner energy levels shield those in the outermost electrons from full attraction by the nucleus.

Ionic radius

Ionic radius is the distance between the nucleus and the outermost energy level in an ion.

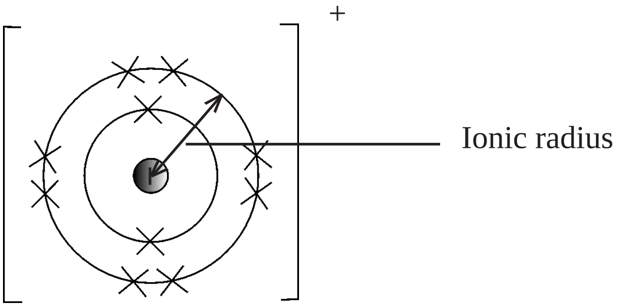


Fig. 2.6 Ionic radius of Na^+ ion

Generally, the ionic radii increase as we go down the group.

- Why does this happen?

Down the group, the number of energy levels increase which increase the ionic radius.

Why is ionic radius smaller than atomic radius for metals?

Let us take sodium as our example to explain this.

Na (atom) 2 .8.1

Na^+ (ion) 2.8

A sodium atom has 11 protons in the nucleus and 11 electrons in the energy levels. In other words each proton has an electron to attract. When sodium loses one electron to form an ion, the 11 protons in nucleus remain but in the energy levels we have 10 electrons. This means one proton in the nucleus has no partner to attract. Its attraction force goes to the 10 electrons. The ionic radius is smaller than the atomic radius because 11 protons are attracting 10 electrons.

The ionic radii of group I elements are smaller than atomic radii of the same element. This is because:

- ion formation involves loss of one energy level.
- there are fewer electrons to be attracted by protons.
- the outermost electrons are attracted more strongly to the nucleus.

Ionisation energy

Ionisation energy is a measure of how difficult it is to remove an electron from an atom in gaseous state. The ionisation energy is that energy that must be absorbed to remove the outermost electron from an atom.

$$M(g) \rightarrow M^+(g) + e^-$$

The higher the ionisation energy the more difficult it is to remove an electron from the atom.

The 1st ionisation energy for group I elements is as follows; Lithium 520kJ, sodium 496kJ and Potassium 419kJ. Notice that the 1st ionisation energy decreases as you move down the group in the periodic table. This means that it is easier to remove an electron from the atom down in Group I. Compare the atomic size (atomic radius) and ionisation energy of Group I elements. The smaller the atom the higher the ionisation energy. Electrons are more attracted to the nucleus in a smaller atom. It becomes more difficult to remove the electron(s) from such atoms. Potassium has the lowest ionisation energy because the outermost electron is far away from the nucleus and therefore is relatively easier to remove.

2.4 Trends in chemical properties of alkali metals

Experiment 2.2

To investigate reaction of alkali metals with air

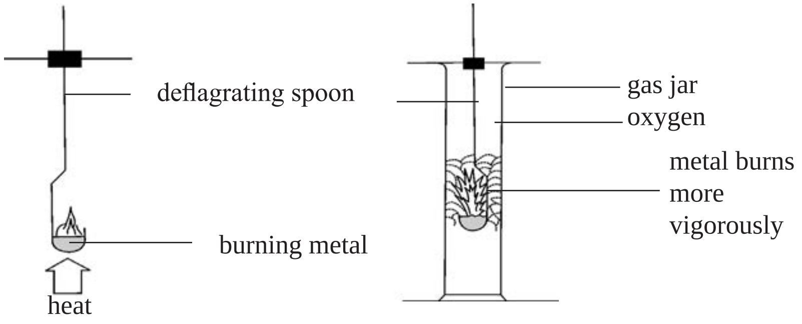
Apparatus and chemicals

- bunsen burner

- pair of tongs
- knife
- tile (ceramic)
- filter paper
- deflagrating spoon
- lithium
- sodium
- potassium

Procedure

1. Place all the requirements on the demonstration bench.
2. With a pair of tongs, remove a small piece of lithium from the bottle.
3. Place it on the ceramic tile covered with filter paper. Remove most of the oil with filter paper.
   Caution: Do not touch the metal with your hands.
4. Cut a small piece of lithium (about 2-3 cubic mm or that of a rice grain) and return the rest into the bottle.
5. Transfer the small piece of lithium into a deflagrating spoon.
6. Heat the lithium metal as illustrated by Fig. 2.7 below.
7. As lithium begins to burn, remove the spoon from the flame and observe the reaction.
8. Repeat the same procedure for sodium and potassium.
9. Record observations in your notebook.
   - What do you conclude about the reactivity of these metals with air (lithium, sodium and potassium)
   - Ali had an element X which he suspected to be in group 1. What would he expect to observe when burning X metal is inserted in a gas jar oxygen? Write a balanced equation for the reaction.



(a) burning metal in air (b) burning metal in oxygen

Fig. 2.7 Reaction of alkali metals in air and oxygen

Reaction of alkali metals with air

When alkali metals are freshly cut, they look shiny and silvery but they tarnish immediately when exposed to air. Group 1 elements react with air to form oxides and peroxides depending on the amount of oxygen available. Lithium reacts with oxygen to form lithium oxide which is white in colour.

Lithium + oxygen → Lithium oxide
$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$
(silvery) (white)

Sodium and potassium react with oxygen in air to form a mixture of white oxide and peroxide.

Sodium + oxygen → Sodium oxide
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
(silvery) (white)

Sodium + excess oxygen → Sodium peroxide
$2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$
(silvery) (yellow-white)

Potassium + oxygen → Potassium oxide
$4K(s) + O_2(g) \rightarrow 2K_2O(s)$
(silvery) (white)

Potassium + excess oxygen → Potassium peroxide
$2K(s) + O_2(g) \rightarrow K_2O_2(s)$
(silvery) (white)

Group I elements also burn with characteristic colours in a Bunsen flame. Lithium burns with a *scarlet flame;* sodium burns with a *yellow flame*; and potassium burns with a *lilac* flame.

Reaction of alkali metals with water

Experiment 2.3

To investigate reaction of alkali metals with water

Apparatus and chemicals

- glass trough
- pair of tongs
- tile (ceramic)
- knife

- filter paper
- lithium
- sodium
- potassium
- litmus solution or universal indicator solution
- water

Procedure

1. Half fill the trough on the demonstration bench with water.
2. Add three drops of universal indicator solution. You can also test the water with red and blue litmus paper instead of universal indicator solution or litmus solution.

- What colour change do you observe? Indicate the pH of the solution as neutral, alkaline or acidic.

3. With a pair of tongs, remove a small piece of lithium from the bottle and place it on the ceramic tile.
4. Cut a small piece of lithium (the size of a rice grain) and return the rest to the bottle.
5. Drop the small piece of lithium into the glass trough with water using a pair of tongs. What do you observe?
6. In case you used litmus paper, repeat the test with red and blue litmus paper. What do you observe?
    - Repeat the same procedure for sodium and potassium.
7. Record your observations in your notebook.
    - Arrange the metals in order of how reactive they are with water
    - State which metal is least reactive. Give the reason.

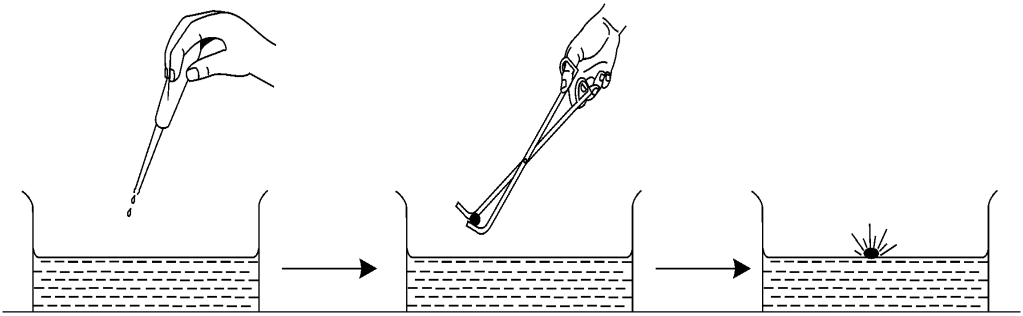


Fig. 2.8 Reaction of alkali metals with water

(i) Lithium

Lithium reacts slowly with water. It does not melt. The reaction produces heat. The heat produced is not sufficient to melt the lithium metal. Hydrogen gas and lithium hydroxide are formed when lithium reacts with water. Since lithium hydroxide is an alkaline solution, the indictor will change colour to show the presence of an alkaline solution.

Lithium + water → lithium hydroxide + hydrogen
$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

(ii) Sodium

Sodium reacts more *vigorously* than lithium. Immediately it is dropped in water, it moves around the surface of the water. The water changes colour depending on the indicator solution used. When the sodium is placed on top of a floating piece of filter paper, it catches fire. Sodium reacts vigorously with water to form hydrogen and sodium hydroxide which is an alkaline solution. Sodium placed on the filter paper ignites because the heat produced is sufficient to ignite the hydrogen gas liberated.

Sodium + water → sodium hydroxide + hydrogen
$2Na(s) + 2H_2O(1) \rightarrow 2NaOH(aq) + H_2(g)$

(iii) Potassium

Potassium reacts *explosively* with water. Immediately it is dropped in water it ignites spontaneously. The heat produced is so much that it ignites the hydrogen produced. Potassium hydroxide solution, which is an alkaline solution, is also formed.

Potassium + water → potassium hydroxide + hydrogen
$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

With all the three metals, depending on the indicator used the change of colour of the solution will be as follows:

- Litmus paper: red turns blue.
- Litmus solution turns blue.
- Universal indicator solution turns purple.
- Methyl orange: orange to yellow.
- Phenolphthalein colourless to pink.

Note

Group 1 elements are stored in paraffin because they can react with oxygen in the atmosphere.

Experiment 2.4

To investigate the reaction of alkali metals with chlorine.

Caution: (i) Do not attempt this reaction with potassium.
(ii) Reaction with chlorine should be done in a fume cupboard or an open space.

Apparatus and chemicals

- deflagrating spoon
- Bunsen burner
- tile (ceramic)
- knife
- pair of tongs
- filter paper
- two jars of chlorine
- lithium
- sodium

Procedure

1. With a pair of tongs, remove a small piece of lithium from the bottle and place it on the ceramic tile covered with a filter paper to remove much of the oil.
2. With a knife, cut a small piece of lithium (size of a rice grain) and return the rest to the bottle.
3. Transfer the small piece of lithium into a deflagrating spoon.
4. Hold the spoon directly into the Bunsen flame with your hand far away towards the flame. *Protect your eyes too.*
5. As soon as the lithium catches fire, lower it quickly into the gas jar of chlorine.
   What do you observe?
6. Repeat this procedure with sodium. *Do not attempt with potassium.*
7. Record your observations in a Table 2.3.Which metal react with chlorine most readily. Explain.

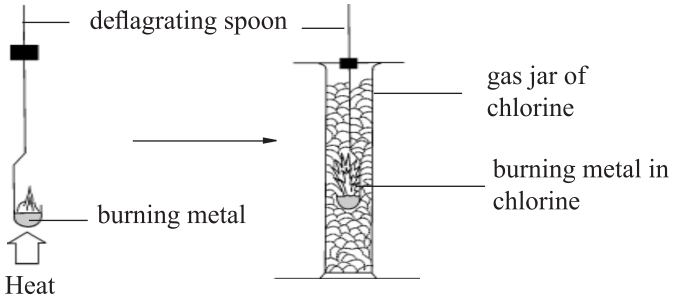


Fig 2.9 Reaction of alkali metals with chlorine

Table 2.3 Reaction of alkali metals with chlorine

Element	Observations (include balanced chemical equations)
Lithium	
Sodium	

Alkali metals react readily with chlorine to form chlorides. Lithium forms a white solid of lithium chloride and sodium forms a white solid of sodium chloride.

Lithium + chlorine → lithium chloride

$2Li(s) + Cl_2(g) \rightarrow 2LiCl(s)$

Sodium + chlorine → sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

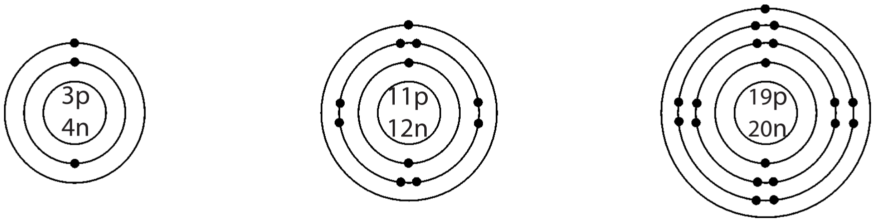
Note: The experiment with potassium is explosive.
Potassium would react explosively with chlorine to form a white solid of potassium chloride.

Potassium + chlorine → potassium chloride

$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

2.5 Similarities of ions and formulae of compounds formed by alkali metals

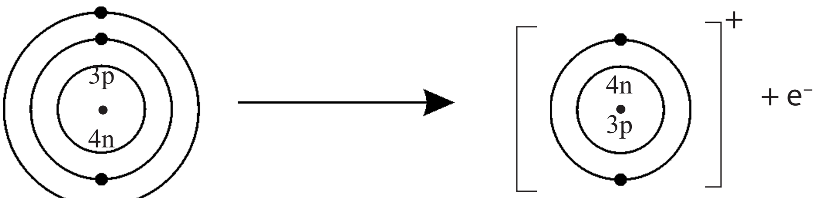
As a reminder, group I elements or alkali metals have one electron in their outermost energy levels.



a) Lithium 2.1 b) Sodium 2.8.1 c) Potassium 2.8.8.1

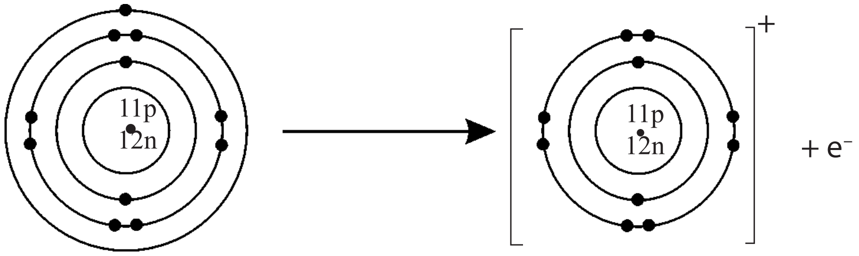
Fig 2.10 Electronic structures of alkali metals

It is easier to lose one electron than gain seven electrons to attain the octet rule of 8 electrons in the last energy level. Therefore, Group 1 elements form their ions by losing one electron. Lithium attain the duplet rule of 2 electrons in the first energy level.



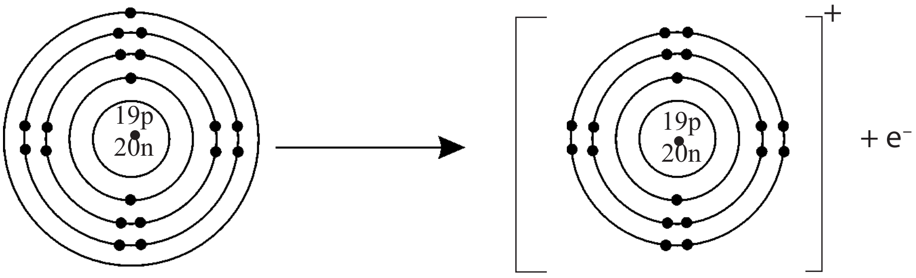
Lithium atom, Li 2.1

Lithium ion, Li$^+$ 2



Sodium atom, Na 2.8.1

Sodium ion, Na$^+$ 2.8



Potassium atom, K 2.8.8.1

Potassium ion, K$^+$ 2.8.8

Fig. 2.11, 2.12 and 2.13: Formation of lithium, sodium and potassium ions

All elements in Group 1 form ions of 1+ charge. They all have a valency of 1. See Table 2.4.

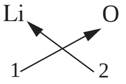
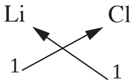
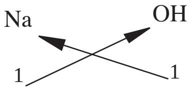
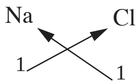
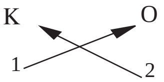
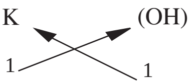
Table 2.4 Ions and valencies of alkali metals

Element	Symbol	Ion	Valency
Lithium	Li	Li$^+$	1
Sodium	Na	Na$^+$	1
Potassuim	K	K$^+$	1

Similarities in chemical formulae of compounds of alkali metals

- What are the similarities of formulae of compounds of alkali metals? Alkali metals will form compounds with the following chemical formulae.

Table 2.5 Formulae of alkali metal compounds

	Oxides (O^{2-})	**Hydroxides (OH^{-})**	**Chloride (Cl^{-})**
Lithium symbol Valency	Lithium Oxide  Li O 1 2 Li_2 O_1 Formula Li_2O		Lithium chloride  Li Cl 1 1 Li_1 Cl_1 Formula LiCl
Sodium		Sodium Hydroxide  Na OH 1 1 Na_1 $(OH)_1$ Formula NaOH	Sodium Chloride  Na Cl 1 1 Na_1 Cl_1 Formula NaCl
Potassium	Potassium Oxide  K O 1 2 K_2 O_1 Formula K_2O	Potassium Hydroxide  K (OH) 1 1 K_1 $(OH)_1$ Formula KOH	

- In Table 2.5 follow the same procedure to fill in the empty spaces. From the above examples, it is clear that Group I elements have similarities in their ions and formulae.
- List down some similarities of ions and formulae of compounds in Group I elements.

2.6 Uses of alkali metals

Some uses of sodium

1. Sodium metal is used in the preparation of tetraethyl lead, an important antiknock reagent in petrol. However, tetraethyl lead is being phased out in many countries because of lead pollution problems.
2. Sodium metal is used in the preparation of titanium metal from titanium chloride.
3. Sodium vapour is used in lamps for street lighting.
4. The sodium compound, NaCl is widely used to season or add taste to food. It is not healthy to eat too much salt.

5. Sodium carbonate is used in the manufacture of glass, detergents and for softening hard water.
6. Sodium hydroxide and sodium chloride are used in the manufacture of soaps and detergents.

2.7 Group II elements:- Alkaline-earth metals

The Group II elements are known collectively as *alkaline-earth metals*. They are found at the left hand side of the periodic table just after Group I as illustrated in Fig. 2.14.

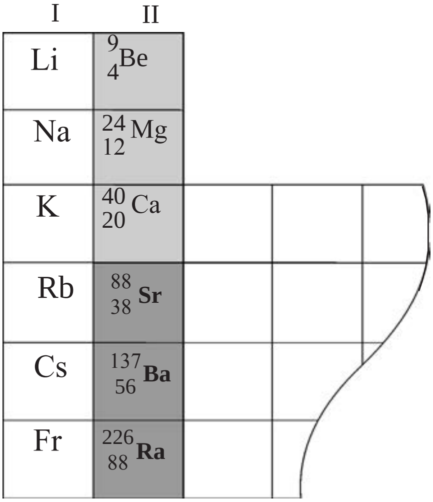


Fig 2.14 Alkaline-earth metals

Note: At this level, only beryllium, magnesium and calcium will be discussed.

2.8 Trends in physical properties of alkaline-earth metals

- State and explain trends in physical properties of alkaline earth metals.

Physical properties

Table 2.6 Physical properties of alkaline-earth metals

Element	Symbol	Electron arrangement	Melting point	Boiling point	1st Ionization energy	Physical appearance	Thermal conductivity	Electrical conductivity
Beryllium	Be	2.2	1278°C	2970°C	899	Silvery metal	Good conductor of heat	Good conductor of electricity
Magnesium	Mg	2.8.2	669°C	1107°C	738	Silvery metal	Good conductor of heat	Good conductor of electricity
Calcium	Ca	2.8.8.2	839°C	1484°C	590	Silvery metal	Good conductor of heat	Good conductor of electricity

Group II elements have two electrons in their outermost energy levels. They have similar or almost similar physical properties and chemical properties. Beryllium is the only member of this group that has unique characteristics. The general properties of Group II elements are as summarised below:

Typical metal properties

- High melting points.
- High boiling points.
- Grey silvery surface.
- Good conductors of heat.
- Good conductors of electricity.

Important trends down the group II (with increase of atomic number)

- Melting point and boiling points generally decrease.
- Atoms get bigger.
- Ions get bigger
- Density increases.

Atomic size and ionic size of alkaline-earth metals

The atomic size of alkaline-earth metals increases down the group as illustrated in Fig. 2.4 and 2.5. This is because down the group, the number of energy levels increase. Another reason is that the electrons in the inner energy levels shield the electrons in the outer energy levels from full attraction by the nucleus. The ionic size of the same elements also increases down the group because of additional energy levels. The atomic radii of the elements are bigger than the ionic radii. Alkaline-earth metals form ions by losing all the electrons in the outermost energy level. The ions end up with one energy level less than that of the atom.

Ionization energy

Ionization energy of alkaline-earth metals decreases down the group because as one goes down the group the size of the atoms increase. Because of this, the attraction of electrons in the outermost energy level by the nucleus decrease and hence the ionization energy decreases.

2.9 Trends in chemical properties of alkaline-earth metals

Reaction of alkaline-earth metals with air.

Experiment 2.5

To investigate reaction of beryllium, magnesium and calcium with air.

Apparatus and chemicals

- Bunsen burner

- pair of scissors
- pair of tongs
- sand paper
- magnesium ribbon
- calcium granules

(a) Reaction of magnesium with air

Procedure

1. With a pair of scissors, cut about 4–5 cm of magnesium ribbon. What is the appearance of the magnesium ribbon?
2. Clean the surface of the magnesium with emery paper. What is the colour of magnesium ribbon after cleaning?
3. Hold one end of the magnesium ribbon with the pair of tongs in a bunsen flame as shown in Fig. 2.15 until the magnesium catches fire.
4. Record your observations in your notebook

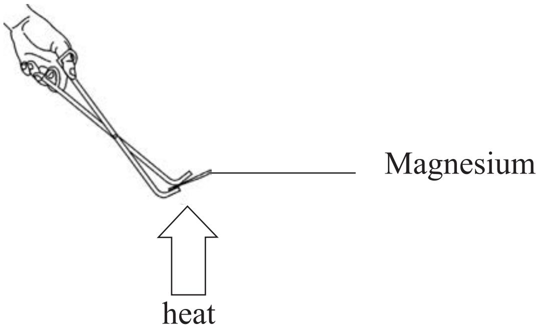


Fig. 2.15 Burning magnesium in air

Caution: Looking directly at burning magnesium will affect your eyes.

(b) Reaction of calcium with air

Procedure

1. Hold a piece of calcium granule tightly with a pair of tongs. What is the appearance of the calcium granule?
2. Hold the pair of tongs with the calcium turning directly in the Bunsen flame and heat strongly.
   - What do you observe?
   - Record your results in Table 2.7.
   - Compare the reaction of magnesium and calcium with air. What is the order of decreasing vigour of reaction with air?
   - Write equations for the reactions between the two metals and air.

3. Record the observations in your notebook.

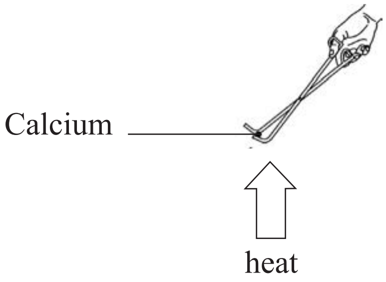


Fig. 2.16 Burning calcium in air

- What do you conclude about reactivity of magnesium and calcium with air.
- How would you expect the reaction of beryllium with air to compare with that of magnesium and calcium.

Table 2.7 Reaction of alkaline-earth metals with air

Alkaline earth-metal	Physical appearance of the metal	Observations of reaction with air	Description of the product formed (write equations for the reaction)	With an arrow show order of decreasing vigour of reaction with air

Generally, alkaline-earth metals burn in air to form simple metal oxides

$$\text{Metal} + \text{oxygen} \rightarrow \text{metal oxide}$$
$$2M(s) + O_2(s) \rightarrow 2MO(s)$$

Note: M represents the metals. They have a valency of two.

Beryllium

Beryllium is a silvery metal. It has a very strong but very thin layer of beryllium oxide on its surface which prevents any further attack from air. Infact getting to the underlying beryllium to react with air is very difficult even at over 600^{o}C. However, powdered beryllium metal does burn in air to give a white beryllium oxide.

$$\text{Beryllium} + \text{oxygen} \rightarrow \text{beryllium oxide}$$
$$2Be(s) + O_2(g) \rightarrow 2BeO(s)$$

Magnesium

Magnesium burns readily with an intense bright white flame to produce a white powder of magnesium oxide.

Magnesium	+	oxygen	→	magnesium oxide
$2Mg(s)$	+	$O_2(g)$	→	$2MgO(s)$

Calcium

Calcium is a silvery metal. The surface of the calcium metal is covered with a thin layer of oxide that protects the metal from further attack by air. At first it is reluctant to burn because of the oxide coating, but then bursts *drastically* into a white flame, which burns intensely to form a white solid of calcium oxide.

Calcium	+	oxygen	→	calcium Oxide
$2Ca(s)$	+	$O_2(g)$	→	$2CaO(s)$

Beryllium, magnesium and calcium do not form peroxides when heated in air. It might not be very clear to see the trend of reactivity of alkaline-earth metals with air due to oxide coating on the surface of the metals. However, the reactivity increases as you go down the group due to increasing ease of removing the electrons in the outermost energy level.

Reaction of alkaline-earth metals (Mg and Ca) with water

Experiment 2.6

To investigate the reaction of alkaline-earth metals (Mg and Ca) with water

Apparatus and chemicals:

- beaker
- boiling tubes
- pair of tongs,
- magnesium
- calcium
- water

Procedure

1. Fill a boiling tube with water and invert it in a beaker half-filled with water. Add a piece of calcium to the beaker. Adjust your boiling tube to collect the gas formed
2. When the boiling tube is full, lift it out of the water, keeping it in an

inverted position. Test the gas collected by bringing a burning splint to the mouth of the test tube.

3. Observe the contents of the beaker. Test the solution in the beaker with red and blue litmus papers.
4. Add a small piece of cleaned magnesium ribbon to a beaker half-filled with water and observe. Heat the water to near boiling and observe any change in the reactivity of magnesium.
5. Record your observations in Table 2.8.

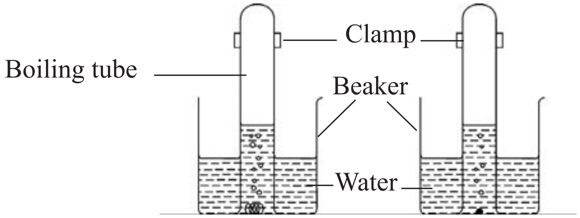


Fig. 2.17 Reaction of alkaline-earth metals with water

Table 2.8 Reaction of alkaline-earth metals with water

Alkaline-earth metals	**Colour change in litmus paper**	**Observations (write equations)**	**Reactivity (show increasing reactivity with an arrow)**
Magnessium			
Calcium			

- What colours do red and blue litmus paper turn in acids? What colours do they turn in bases?
- Rank the metals tested with respect to their reactivity with water.
- What gas was produced when calcium reacted with water? How do you know? Where does this gas come from?
- When alkaline-earth metals react with water, are the resulting solutions acidic or basic? What class of metal compounds is responsible for this property?
- Although the metals tested varied in their rate of reaction with water, they all produced the hydrogen gas and an alkaline solution.
- Write word equations and chemical equations for the reactions of each of the metals with water.
- From your observations are the metals tested placed appropriately in the activity series?

Beryllium

Beryllium does not react with water or steam even when it is red-hot.

Magnesium

Very clean magnesium slightly reacts with cold water to form magnesium hydroxide solution which turns red litmus blue and hydrogen gas. The reaction stops after sometime because the magnesium hydroxide formed is almost insoluble in water and forms a barrier on the magnesium preventing further reaction.

Magnesium + water → magnesium hydroxide + Hydrogen

$$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$$

Calcium

Calcium reacts with cold water with increased vigour than magnesium to give calcium hydroxide solution and hydrogen gas.

Calcium + Water → Calcium hydroxide + Hydrogen

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

The reactivity of alkaline-earth metals with water shows increasing reactivity as you go down the group. This is mainly due to decreasing ionization energy down the group.

Reaction of alkaline-earth metals (Mg and Ca) with chlorine

Experiment 2.7

To investigate the reaction of beryllium, magnesium and calcium with chlorine.

Note: Chlorine is a very poisonous gas that can cause irritation and trigger an asthmatic attack. This experiment should be done in a fume chamber or an open space like a field.

Apparatus and chemicals

- Bunsen burner
- pair of tongs
- sand paper
- gas jars full of chlorine

- magnesium ribbon
- calcium turnings

Procedure

1. Clean the surface of the metal to remove the oxide coating.
2. Coil a small piece of metal (magnesium ribbon) on a deflagrating spoon and light it in the Bunsen flame.
3. Insert the burning magnesium into a gas jar of dry chlorine as illustrated in Fig. 2.18. What do you observe? What is the colour of the product?
4. For calcium, place the metal in a deflagrating spoon and burn.
5. Insert the burning metals into a gas jar of chlorine as illustrated in Fig. 2.19.
6. Record your observation in a table like Table 2.9.
    - compare the reactivity of magnesium and calcium with chlorine

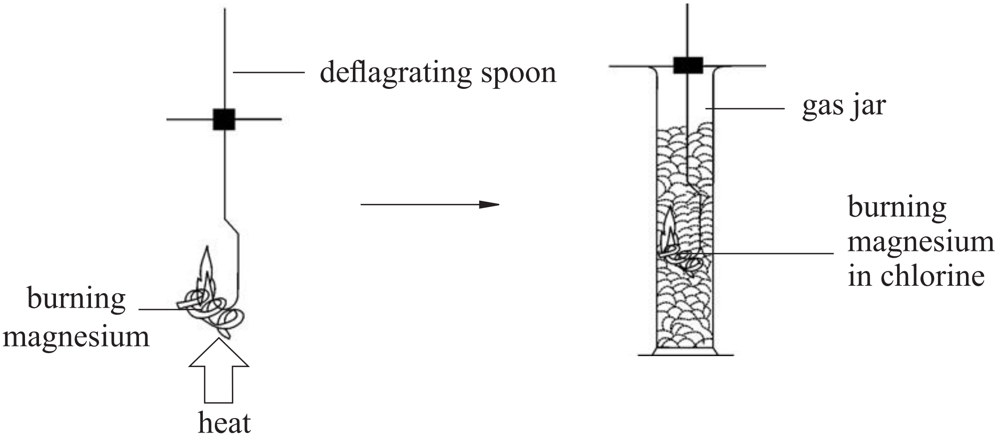


Fig 2.18 Reaction of magnesium with chlorine

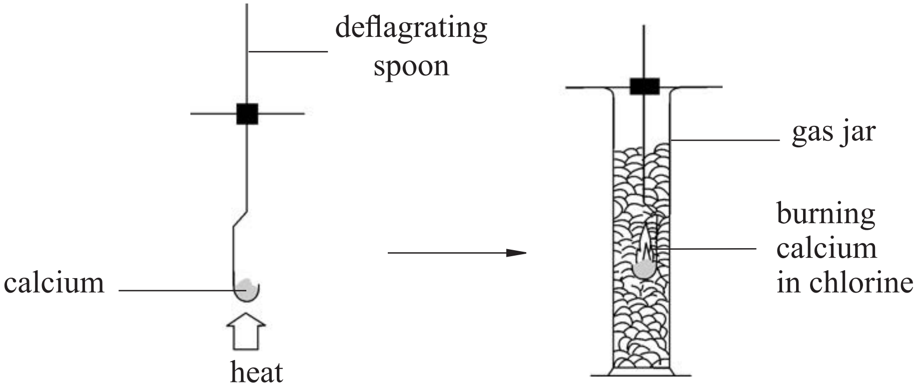


Fig 2.19 Reaction of calcium with chlorine

Table 2.9 Reaction of alkaline-earth metals with chlorine

Elements	Observation of reactions of alkaline-earth metals with chlorine	Write equations for the reaction
Magnesium		
Calcium		

Reaction of alkaline-earth metals with chlorine

Alkaline-earth metals react with chlorine to form chloride salts.

Beryllium

Heated beryllium metal reacts with chlorine to form beryllium chloride.

Beryllium + chlorine → beryllium chloride

$Be(s) + Cl_2(g) \rightarrow BeCl_2(s)$

Magnesium

Magnesium burns in chlorine with a bright light to form a white solid of magnesium chloride.

Magnesium + chlorine → magnesium chloride

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
(white)

Calcium

Calcium appears to burn slowly in chlorine compared to magnesium to form a white solid of calcium chloride.

Calcium + chlorine → calcium chloride

$Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
(white)

Reactivity of alkaline-earth metals with chlorine increases as you go down the group. It appears from the above experiment that magnesium is more reactive than calcium. This is not true. The calcium oxide coating formed on the surface as a result of its reactivity with oxygen forms a protective coat that make the reaction of calcium to appear slower. The inner silvery calcium metal cannot be easily reached by the chlorine.

Reaction of alkaline-earth metals with dilute acids

Experiment 2.8

To investigate reaction of alkaline-earth metals (Mg and Ca) with dilute acids.

Apparatus and chemicals

- Bunsen burner
- test tubes
- test tube rack
- wooden splint
- magnesium ribbon
- calcium
- dilute hydrochloric acid
- dilute sulphuric acid
- dilute nitric acid

Procedure

1. In separate test tubes, place a 2 cm strip of clean magnesium ribbon and a small piece of calcium.
2. Add 5cm^3 of dilute hydrochloric acid to each of the test tubes and observe the reactions.
3. If a gas is produced, hold your thumb over the mouth of the test tube until you feel the pressure build up. Release your thumb and immediately bring a burning splint to the mouth of the test tube.

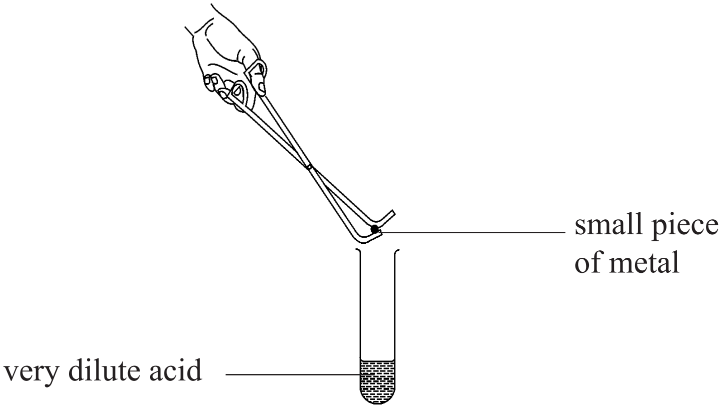


Fig. 2.20 Reaction of alkaline-earth metals with dilute acids

4. Repeat procedures 2 and 3 using dilute sulphuric acid, and then dilute nitric acid in place of the hydrochloric acid with magnesium and calcium.
5. Record your observations in Table 2.10

Table 2.10 Reaction of alkaline-earth metals with dilute acid

Element	**Hydrochloric acid**	**Sulphuric acid**	**Nitric acid**
Magnesium			
Calcium			

- Did all the metals tested react with dilute hydrochloric acid? Rank the metals in order of their reactivity.
- In reactions that produced a gas, was the same gas formed in each case?
- Identify the gas.
- Compare the reactivity of each metal in the three different acids. Is there any difference in the vigour of the reactions in the different acids?
- Do you notice any pattern?
- Compare the order of reactivity of the metals in each of the acids.
- When metals react with acids, two products are produced. One is a gas and the other is a salt. Write a general word equation and chemical equation for the reaction of a metal with an acid.

Note: Reaction of calcium and acids is very explosive. Very dilute acids should be used. Alkaline-earth metals react with dilute acids to form a salt and hydrogen only. The evolution of hydrogen gas is evident by the 'pop' sound produced by the burning splint. Very dilute nitric acid should also be used with magnesium.

Table 2.11 Reaction of alkaline earth metals with dilute acids

Alkaline-earth metal	Dilute acid	Reaction
Berryllium (Be)		The surface of berrylium metal is covered with a thin layer of beryllium oxide that helps protect the metal from attack by acids, but powdered beryllium metal dissolves readily in dilute acids to form a salt and hydrogen gas.
	Hydrochloric acid $HCl(aq)$	Beryllium + Hydrochloric acid → Beryllium chloride + Hydrogen $Be(s) + 2HCl(aq) \rightarrow BeCl_2(aq) + H_2(g)$
	Sulphuric acid $H_2SO_4(aq)$	Beryllium + Sulphuric acid → Beryllium sulphate + hydrogen $Be(s) + H_2SO_4(aq) \rightarrow BeSO_4(aq) + H_2(g)$
	Nitric acid $HNO_3(aq)$	Beryllium + nitric acid → Beryllium nitrate + hydrogen $Be(s) + 2HNO_3(aq) \rightarrow Be(NO_3)_2(aq) + H_2(g)$
Magnesium (Mg)	Hydrochloric acid $HCl(aq)$	Magnesium reacts with dilute hydrochloric acid to form magnesium chloride and hydrogen $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
	Sulphuric acid $H_2SO_4(aq)$	Magnesium also reacts with dilute sulphuric acid to form magnesium sulphate and hydrogen gas $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
	Nitric acid $HNO_3(aq)$	Magnesium reacts with very dilute nitric acid to form magnesium nitrate and hydrogen gas Magnesium + Nitric acid → Magnesium nitrate + Hydrogen $Mg(s) + 2HNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$
Calcium (Ca)	Hydrochloric acid $HCl(aq)$	Very dilute hydrochloric acid reacts with calcium vigorously to form calcium chloride and hydrogen gas Calcium + Hydrochloric acid → Calcium Chloride +Hydrogen gas. $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(s) + H_2(g)$
	Sulphuric acid $H_2SO_4(aq)$	Dilute sulphuric acid reacts much slower with calcium compared with the other acids to form calcium sulphate and hydrogen gas The reaction is slower because the calcium sulphate produced is slightly soluble and forms a coating on the surface of the calcium metal preventing any further reaction. Calcium + Sulphuric acid → Calcium sulphate + Hydrogen $Ca(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2(g)$
	Nitric acid $HNO_3(aq)$	Very dilute nitric acid reacts with calcium to form calcium nitrate and hydrogen gas $Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$

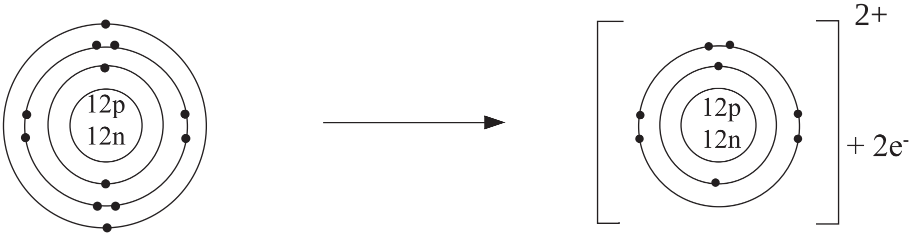
2.10 Similarities of ions and formulae of compounds formed by alkaline-earth metals

As a reminder, Group II elements have two electrons in their outermost energy level. Let us consider magnesium and calcium.

Magnesium Calcium

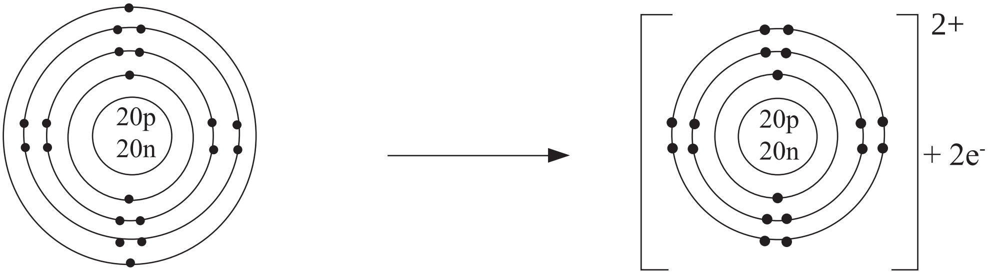
Fig 2.21: Atomic structure of magnesium and calcium

It is easier for Group II elements to lose two electrons than gain 6 electrons to attain the octet rule of 8 electrons in their outermost energy level. They therefore form ions by losing the two electrons in their outermost energy level.



Magnesium atom Mg 2.8.2 Magnesium ion Mg^{2+} 2.8

Fig 2.22: Formation of magnesium ion



Calcium atom Ca 2.8.8.2 Calcium ion Ca^{2+} 2.8.8

Fig 2.23: Formation of calcium ion

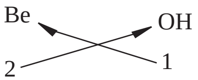
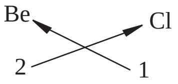
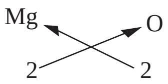
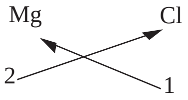
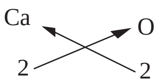
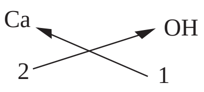
All Group II elements form ions that are positively charged. The ions have a charge of +2 and they have valency of 2.

Table 2.12 Ions and valencies of alkaline-earth metals

Element	**Symbol**	**Ion**	**Valency**
Magnesium	Mg	Mg^{2+}	2
Calcium	Ca	Ca^{2+}	2

- What are the similarities of formulae of oxides, hydroxides and chlorides of alkaline-earth metals?

Table 2.13 Formulae of some alkaline-earth metal oxides, hydroxides and chlorides

Element	**Oxide (O^{2-})**	**Hydroxide (OH^-)**	**Chloride (Cl^-)**
Beryllium symbol Valency Exchange Valency	**(a)**	Beryllium hydroxide Be OH 2 1 Be_1 $(OH)_2$ Formula $Be(OH)_2$	Beryllium chloride Be Cl 2 1 Be_1 Cl_2 Formula $BeCl_2$
Magnesium symbol Valency Exchange Valency	Magnesium oxide Mg O 2 2 Mg_2 O_2 Formula MgO	**(b)**	Magnesium chloride Mg Cl 2 1 Mg_1 Cl_2 Formula $MgCl_2$
Calcium symbol Valency Exchange Valency	Calcium oxide Ca O 2 2 Ca_2 O_2 Formula CaO	Calcium hydrxide Ca OH 2 1 Ca_1 $(OH)_2$ Formula $Ca(OH)_2$	**(c)**

- Follow the same procedure to fill in the empty spaces a, b and c. From the above examples, it is very clear that Group II or alkaline- earth metals have similar ions and formulae of their compounds.

- List down some of the similarities of ions and formulae of Group II elements.

2.11 Uses of alkaline-earth metals

(a) Beryllium

- Used in transmission of X-rays (beryllium transmits X-rays better than aluminium).
- Beryllium alloyed with copper gives a hard strong alloy with high resistance to wear. It is therefore used in computer parts, and other instruments with desirable lightness and stiffness.
- Alloys of beryllium are used as structural materials for high performance aircrafts, missiles, spacecraft and communication satellites among many other things.
- The oxide is used in the nuclear industry.

(b) Magnesium

- It is lighter than aluminium, and it is used to make alloys used for aircraft, car engine casings, and missile construction.
- It is used as a reducing agent for the production of uranium and other metals from their salts.
- Magnesium hydroxide (milk of magnesia), chloride, sulphate (Epsom salts), among others are used in medicine.
- Magnesium oxide is used as brick-liners in furnaces.
- Used in computers for radio-frequency shielding.

(c) Calcium

- It is used as a reducing agent in the preparation of metals such as thorium, uranium, zirconium, etc.
- Calcium forms calcium carbonate which is a component of portland cement.
- Calcium carbonate is used as antacid tablets.

2.12 Group VII - Halogens

The elements in Group VII are collectively known as *halogens*. Their group name comes from the Greek word <u>hal</u>, meaning ‘salt’ and <u>gen</u>, meaning to produce. They are found at the right hand side of the periodic table as shown in Fig. 2.24.

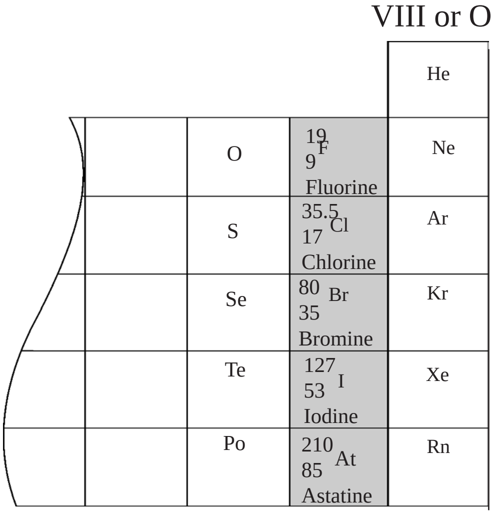


Fig 2.24 The halogens

Table 2.14 Physical properties of halogens

2.13 Trends in physical properties of halogens

Element	Symbol	Electron arrangement	Boiling point	Melting point	Physical appearance	Physical state
Fluorine	F	2.7	-223°C	-187°C	pale yellow	gas
Chlorine	Cl	2.8.7	-102°C	-35°C	pale green-yellow	gas
Bromine	Br	2.8.18.7	7°C	59°C	deep red	liquid
Iodine	I	2.8.18.18.7	114°C	183°C	dark purple	solid

Fluorine (F) and chlorine (Cl), are *diatomic gases* (F_2, Cl_2), at room temperature. Fluorine has a pale yellow colour, while chlorine is a pale green-yellow gas. Both have a choking smell that is harmful to the lungs and nasal passages. Bromine is a dense, deep red liquid that exists as bromine (Br_2) molecule and iodine is a shiny, dark purple solid, existing as diatomic molecule (I_2).

Note: Bromine gives off red brown fumes and iodine gives off purple fumes, both of which are harmful to breathe. All of the halogens are highly reactive and occur only as compounds in nature. The halogens show great similarities to each other in their properties.

The melting and boiling points of group seven elements increase down the group. Fluorine has the lowest and iodine has the highest. The intermolecular forces of attraction become stronger as one goes down group VII. This happens because the molecules become larger and larger down the group. For instance fluorine molecule is small and therefore the forces of attraction between the molecules are weak. Less energy is required to break down these forces of attraction. That is why the melting and boiling point of fluorine is low. Iodine is lower in the group. Iodine molecules are large and intermolecular forces of attraction are strong. It needs more energy to break these forces to separate the molecules.

Group seven elements do not conduct electricity or heat. This is because they do not have free mobile electrons in their outermost energy levels.

The atomic radius of the halogens increases down the group. The size of an atom depends on how strongly the protons in the nucleus attract the electrons in the outermost energy level. The atomic sizes of the halogen atoms are smaller compared to the atomic size of alkali and alkaline-earth metals.

Increase in the nuclear charge increases attraction of the outermost electrons, but as one goes down group seven, there is an additional energy level that makes the positive nuclear charge move further away from the last energy level.

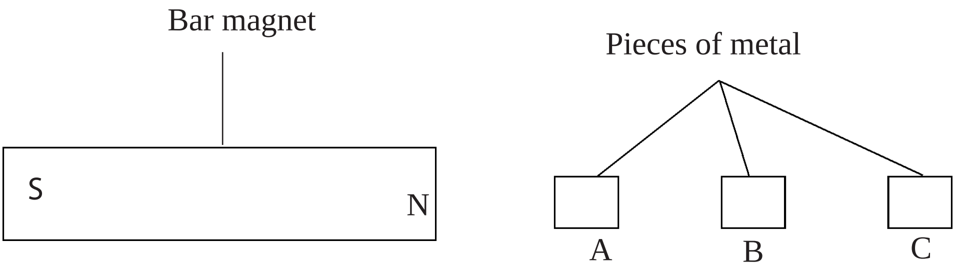
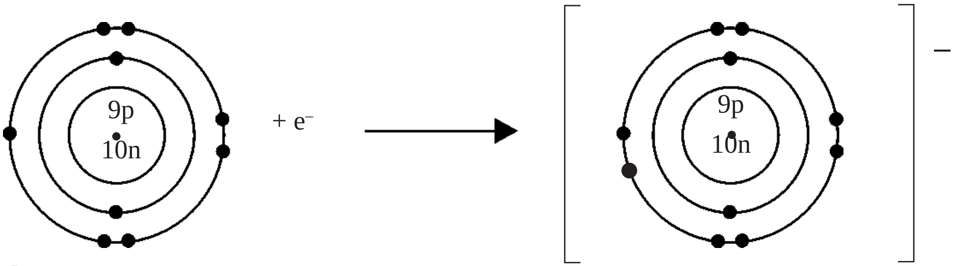


Fig. 2.25 Shielding effect

The magnet attracts metal A more than B or C as illustrated in Fig. 2.25 above. Metal A is shielding the total “feel” of attraction for B and C. Electrons in the inner energy levels shield the outermost from the full feel of nuclear attraction.The size of an atom depends on how strongly the positive protons in the nucleus attract the electrons in the outermost energy level. An addition of an extra energy level with electrons reduces the full effect of attraction for the outermost electrons. The electron within the inner energy levels are said to shield the outermost electrons from the full attraction by protons.

Shielding or screening effect is the reduced attraction by the protons for the outermost electrons caused by the electrons within the inner energy levels.

Ionic sizes of halogens increase down the group but they are larger compared to their atomic sizes. Let us consider how their ions are formed.



Fluorine atom, F 2.7

Fluoride ion, F^- 2.8

Fig. 2.26 Ion formation in fluorine

Fluoride ion has more electrons than protons. The sudden increase in ionic size of fluoride ion in this case is due to the increase in repulsion among the electrons in the complete outer energy level because nuclear charge remains the same. Furthermore,there is a larger number of electrons to be attracted by the same number of protons in the nucleus.

2.14 Trends in chemical properties of halogens

CAUTION: Bromine liquid is very corrosive and can cause burns. It has a very irritating vapour. In case of skin contact, add sodium thiosulphate and large amounts of water and obtain medical attention. For spillages, add sodium carbonate and a lot of water.

- Chlorine is extremely irritating and sometimes may trigger off asthma attacks. Chlorine should always be handled within a fume chamber or in an open space.
- Iodine solid is corrosive and may cause burns. With skin contact wash with large amounts of water.

Reaction of halogens with water

Experiment 2.11

Reaction of halogens with water.

Apparatus and chemicals

- a gas jar of chlorine
- a small amount of bromine liquid
- a few small crystals of iodine solid
- beakers

- Universal indicator solution or blue litmus paper
- water
- stirring rod
- spatula
- boiling tube

Procedure

I. Reaction of chlorine with water.

(Note: This should be done in a fume chamber or open space).

1. To a gas jar full of chlorine, add a small amount of water and universal indicator solution. Litmus paper can also be used.
2. Replace the cover slip and shake as illustrated in Fig. 2.27.
3. Observe any changes.
4. Record your observations as illustrated in Table 2.15.
   - What is the colour of water with indicator?
   - What do you observe when the water is added to the gas and shaken? Explain your observations.

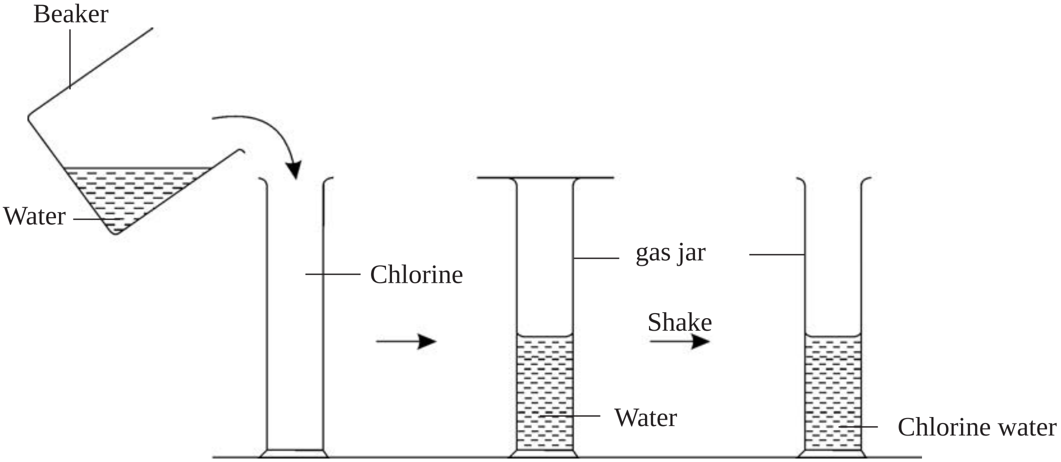


Fig 2.27 Reaction of chlorine with water

II. Reaction of bromine with water.

1. Carefully place 1-2 cm^3 of bromine liquid in a large beaker containing about 250 cm^3 of water.
2. Stir with a glass rod as illustrated in Fig. 2.28.
3. Observe any changes.
4. Put blue litmus paper in the solution formed.
5. Record your observation as illustrated in Table 2.15.
   - What is the initial colour of bromine liquid?
   - What happens when bromine liquid is dissolved in water?

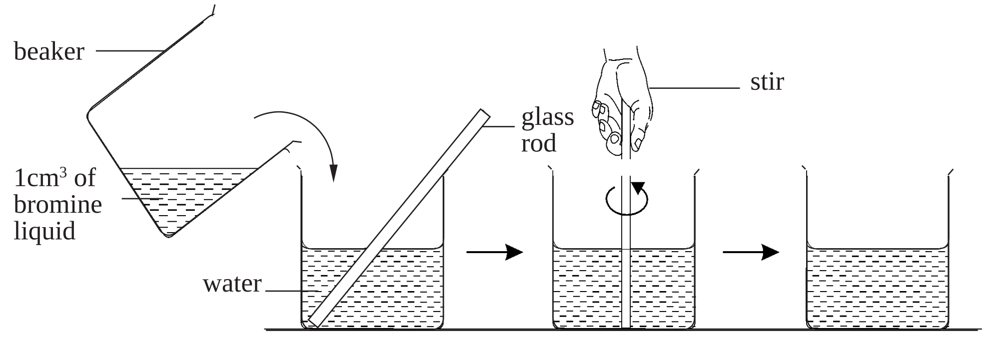
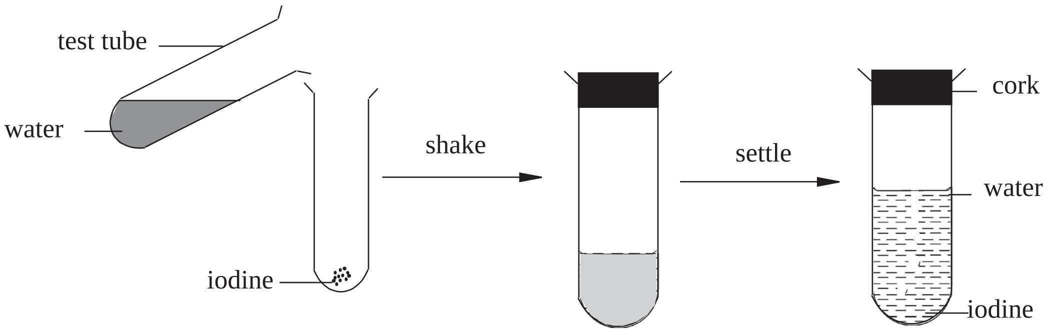


Fig. 2.28 Reaction of bromine with water

III. Reaction of iodine with water



1. Put a spatula full of iodine solid into a boiling tube.
2. Add about 10 cm^3 of water.
3. Put a stopper on and shake.
4. Allow to settle.
5. Record your result in a Table 2.15 below.

Table 2.15 Reaction of halogens with water

Element	**Physical appearance and state of element before reaction**	**Observation (reaction with water)**	**Names of products formed**
Chlorine			
Bromine			
Iodine			

Fluorine reacts with water to form hydrogen fluoride and oxygen gas.

Fluorine + water → hydrogen fluoride + oxygen

$2F_2(g) + 2H_2O(l) \rightarrow 4HF(aq) + O_2(g)$

When chlorine dissolves in water, the colour of universal indicator rapidly changes and chlorine in water dissolves to form a pale yellow solution which has the smell of bleach. Blue litmus paper turns red and is then bleached to white. Chlorine water is acidic and has bleaching properties. The reaction can be summarised in the following equation.

Chlorine + water → hydrochloric acid + chloric(I) acid

$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$

The presence of chloric (I) acid (hypochlorous acid) gives the solution bleaching properties.

Bromine liquid is slightly soluble in water in which it forms a light brown solution which is acidic. It has less bleaching characteristics.

Iodine is almost insoluble in water.

Reaction of halogens with metals

Experiment 2.12

Reaction of halogens with sodium, zinc and iron

Apparatus and chemicals

- thistle funnel
- round bottomed flask
- boiling and side arm tubes
- boiling tubes
- delivery tubes
- clamp stand
- Bunsen burner
- corks
- concentrated hydrochloric acid
- potassium manganate(VII)
- bromine liquid
- iodine
- sodium
- zinc
- iron wool (steel wire)

Procedure

I. Reaction of chlorine with sodium, zinc and iron

1. Prepare chlorine gas as illustrated by the set up in Fig. 2.30.
2. Heat sodium in the deflagrating spoon and insert the burning sodium into a gas jar of chlorine.

3. In case of zinc or iron, place the metal in the tube, as in Fig. 2.30.
4. Pass a steady steam of chlorine gas over heated metal.
5. Heat the metal until the reaction starts and then remove the Bunsen burner.
6. Write down your observations and equations for the reactions.

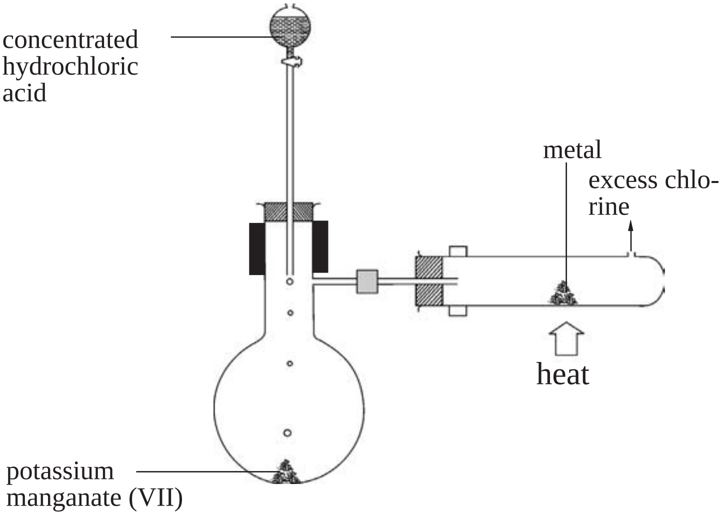


Fig 2.30 Reaction of chlorine gas with zinc or iron

II. Reaction of bromine with metals (sodium, zinc and iron)

1. Bromine is a highly volatile liquid. Place bromine liquid in a corked conical flask with side arm.
2. Place the metal in the tube as shown.
3. Pass a steady stream of bromine gas over heated metals as illustrated in Fig. 2.31.
4. Remember to heat the metal until the reaction starts and continue heating for the duration of the experiment.

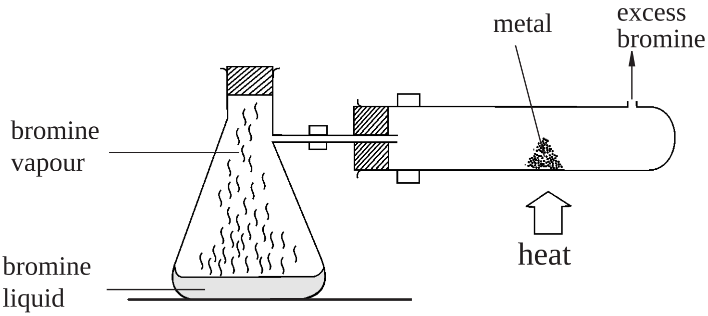


Fig 2.31 Reaction of bromine with metals (sodium, zinc and iron)

- What do you observe when heated metal reacts with bromine?
- What colour are the products?
- Write equations for the reactions.

III. Reaction of iodine with metals (sodium, zinc and iron)

1. Place a few crystals of iodine in a boiling tube.
2. Place the metal a few centimetres above the crystals as illustrated in Fig. 2.32.
3. Heat the metal and the iodine.
    - What happens when the iodine crystals are heated?
    - What do you observe when heated metal reacts with iodine?
    - What are the colours of the products?

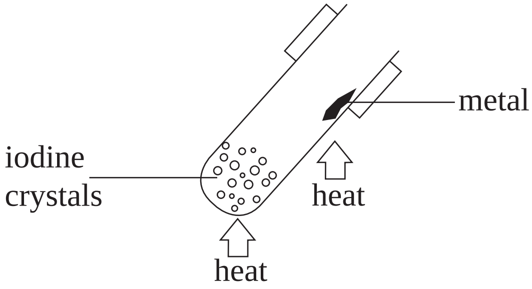


Fig. 2.36 Reaction of metals with iodine

- Write word and balanced chemical equations for the reactions.
- Compare reactivity of iodine, bromine and chlorine with metals.

Note: The metal in this experiment will require to be heated for the duration of the experiment.

Table 2.16 Reaction of halogens (F, Cl, Br and I) with metals (Na, Zn and Fe)

Halogen	Metal	Reaction	Reactivity
Fluorine	Sodium, zinc and iron	Fluorine generally is a highly toxic and very reactive pale yellow gas. Because of this reactivity, Fluorine element or molecule is never found in nature. In fact no other chemical element can displace fluorine. Reactions of fluorine with sodium, zinc and iron cannot be done in an ordinary laboratory.	Reaction of metals with fluorine is explosive and therefore dangerous
Chlorine	Sodium	Heated sodium metal reacts vigorously with chlorine to form a white salt of sodium chloride. Sodium + Chlorine → Sodium chloride $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ (white) The sodium chloride salt produced is soluble in water to give a neutral solution of pH=7, **NOTE:** With sodium once the reaction is on, the heat can be removed and the reaction will proceed.	Reaction of metals with chlorine is vigorous
	Zinc	Zinc reacts when heated in chlorine to form a white salt of zinc chloride. Zinc + Chlorine → Zinc chloride $Zn(s) + Cl_2(g) \rightarrow ZnCl_2(s)$ (white) Zinc chloride salt is also soluble in water to form a neutral solution of pH = 7. **NOTE:** Continued heating is not necessary for this reaction to continue.	
	Iron	When iron wool is heated, it glows red hot. Chlorine gas reacts with the red hot iron wool to form a brown solid of iron(III) chloride. Iron + Chlorine → Iron(III) chloride $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$ (brown) **NOTE:** Continued heating is not required for the reaction to continue.	

Halogen	**Metal**	**Reaction**	**Reactivity**
Bromine	Sodium	Sodium when heated reacts with bromine vapour (gas) to form a white solid, sodium bromide. Sodium + Bromine → Sodium bromide $2Na(s) + Br_2(g) \rightarrow 2NaBr(s)$ (white) **NOTE:** Heating must be continued for the reaction to proceed.	The reaction of bromine with metals is slower compared to the one with chlorine.
	Zinc	Heated zinc reacts with bromine vapour to form a white salt, zinc bromide. Zinc + Bromine → Zinc bromide $Zn(s) + Br_2(g) \rightarrow ZnBr_2(s)$ (white)	
	Iron	Iron must be heated continuously to react with bromine slowly to form a brown salt, iron(III) bromide Iron + Bromine → Iron(III) bromide $2Fe(s) + 3Br_2(g) \rightarrow 2FeBr_3(s)$ (brown)	
Iodine	Sodium	When iodine is heated, it sublimes to form purple iodine vapour which reacts with heated sodium to form a salt, sodium iodide. Sodium + Iodine → Sodium iodide $2Na(s) + I_2(g) \rightarrow 2NaI(s)$ (white)	Iodine reacts with metals very slowly even with strong and continuous heating.
	Zinc	Heated zinc reacts slowly even with continuous heating with iodine vapour to form a salt, zinc iodide. Zinc + Iodine → Zinc iodide $Zn(s) + I_2(g) \rightarrow ZnI_2(s)$ (white)	
	Iron	Iron reacts also very slowly even with continuous heating to form iron(II) iodide instead of iron(III) iodide as in the case of chlorine and bromine. Iron + Iodine → Iron(II) iodide $Fe(s) + I_2(g) \rightarrow FeI_2(s)$ (white)	

2.15 Similarities of ions and formulae of compounds of halogens

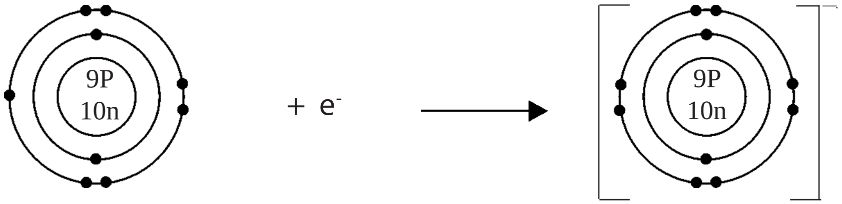
As a reminder, the halogens have seven electrons in their outermost energy level as shown by the following atomic structures of fluorine and chlorine.

Fluorine atom (F) 2.7

Chlorine atom (Cl) 2.8.7

Fig 2.33: Atomic structures of fluorine and chlorine

Since they have the same number of electrons in their outermost energy levels, they form ions in the same way. Halogens form ions by gaining one electron to form negatively charged ions.

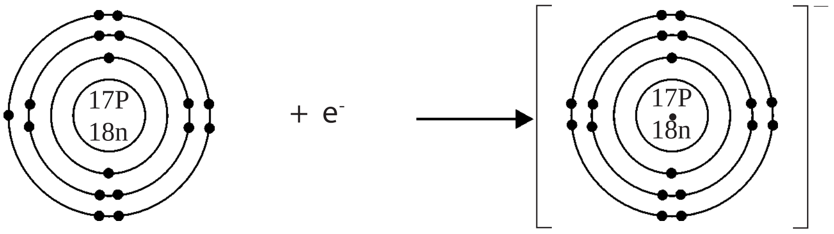


F (Fluorine atom) 2.7

F^- (Fluoride ion) 2.8

Fig 2.34: Ion formation by flourine

All halogens form negatively charged ions and their valency is 1 (one).



Cl (Chlorine atom) 2.8.7

Cl^-(Chloride ion) 2.8.8

Fig. 2.35: Ion formation by chlorine

Table 2.17 Ions and valencies of halogens

Halogen name	Symbol	Ion	Valency
Fluorine	F	F^-	1
Chlorine	Cl	Cl^-	1
Bromine	Br	Br^-	1
Iodine	I	I^-	1

Similarities in formulae of halogen compounds

Halogens will form compounds with the following formulae with metals of valency 1.
Consider sodium (Na), Potassium (K) and Lithium (Li).

Table 2.18 Chemical formulae of metal halides

(a)

Fluorine	Chlorine	Bromine	Iodine
Na^1 F^1 Na_1 F_1 NaF	Na^1 Cl^1 Na_1 Cl_1 NaCl	Na^1 Br^1 Na_1 Br_1 NaBr	Na^1 I^1 Na_1 I_1 NaI

- What are the formulae of compounds of the halogens with potassium and lithium.
- Do you see any similarities?
  Halogens will form compounds with the following formulae with metals of valency 2.
  Consider Magnesium (Mg), Lead (Pb) and Calcium (Ca)

(b)

Fluorine	Chlorine	Bromine	Iodine
Mg^2 F^1 Mg_1 F_2 MgF_2	Mg^2 Cl^1 Mg_1 Cl_2 $MgCl_2$	Mg^2 Br^1 Mg_1 Br_2 $MgBr_2$	Mg^2 I Mg_1 I_2 MgI_2

- What are the formulae of compounds of the halogens with lead and calcium?
- Do you see any similarities?

Consider metals with valency 3 like aluminium (Al) and iron(III) (Fe)
(c)

Fluorine	**Chlorine**	**Bromine**	**Iodine**
Al^3 F^1 Al_1 F_3 AlF_3	Al^3 Cl^1 Al_1 Cl_3 $AlCl_3$	Al^3 Br^1 Al_1 Br_3 $AlBr_3$	Al^3 I^1 Al_1 I_3 AlI_3

- Complete the formulae of the halogens with iron(III) ion.
- Can you observe any similarities?

From the above discussion, it is quite clear that halogens have similar ions and formulae of their compounds.

2.16 Uses of halogens

Uses of fluorine

- It is used in the manufacture of plastics.
- Hydrofluoric acid is extensively used in air conditioning and refrigerators.
- In small amounts, fluoride in water sources prevent tooth decay. It is a constituent of toothpaste. Soluble fluoride in drinking water may damage the teeth of children growing permanent teeth.

Uses of chlorine

- It is used in water treatment all over the world. Even the smallest drinking water supply is now usually chlorinated.
- It is extensively used in the production of paper products, dyestuffs, textiles, petroleum products, medicines, antiseptics, insecticides, foodstuffs, solvents, paints, plastics, and many other products.
- Most chlorine is used in the manufacture of chlorinated cleaning compounds, pulp bleaching, disinfectants, and textile processing.
- Chlorine is used in the manufacture of PVC pipes.

Uses of bromine

- Much bromine is used in making petrol engine anti-knock compounds.
- Used in manufacture of fumigants.
- Flame proof agents.
- Water purification compounds.
- Dyes.

- Medicines.
- Inorganic bromides like silver bromide used in photography.
- Pesticides
- Water purification

Uses of Iodine

- Iodide is used in medicine. A solution of potassium iodide and iodine in alcohol is used as a disinfectant for external wounds.
- Silver iodide is used in photography.
- It is added to table salt to prevent goiter.
- It is used in testing for starch. A deep blue colour is observed when starch reacts with iodine.

2.17 Group O or VIII elements-noble gases

Group O elements are also known as *noble gases*. They were initially known as 'inert' gases until it was discovered that some of its members like krypton and xenon can form compounds.

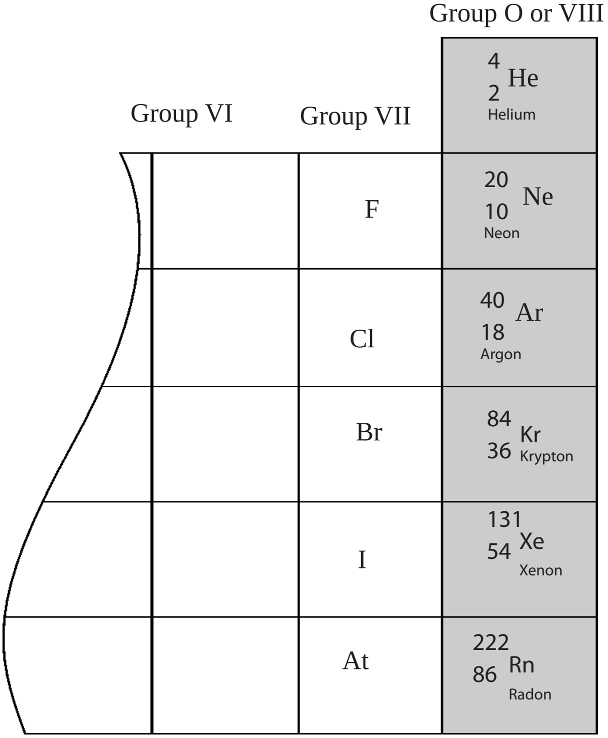


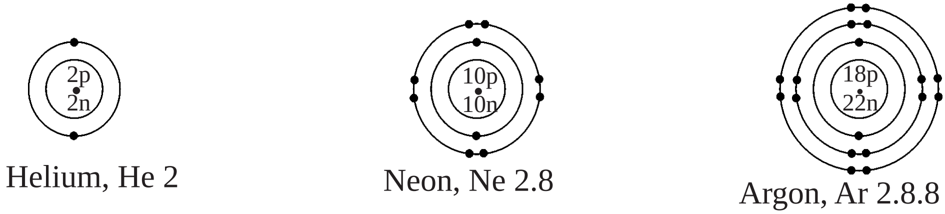
Fig. 2.36 The noble gases

NOTE: At this level we are only going to discuss helium, neon and argon.

Table 2.19 Physical properties of noble gases

Element	Symbol	Electron arrangement	Melting point	Boiling point	Atomic Radius/nm	Physical appearance	Physical state	Thermal conductivity	Electrical conductivity
Helium	He	2	–272˚C	269˚C	0.049	Colourless	Gas	Do not conduct heat	Do not conduct electricity
Neon	Ne	2.8	–249˚C	–246˚C	0.051	Colourless	Gas	Do not conduct heat	Do not conduct electricity
Argon	Ar	2.88	–189˚C	–186˚C	0.088	Colourless	Gas	Do not conduct heat	Do not conduct electricity

Noble gases have full outer energy levels as far as electrons are concerned as illustrated in Fig. 2.37. Due to this they are very unreactive because they have two or eight electrons in their outermost energy levels hence they do not gain or lose electrons. The atomic size increases down the group because of the increasing number of energy levels. Noble gases exist as monoatomic gases at room temperature. They are non-conductors of heat and electricity.



Helium, He 2

Neon, Ne 2.8

Argon, Ar 2.8.8

Fig 2.37: Atomic structure and electron arrangement of noble gases

2.18 Uses of noble gases

Uses of helium

- It is used in filling metrological balloons as it is very light and safer than hydrogen.
- Cooling medium for nuclear reactors.
- A mixture of 80% helium and 20% oxygen is used as an artificial atmosphere for divers and others working under pressure.
- A protective gas for semiconductor materials.
- Pressurizing rockets fuels to liquids.

Uses of neon

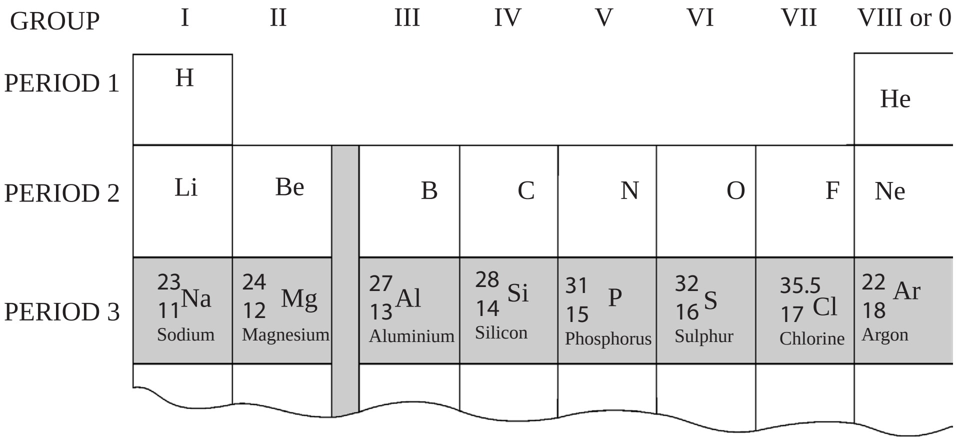
- Used in making neon advertising coloured signs. This is its largest use.
- Used to make high-voltage indicators.
- Neon and helium are used in making gas lasers.
- Liquid neon is an economical refrigerant.

Uses of argon

- Used in electric light bulbs and in fluorescent tubes.
- Used as an inert gas shield for arc welding and cutting .

2.19 Properties and trends across the period three

- The third period has the following elements



GROUP	I	II	III	IV	V	VI	VII	VIII or 0
PERIOD 1	H							He
PERIOD 2	Li	Be	B	C	N	O	F	Ne
PERIOD 3	$^{23}_{11}$Na Sodium	$^{24}_{12}$Mg Magnesium	$^{27}_{13}$Al Aluminium	$^{28}_{14}$Si Silicon	$^{31}_{15}$P Phosphorus	$^{32}_{16}$S Sulphur	$^{35.5}_{17}$Cl Chlorine	$^{22}_{18}$Ar Argon

Fig 2.38 Elements of the period three

2.20 Trends in physical properties of elements across period three

Experiment 2.13

To investigate the conduction of electric current by elements in the period three.

Apparatus and chemicals

- torch cells or battery (a source of D.C)
- crocodile clips
- bulb
- connecting wires
- sodium (**Caution:** sodium is a very reactive metal, handle with a lot of care)
- magnesium
- aluminium foil
- phosphorus (**Caution:** phosphorus is a very reactive element, handle with a lot of care)
- sulphur

Procedure

1. Set up the apparatus as in Fig. 2.39.

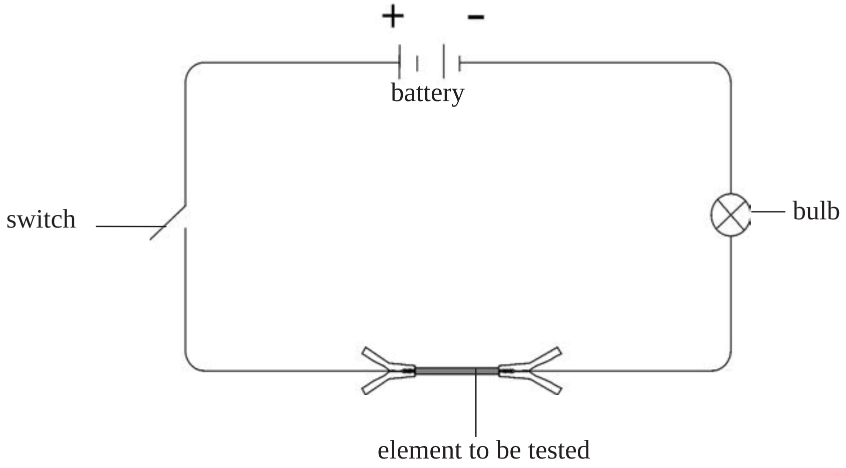


Fig. 2.39: Testing for conductors and non-conductors among period 3 elements

2. Bring the crocodile clips close but do not let them touch. What do you observe in the bulb?
3. Let the crocodile clips touch. What do you observe in the bulb?
4. Now clip magnesium and observe the bulb.
5. Repeat the experiment using the other elements.
6. Record your observation as illustrated in Table 2.19.

Table 2.20: Effect of electricity on elements of the period three

Name of element	**Put a tick (√) if the bulb lights and cross (x) if the bulb does not light**
Sodium	
Magnesium	
Aluminium	
Phosphorus	
Sulphur	

Table 2.21: Trends in physical properties across period three

	Sodium	Magnesium	Aluminium	Silicon	Phosphorus	Sulphur	Chlorine	Argon
Symbol	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic number	11	12	13	14	15	16	17	18
Electronic configuration	2.8.1	2.8.2	2.8.3	2.8.4	2.8.5	2.8.6	2.8.7	2.8.8
Melting points	98˚C	650˚C	660˚C	1410˚C	44˚C	113˚C	–101˚C	–189˚C
Boiling points	883˚C	1090˚C	2467˚C	2680˚C	280˚C	445˚C	–35˚C	–18˚C
Atomic radius/ nm	0.157	0.136	0.125	0.117	0.110	0.104	0.099	0.19
Ionic radius/nm	0.102	0.072	0.053	–	0.190	0.184	0.181	–
Ionization energy/kj/mol	502	744	584	793	1017	1006	1257	1526
Physical appearance	Silvery	Silvery	Silvery	Black	White/red	Yellow	Pale green-yellow gas	Colourless gas
Physical state	Solid	Solid	Solid	Solid	Solid	Solid	Gas	Gas
Electrical conductivity	Good conductor of electricity	Good conductor of electricity	Good conductor of electricity	Does not conduct electricity	Does not conduct electricity	Does not conduct electricity	Does not conduct electricity	Does not conduct electricity
Thermal conductivity	Good conductor of heat	Good conductor of heat	Good conductor of heat	Does not conduct heat	Does not conduct heat	Does not conduct heat	Does not conduct heat	Does not conduct heat
Metallic character	Metal	Metal	Metal	Metalloid (semi-metal)	Non-metal	Non-metal	Non-metal	Non-metal

The periodic table is arranged in order of increasing atomic numbers. As one moves from one element to another from left to right, one more proton is added to the nucleus and one more electron is added to the same outermost energy level as shown in Fig. 2.40.

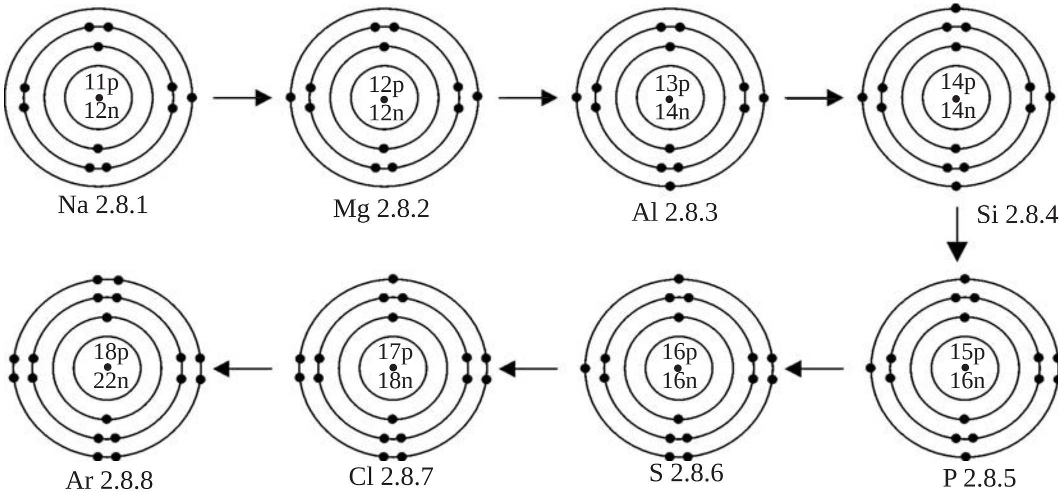


Fig. 2.40: Period three elements showing increasing atomic numbers

All elements in the third period have the first energy level filled with 2 (two) electrons and the second energy level filled with 8 (eight) electrons as illustrated in Fig. 2.40. Sodium begins a 3rd energy level with 1 (one) electron, magnesium has 2 (two) and so on until argon, the last in this periodic. Its third energy level is filled with 8 (eight) electrons. Argon which has a stable electron arrangement is a noble gas.

Atomic radius

Atomic radius of the elements decreases across the period from the left to right see Table 2.21. As we move from left to right across the period, one more proton is added to the nucleus of each successive atom, and one electron is added in the same energy level. The increased positive charge in the nucleus of each successive atom attracts all the electrons towards the nucleus more strongly. Electrons are then drawn more closely towards the nucleus. The nucleus therefore attract the electrons more effectively with an increased positive charge (increased number of protons).

Ionisation energy

The *ionisation energy* is the energy required to remove an electron from a gaseous atom. It increases across the period three. This is because across the period electrons are added into the same energy level. At the same time protons also increase by the same number in the nucleus. This increases the positive (+ve) charge pulling the electrons towards the nucleus hence making the radius reduce across the period. The further away from the positively charged nucleus an electron is, the less strongly the electron is attracted to the nucleus. Therefore, the more easily that electron can be removed. Since atomic radius decreases across the period, the ionisation energy increases across a period.

Other characteristics of elements in the period three

In general metals are hard (except Group I metals), they have metallic lustre, high melting points and boiling points. They are also good conductors of heat and electricity. These characteristics are possessed by sodium, magnesium and aluminium which are the only metals in period three.

The rest of the elements are non-metals. They have low melting and boiling points and are electrical insulators or non-conductors of heat and electricity. Elements to the left of period three exhibit metallic properties; elements to the right show non-metallic properties. Silicon is a semi-metal (metalloid).

2.21 Trends in chemical properties of elements across the period three

The reactivity of elements of period 3 is related to the number of electrons which must be lost or gained during chemical reactions.

Reaction of period three elements with oxygen

Experiment 2.14

To investigate reaction of period three elements with oxygen (Na, Mg, Al, P and S).

Apparatus and chemicals

- Bunsen burner
- deflagrating spoons
- knife
- tile (ceramic)
- filter paper
- pair of tongs
- sand paper
- gas jars full of oxygen
- sodium
- magnesium
- aluminium
- phosphorus (red)
- sulphur

Procedure

I. Reaction of (sodium, phosphorus and sulphur)

1. Cut a small piece, (the size of a rice grain) of sodium. Soak up the oil on the surface of the sodium using a filter paper.
2. Transfer the sodium to a deflagrating spoon using the pair of tongs.
3. Heat the sodium with a Bunsen burner flame until it catches fire.
4. Transfer the burning sodium into a gas jar of oxygen as shown in Fig. 2.41 below.

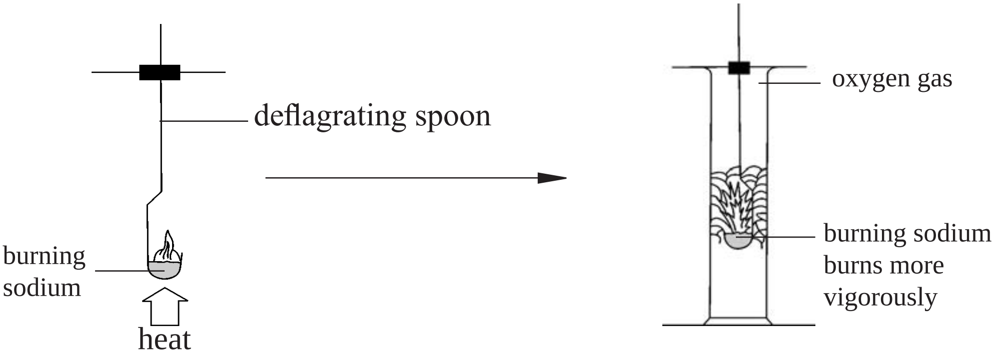


Fig. 2.41: Reaction of sodium with oxygen

5. Repeat the same with phosphorus and sulphur
    - What do you observe?
    - How fast and vigorous are the reactions?
    - Record your observations in Table 2.22.

II: Reaction of (magnesium and aluminium) with oxygen

1. Cut about 4–5 cm of magnesium ribbon.
2. With an emery paper or a knife, clean the surface of the magnesium.
3. Hold the magnesium ribbon securely with a pair of tongs.
4. Hold the end of the tongs in a bunsen flame until the magnesium ribbon catches fire.
5. Transfer the burning magnesium to a gas jar of oxygen as shown in Fig. 2.42.
6. Repeat the same with aluminium foil.

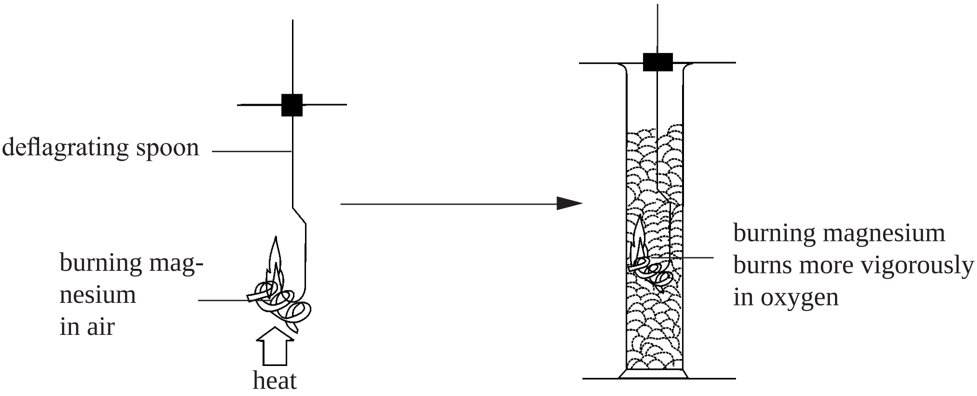


Fig 2.41 Reaction of magnesium with oxygen

Table 2.22: Reactivity of period three elements with oxygen

Elements	Observations
Sodium	
Magnesium	
Aluminium	
Silicon	
Phosphorus	
Sulphur	

- Which metal reacts most easily with oxygen?
- Which metal reacts least easily with oxygen?
- Which non-metal reacts most easily with oxygen?
- Which non-metal reacts least easily with oxygen?
- Suggest reason(s) for your answers.

Sodium metal is the most reactive among the three metals. Aluminium is the *least reactive*. This is because as we move across the period, the atomic sizes decrease. Consequently, the electrons in the outermost energy level are more strongly attracted by the nucleus. Since metals react by donating electrons, as we move across the period, electrons are not so easily released. This decreases reactivity of metals.

Non-metals on the other hand react by gaining or accepting electrons. As we move across the period, the atomic sizes also decrease. This makes sulphur more reactive than phosphorus. The reason is, as the incoming electron comes near the outermost energy level, though it is repelled by other electrons, it will be attracted more by the nucleus since sulphur has a smaller atomic radius than phosphorus. This can be illustrated as in Fig. 2.43.

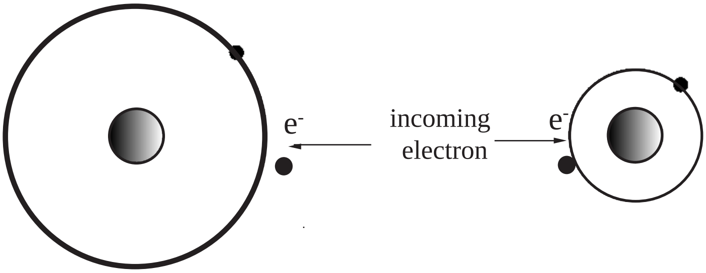


Fig. 2.43 An atom with a smaller atomic size attracts an electron more easily

All the elements in period 3 *except argon* react with oxygen to form oxides. The ratio of oxygen atoms to metal atom in the oxide increases steadily across the period. Some elements form more than one oxide.

Table 2.23: Reaction of period 3 elements with oxygen

	Sodium	**Magnesium**	**Aluminium**	**Silicon**	**Phosphorus**	**Sulphur**	**Chlorine**	**Argon**
Reaction with oxygen	Burns vigorously with a yellow flame to form a white solid	Burns with a bright white light to form a white solid	Cleaned aluminium will burn in oxygen to form a white solid	Silicon powder burns in oxygen at about 400°C to form a solid	Phosphorus burns with a white bright light to form white fumes	Sulphur burns with deep blue flame. A pungent gas is formed	Burns in oxygen to form a gas	No reaction
Oxides formed	Sodium oxide Na_2O Sodium peroxide Na_2O_2	Magnesium oxide Mg0	Aluminium oxide Al_20_3	Silicon(IV) oxide $Si0_2$	Phosphorus(III) oxide P_2O_3 Phosphorus(V) oxide P_2O_5	Sulphur(IV) oxide SO_2	Chlorine(I) oxide Cl_2O Chlorine(VII) oxide Cl_2O_7	—
Nature of the oxides	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic	Acidic	No reaction

Equations for reactions of period three elements with air.

Sodium + oxygen → sodium oxide
$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

Magnesium + oxygen → magnesium oxide
$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Aluminium + oxygen → aluminium oxide
$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

Silicon + oxygen → silicon(IV) oxide
$Si(s) + O_2(g) \rightarrow SiO_2(s)$

Phosphorus + oxygen → phosphorus (III) oxide
$4P(s) + 3O_2(g) \rightarrow 2P_2O_3(s)$

Phosphorous + oxygen → phosphorus(V) oxide
$4P(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$

Sulphur + oxygen → sulphur (IV) oxide
$S(s) + O_2(g) \rightarrow SO_2(g)$

Chlorine + oxygen → chlorine(I) oxide
$2Cl_2(g) + O_2(g) \rightarrow 2Cl_2O(g)$

Chlorine + oxygen → chlorine(VII) oxide
$2Cl_2(g) + 7O_2(g) \rightarrow 2Cl_2O_7(g)$

Note: Burning sodium will also form sodium peroxide (Na_2O_2). Write the chemical equation for this reaction.

Oxides of period three elements

Oxides of elements on the left side of the periodic table are *basic oxides*. Sodium oxide (an alkali) dissolves in water to form sodium hydroxide solution which is alkaline.

Sodium oxide + water → sodium hydroxide
$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

Metal oxides react with acids to form a **salt** and **water** only.

Sodium oxide + hydrochloric acid → sodium chloride + water

$Na_2O(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l)$

Magnesium oxide + sulphuric acid → magnesium sulphate + water

$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l)$

- Write equations for the reaction of magnesium oxide with hydrochloric acid and nitric acid.

Aluminium oxide

Oxides that have the properties of both acid and bases are known as *amphoteric oxides*. Amphoteric oxides react with both basic solutions (alkalis) and acids to form salts and water only.
Aluminium oxide reacts with both acids and alkalis.

Basic character

Aluminium oxide + sulphuric acid → aluminium sulphate + water

$Al_2O_3(s) + 3H_2SO_4(aq) \rightarrow Al_2SO_4(aq) + 3H_2O(l)$

- Write an equation for the reaction of aluminium oxide with hydrochloric acid and nitric acid.

Acidic character

Aluminium oxide + sodium hydroxide + water → sodium aluminate

$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$

- Write equations for the reaction of aluminium oxide with potassium hydroxide solution.

Other oxides are acidic. They dissolve in water forming acids

Phosphorus(III) oxide + water → phosphoric(III) acid

$P_2O_3(s) + 3H_2O(l) \rightarrow 2H_3PO_3(aq)$

Phosphorus(V) oxide + water → phosphoric(V) acid

$P_2O_5(s) + 3H_2O(l) \rightarrow 2H_3PO_4(aq)$

Experiment 2.15

To investigate reaction of period 3 elements with water.

Apparatus and chemicals

- trough
- pair of tongs
- wood splint
- Bunsen burner

- boiling tube
- sand paper
- sodium
- magnesium
- aluminium
- silicon
- phosphorus
- sulphur powder
- gas jar of chlorine

Procedure

I. Reaction of sodium with water

1. Place water in the trough to about three quarters.
2. Add 2–3 drops of universal indicator.
3. Cut a small piece of sodium, about the size of a grain of rice.
4. Pick the piece of sodium with a pair of tongs and drop it into the trough of water as illustrated in Fig. 2.44.

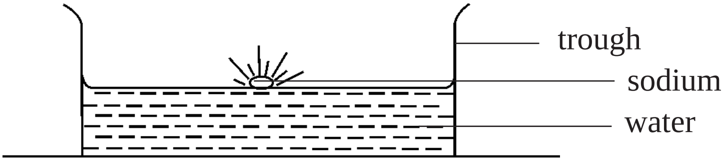


Fig. 2.44 Reaction of sodium with water

   - What do you observe?
   - What colour change occurs with the universal indicator?
5. Record your observations in a table like shown in table 2.24.

II. Reaction of magnesium with water

1. Cut about 4-5 cm of magnesium ribbon. Clean it with the sandpaper.
2. Put water into a boiling tube to about three quarters level. Add magnesium.
   - What do your observe?
3. Heat the boiling tube with the magnesium and water over a bunsen flame.
   - Is there any change in the reactivity of magnesium and water with increase in temperature of the water?
   - How is the reaction of magnesium with steam?
4. Record your results as shown in Table 2.24.

III. Reaction of aluminium, phosphorus and sulphur with water

1. Put water into three test tubes.

2. Put small amounts of aluminium foil, phosphorus and sulphur into three boiling tubes with water.
    - What do you observe?
3. Record your observation as in Table 2.24.

Caution: Phosphorus is very reactive in air. Make sure that you return any remaining phosphorus into the bottle. Do not touch it with your bare hands.

IV: Reaction of chlorine with water

1. Add some water into a gas jar of chlorine and shake.
2. Put red and blue litmus papers into the solution
    - What is the colour of the solution formed?
    - What happens to the red and blue litmus papers?

Table 2.24 Record of period 3 elements reactions with water

Elements	**Observations and equations**
Sodium	
Magnesium	
Aluminium	
Silicon	
Phosphorus	
Sulphur	
Chlorine	

Reaction of period 3 elements with water

Table 2.25 Reaction of period 3 elements with water

	Sodium	**Magnesium**	**Aluminium**	**Silicon**	**Phosphorus**	**Sulphur**	**Chlorine**	**Argon**
Reaction with water	Reacts violently with cold water. It melts forming a ball on surface of water, darts over the surface producing a hissing sound	Reacts very slowly with cold water. Reaction is faster with steam	Does not react with cold water	No reaction	No reaction	No reaction	Reacts with water to form a yellow green solution	No reaction
Products	NaOH(aq) and Hydrogen gas(H_2)	With cold water magnesium hydroxide Mg$(OH)_2$(aq) and Hyrogen (H_2)	—	—	—	—	Chloric (l) acid (HOCl)(aq) and Hydrochloric acid HCl(aq)	
Reactivity	⟶ Decrease across period							

Equations for the reactions

Sodium + water → sodium hydroxide + hydrogen
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Magnesium + water → magnesium hydroxide + hydrogen
$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

Magnesium + water → magnesium oxide + hydrogen
$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

Chlorine + water → hydrochloric acid + hypochlorous acid (chloric(i) acid)
$Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HOCl(aq)$

The reactivity of the metals with water decreases across period 3 as shown in Table 2.25. As you move from sodium to aluminium, the ionization energy increases meaning that it is becomes more difficult to remove an electron from the outermost energy level. The reaction of water with these metals involves formation of ions then other aqueous compounds are formed. Non-metals (silicon to argon) except chlorine do not react with water.

Reaction of period 3 elements with dilute acids

Experiment 2.16

To investigate reaction of period 3 elements of the periodic table with dilute acids (dilute H_2SO_4, HCl and HNO_3).
Caution: Do not attempt to react sodium with acid. It is very explosive.

Apparatus and chemicals

- test tubes
- test tube racks
- splint
- Bunsen burner
- labels
- magnesium ribbon
- aluminium foil
- silicon
- phosphorus
- sulphur powder

Procedure

1. Put 30cm^3 of dilute sulphuric acid into 4 test tubes placed in a test tube rack.
2. Add small amounts of the elements (Mg, Al, P, S) to the test tube with the acid.

- What do you observe?

3. Test any gas produced with a lit splint as illustrated in Fig. 2.45(c).
4. Repeat the same procedure for dilute hydrochloric acid and dilute nitric acid.

   NOTE: The gas must be tested immediately it is evolved.
5. Record your observation as illustrated in Table 2.26.
   - Arrange the reactivity of the following elements (Mg, Al, P and S) with acids starting with the least reactive

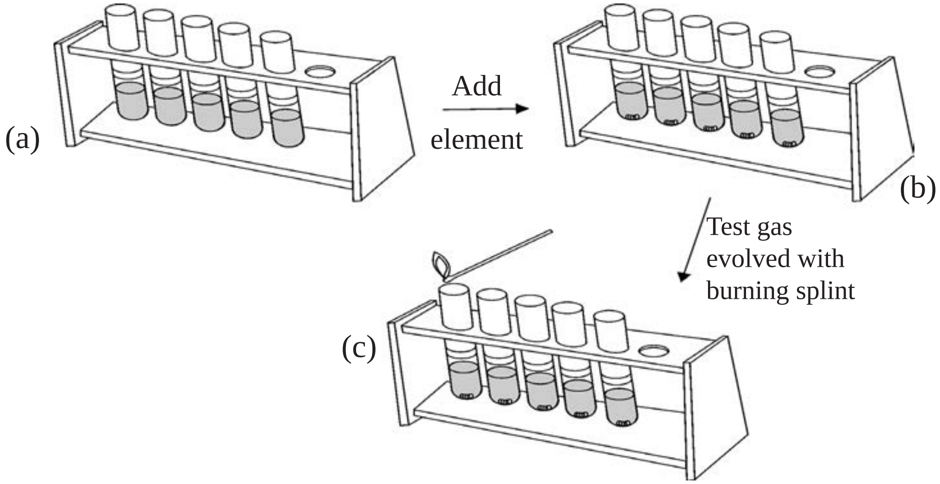


Fig 2.45 Reaction of period 3 elements with dilute acids

Table 2.26 Record of reactions of period 3 elements with dilute acids (H_2SO_4, HCl and HNO_3)

Element	Dilute acid	Observations
Magnesium (Mg)	H_2SO_4	
	HCl	
	HNO_3	
Aluminium (Al)	H_2SO_4	
	HCl	
	HNO_3	
Silicon (Si)	H_2SO_4	
	HCl	
	HNO_3	
Phosphorus (P)	H_2SO_4	
	HCl	
	HNO_3	
Sulphur (S)	H_2SO_4	
	HCl	
	HNO_3	

Table 2.27 Reaction of period 3 elements with dilute acids

	Sodium	Magnesium	Aluminium	Silicon	Phosphorus	Sulphur	Chlorine	Argon
Reaction with dilute acids	**Note:** Do not attempt this experiment. Explosive with dilute acids to form salt and hydrogen	Reacts vigorously with dilute acids to form salt and hydrogen gas	Clean metal of aluminium reacts slower to form salt and hydrogen	— **No reaction at all**	—	—	—	—
Products	HCl Sodium Chloride (NaCl)(aq) Hydrogen gas (H_2) H_2SO_4 Sodium sulphate (Na_2SO_4) Hydrogen gas (H_2) HNO_3 Sodium nitrate ($NaNO_3$) Hydrogen gas (H_2)	Magnesium Chloride ($MgCl_2$)(aq) Hydrogen gas (H_2) Magnesium sulphate ($MgSO_4$) Hydrogen gas (H_2) Magnesium nitrate ($Mg(NO_3)_2$ Hydrogen gas (H_2)	Aluminium chloride ($AlCl_3$)(aq) Hydrogen gas (H_2) Aluminium sulphate $Al_2(SO_4)_3$(aq) Hydrogen gas (H_2) Aluminium nitrate $Al(NO_3)_3$	—	—	—	—	—
Reactivity	⟶ Decreasing reactivity of metals							

Equation for the reactions

CAUTION: Sodium reacts explosively with acids. *Do not attempt this reaction.*

Sodium + sulphuric acid → sodium sulphate + hydrogen

$$2Na(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2(g)$$

- Write balanced equations for reactions of sodium with hydrochloric and nitric acids.

Magnesium + hydrochloric acid → magnesium chloride + hydrogen

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

- Write balanced equations for reactions of magnesium with sulphuric acid and nitric acids.

Aluminium + nitric acid → aluminium nitrate + hydrogen

$$2Al(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3 + 3H_2(g)$$

- Write balanced equations for reactions of aluminium with sulphuric acid and hydrochloric acid.

Reactivity of the metals of period 3 *decreases* from sodium to aluminium due to increasing ionization energy. Ionization energy increases due to increasing nuclear attraction of the electrons in the outermost energy level, making them very difficult to be removed.

Silicon, phosphorus, sulphur, chlorine and argon don't react with dilute acids.

Project

- Construction of models of the periodic table

2.23 Summary

1. Alkali metals are in Group I. They include lithium (Li), sodium (Na), and potassium (K) and others.
2. Alkali metals are soft and can be cut with a knife except lithium. They have low melting points and boiling points.
3. Alkali metals are very reactive compared with other metals. This is because they easily lose their outermost energy level electron.
4. Alkali metals form ions with one positive charge (+1). This is because their atoms lose one electron from their outermost energy levels.
5. The reactivity of alkali metals increases down the group. This is because the atomic radii increase down the group because of the increase in the number of energy levels.
6. Alkali metals form compounds with similar formulae, e.g. lithium sulphate (Li_2SO_4), sodium sulphate (Na_2SO_4), potassium sulphate (K_2SO_4).
7. Alkali metals are very reactive. They tarnish easily in air due to formation of an oxide coating. That is why they are stored in paraffin or oil.
8. Alkaline-earth metals include beryllium, magnesium, calcium and others.
9. Alkaline-earth metals have high melting points and boiling points compared with alkali metals.
10. The reactivity of elements of Group II of the periodic table increases down the group just like in alkali metals.
11. Alkaline-earth metals form ions with two positive charges (+2). This is because they lose two electrons from their outermost energy level.
12. Alkaline-earth metals form compounds with similar formulae e.g. beryllium hydroxide, $Be(OH)_2$, magnesium hydroxide $Mg(OH)_2$ and calcium hydroxide, $Ca(OH)_2$.
13. The halogens or Group VII elements include fluorine, chlorine, bromine and iodine.
14. They are a group of reactive non-metals with low melting points and boiling points. Fluorine and chlorine are gaseous at room temperature. Iodine sublimes on heating.
15. The noble gases, in Group O or VIII include helium (He), neon (Ne) and argon (Ar). They are the least reactive.
16. Their atoms have the most stable arrangement of electrons in the outermost energy level. Therefore they do not gain or lose electrons.

17. Period three elements are sodium (Na), magnesium (Mg), aluminium (Al), silicon (Si), phosphorus (P), sulphur (S), chlorine (Cl) and argon (Ar).
18. Atomic radius decreases as you move across the third period, that is from left to right.
19. The number of protons increases across the third period. This is also the trend of positive attraction of the outermost energy level electrons.
20. The ease of loss of electron(s) decreases across the third period for metals. This is because of attraction of outermost energy level electrons to the nucleus.
21. The ease of gain of electron(s) increases across the period for non-metals. This is because of increase in positive attraction of the incoming electron(s), by the positive nucleus.
22. Ionic radius for metals is smaller than atomic radius. This is because loss of electrons in the outermost energy level involves loss of one energy level.
23. Ionic radius for non-metals is larger than atomic radius. The nucleus of ion "hold" extra electron(s). The total number of negative charges become greater than the number of positive charges in the nucleus. The electrons in the ion are not attracted by the nucleus as strongly as in the neutral atom, and the ion "expands". Also, electrons repel one another and so extra electrons cause extra repulsion.
24. Ionization energy increases across the period.
25. Atomic size decreases across the period.
26. Atomic size/atomic radius determines the ionization energy required for metal to lose electron(s) and therefore their reactivity.
27. Atomic size/atomic radius determine the electron affinity and therefore the reactivity for non-metals.

Revision Exercise 2

1. A student from Chemalel Secondary School discovered a new element in the school compound and named it "Pelounium". It was a solid she could easily cut with a knife. When she put it into water, it reacted vigorously and caught fire, forming an alkaline solution. Answer the following questions concerning "pelounium" (symbol "Pe")
    (a) How did she store it?
    (b) Was it a metal or a non-metal?
    (c) Where would she place it in the metal reactivity series?
    (d) What would be its valency?
    (e) Write the formulae for its chloride, sulphate, carbonate and nitrate.

2. Use the table below to answer the questions that follow.

Element	**Electron Arrangement**	**Atomic Radii (nm)**	**Ionic Radii (nm)**
A	2.8.2	0.136	0.065
B	2.8.7	0.099	0.181
C	2.8.8.1	0.203	0.133
D	2.8.8.2	0.174	0.099

(a) Which elements have similar chemical properties? Explain.
(b) Write the formula of the oxide of the element C.
(c) Which element is a non-metal? Explain.
(d) Name the elements that belong to the same group. Explain.
(e) Explain the trend in the atomic and ionic radii of the elements in (d) above. Which one is most reactive? Explain.

3. Balance the following equations and indicate the state symbols.

(a) $CaCO_3 + HNO_3 \rightarrow Ca(NO_3)_2 + CO_2 + H_2O$
(b) $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

4. Write balanced equations for each of the following reactions. Show state symbols in each equation.

(a) Heating potassium in air to get potassium oxide.
(b) Reaction of sodium in water.
(c) Reaction of copper(II) oxide with hydrochloric acid.

5. Briefly explain the following observations.

(a) Alkaline-earth metals are generally less reactive than alkali metals.
(b) The order of reactivity increases down group I but decreases down group VII.
(c) Noble gases are generally unreactive.

UNIT 3: Structure and bonding

Introduction

The electronic arrangement of noble gases, as discussed earlier, is very stable. For example, helium 2, neon 2.8 and argon 2.8.8. They have the stable duplet (2) for helium and octet (8) for the others, in the outermost energy levels. Consequently, noble gases are in most cases chemically unreactive and do not form compounds with other elements.

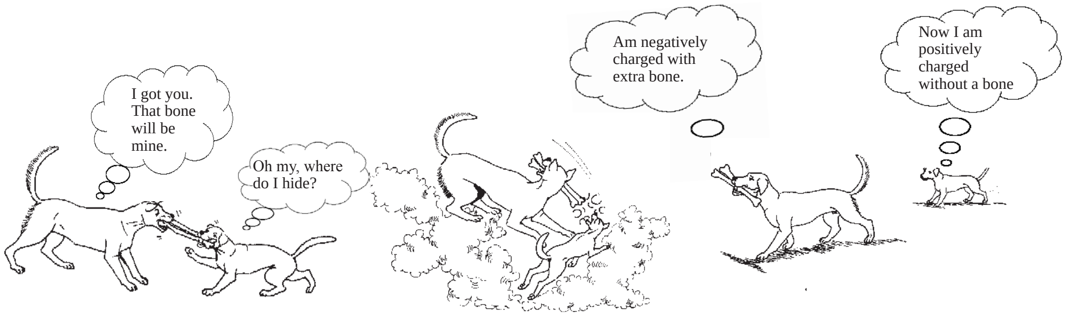
The tendency among other elements is to strive to attain the stable noble gas electronic arrangement. Some will do this by donating electrons; others by accepting them; and still others will share electrons. In this unit, we shall discuss what happens when atoms of elements donate, accept or share electrons, so as to achieve stability of noble gases.

3.1 Ionic chemical bonding

We cannot see ionic bonds. However, ionic bonding can be best imagined as one big greedy dog stealing a bone from a little dog.

Suppose the bone that is up for grabs represents an electron, and the dogs represent atoms of two different elements. When the big dog gains possession of the bone from the smaller dog, it is similar to one atom gaining an electron from another atom.

The bigger dog will be a bone richer, while the smaller dog will be a bone poorer. In atoms, the atom that loses an electron becomes a positively charged ion and the one that gains the electron becomes a negatively charged ion.



a) Two dogs with different strengths each holding a bone of same size. The huge dog is trying to snatch the bone from the smaller dog angrily.

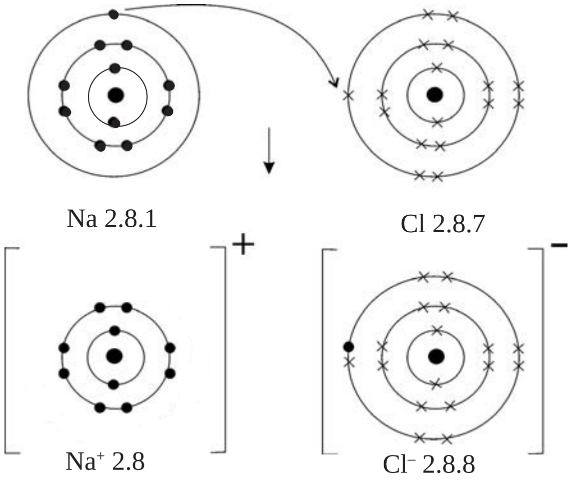
b) The dogs are in a big fight, there is a lot of dust.

c) The huge dog walks away with both bones while the small dog is left very sad.

Fig 3.1 Illustration of ionic bonding

Consider a specific case of ionic bonding between real elements, such as sodium and chlorine. A sodium atom contains 11 protons and has an electronic arrangement of 2.8.1. The arrangement differs from the nearest noble gas electronic structure, that of neon, 2.8 by the presence of one extra electron in the third energy level.

On the other hand, a chlorine atom contains 17 protons and has the electronic arrangement of 2.8.7. It differs from the nearest noble gas electronic arrangement, that of argon, 2.8.8, by missing one electron in the third energy level. In order to attain the stable noble gas electron arrangement, a sodium atom would have to lose the electron in the outermost energy level. The chlorine atom would need to take one electron into its outer energy level to gain the noble gas structure. During ionic chemical bonding of sodium and chlorine atoms, the single electron from the outermost energy level of sodium atom is transferred to the outermost energy level of the chlorine as shown in Fig. 3.2.


(a) Before an electron is transferred.

(b) After electron transfer.

Fig. 3.2 Ionic bonding in sodium chloride (NaCl)

Sodium atom has 11 positive charges (protons) balanced by 11 negative charges (electrons). Sodium ion has only 10 electrons. Therefore, because the positive charges in the nucleus are unchanged, there is **one excess** positive charge on the sodium ion. Similarly, chloride ion has one **negative charge** in **excess** of the positive charges.

A sodium ion is positively charged because of the one excess positive charge. Similarly chloride ion is negatively charged because of the one excess negative charge. These ions have opposite charges.

Because of the attraction of the oppositely charged Na^+ and Cl^- ions they attract and form a **bond** called **ionic or electrovalent bond.** This type of combination is called *ionic bonding*.

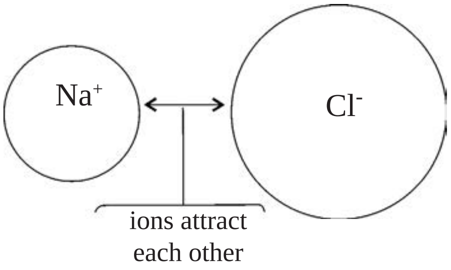


Fig 3.3 Sodium ion combine with chloride ion

Structure of sodium chloride

The structure of sodium chloride contains numerous sodium and chloride ions in equal proportions. The electrical attraction (electrostatic attraction) resulting from their opposite charges constitutes the ionic bond.

The ions arrange themselves into a rigid solid shape called a **crystal**. Each sodium ion is surrounded by six (6) equidistant chloride ions and vice versa.

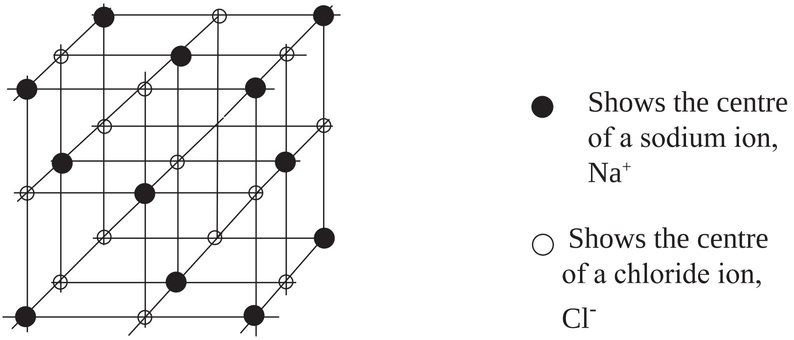


Fig 3.4: Arrangement of sodium and chloride ions in sodium chloride crystal

Sodium and chloride ions crystallise in a pattern (crystal lattice) forming a cube. In an end face of the cube a Na^+ ion occupies the centre, with six Cl^- ions, spaced equally between them. The ions form a **giant ionic structure.**

The attraction forces between the ions are strong and therefore the ions are not free to move but they vibrate within a given space. Consequently the melting point of sodium chloride is high. In solid form, it is a non-conductor of electricity. We will see later that when sodium chloride is melted sodium ions and chloride ions separate and thus their forces of attraction are greatly reduced. When an electric current is applied the ions in molten sodium chloride are free to move thereby conducting electricity.

The positive ions formed as a result of loss of one or more electrons are called **cations**, and their positive charges are equal to the number of electrons lost. Likewise, the negative ions formed as a result of gain of one or more electrons are called **anions** and their negative charges are equal to the number of electrons gained.

The number of electrons lost from, or added to, the outermost energy level of the atom of an element during ionic bonding is equal to the combining power (valency) of that element. Only the **outermost** energy level electrons are involved in ionic bonding. The number of ions involved must balance the valency requirements of elements as shown in the following examples.

(i) Sodium sulphide

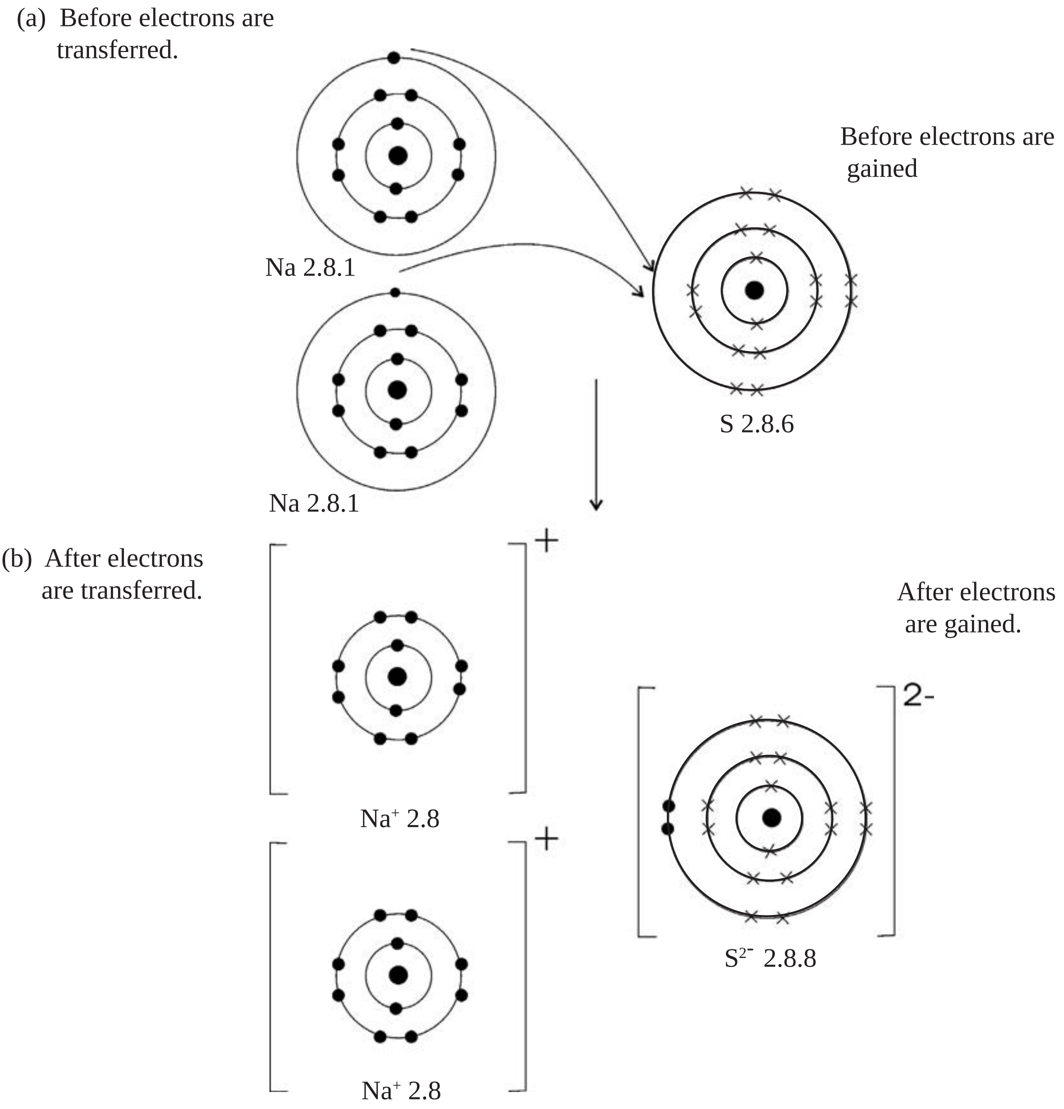


Fig. 3.5 Ionic bond in sodium sulphide (Na_2S)

Therefore the formula of sodium sulphide is Na_2S. The valency electrons (outermost energy level electrons) from the two sodium atoms are transferred to the outermost energy level of sulphur as shown in Fig. 3.5.

(ii) Magnesium oxide

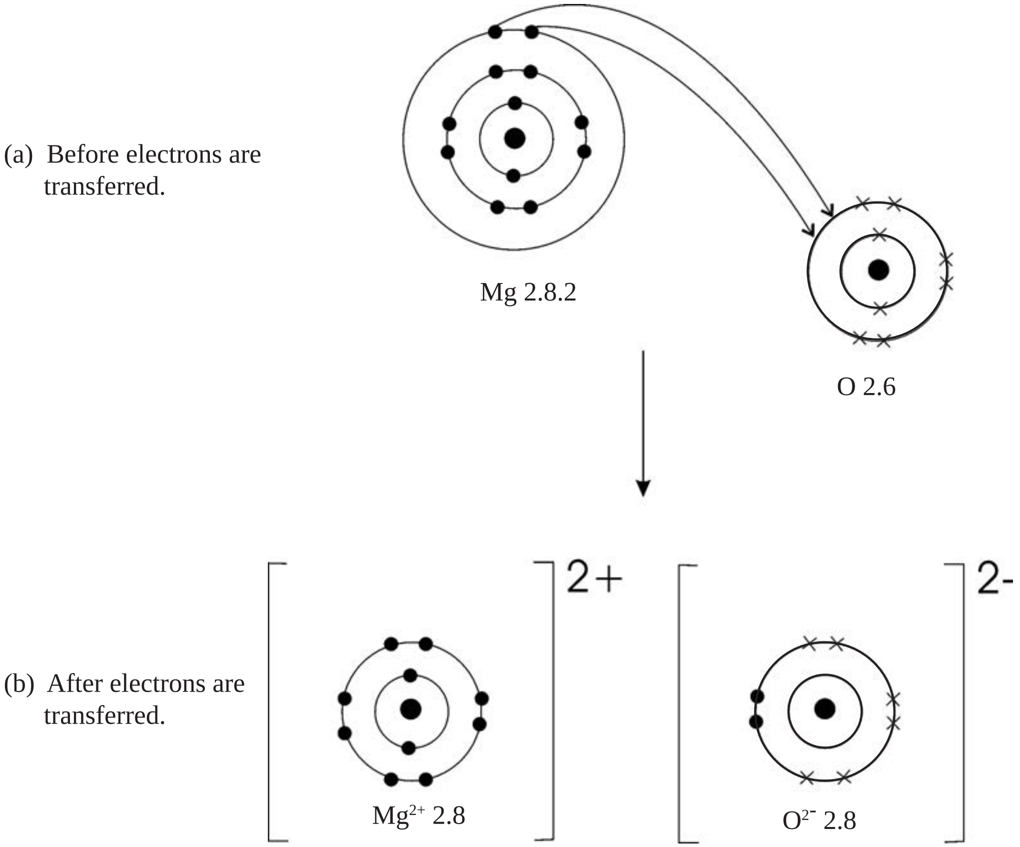


Fig. 3.6 Ionic bond in magnesium oxide

Therefore the formula of magnesium oxide is (MgO). The two valency electrons from one magnesium atom are transferred to one oxygen atom.

Test yourself

Explain using electron arrangement diagrams as in Fig. 3.5 and 3.6 how the following ionic compounds are formed:

(i) Calcium oxide
(ii) Lithium oxide
(iii) Calcium chloride

Properties of ionic compounds

1. Ionic compounds are **electrolytes** (solutions that conduct electric current) when they are in solution or molten form.

2. They have high melting points (non-volatile).
3. They are generally soluble in water.
4. Ionic compounds are insoluble in organic liquids like benzene or propanone.
5. They are usually crystalline solids.

3.2 Covalent chemical bonding

Sharing a pair of electrons forms a *covalent bond.* Imagine two dogs of equal strength fighting over two bones The two dogs have equal attraction to the bones. Since they have equal strength, the dogs end up sharing the pair of bones evenly. If one dog bites both bones on one end and the other bites the other end, no dog will have more bone than the other.

Suppose the bones were electrons and the dogs were atoms. If each bone represents one electron then the two atoms will be sharing a pair of electrons. A pair of electrons shared constitutes a **covalent bond.**

(a) Two dogs with the same strength holding same size of bones in their mouths gazing at each other angrily.

(b) The dogs are in a big fight

(c) None of the dogs wants to lose the bones, hence the dogs are held together by the bones by sharing them.

Fig 3.7 Illustration of a covalent bond

Now let us illustrate covalent bonding in a hydrogen molecule.

Examples

- A hydrogen molecule, (H_2), has two hydrogen atoms linked by a covalent bond.

A hydrogen atom has 1 electron in its only energy level. It is unstable. Therefore two hydrogen atoms combine by each contributing an electron each. Then they share the electron pair equally. The shared pair revolves around both atoms. In effect they have the stable electron duplet arrangement i.e. first energy level with 2 electrons like helium see Fig. 3.8.

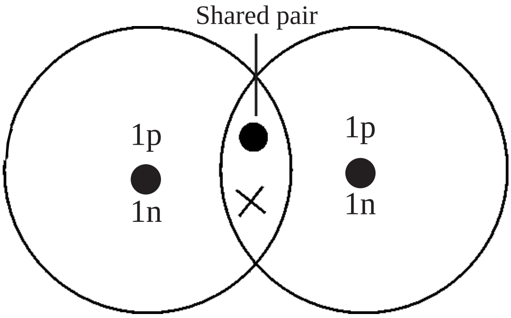


Fig. 3.8: Hydrogen molecule

The shared pair is the covalent bond. It is attracted by the proton of each H atom. Sometimes the shared pair is represented by a short line (–) H – H or the pair can be shown as H $\overset{\bullet}{\times}$ H.

- A molecule of chlorine, Cl_2, contains two chlorine atoms linked by a covalent bond.

These two chlorine atoms have each an electronic arrangement of 2.8.7 with 17 protons. If no other element is available from which electrons may be obtained to make these two atoms have a noble gas electronic structure such as argon (2.8.8), a "shared pair" of electrons is formed. Each chlorine atom contributes **one electron** to the shared pair. This idea of sharing can be shown in a diagram as in Fig. 3.9.

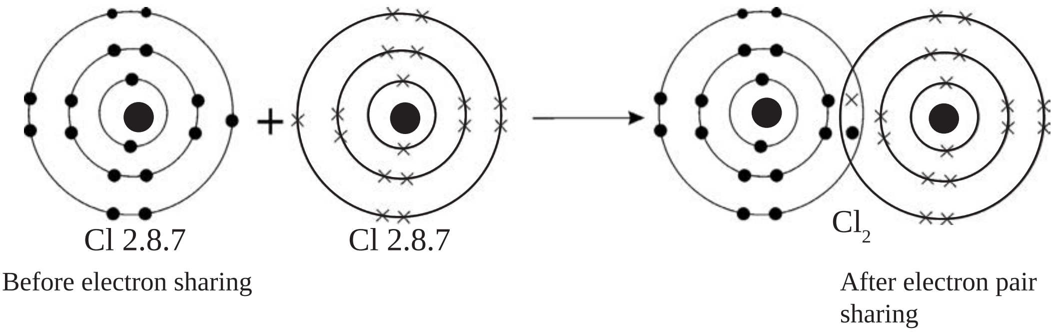


Fig 3.9: Chlorine molecule

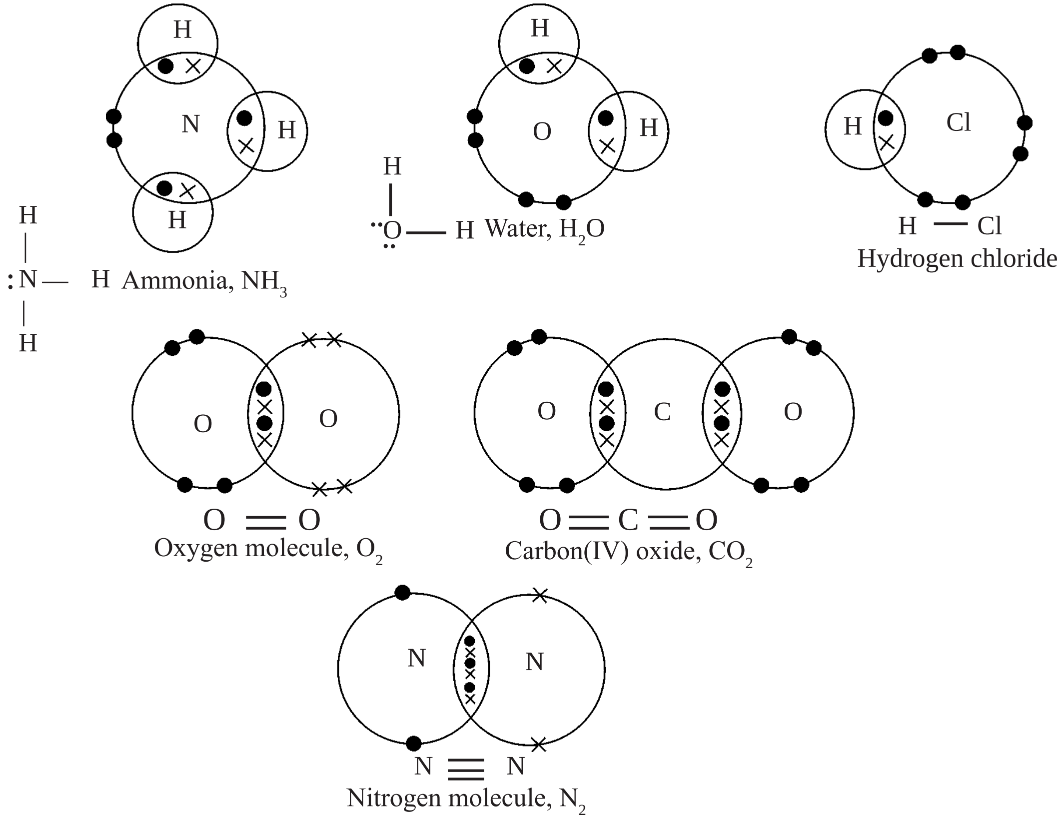
In the chlorine molecule the stable electron octet arrangement for each chlorine atom is achieved.

Again we can see a covalent bond is as a result of sharing two electrons, one from each atom. The nucleus of each atom attracts the shared pair strongly. Like the analogy of the two dogs (atoms) and the two bones (2 electrons) if none of the two dogs release the bones, they will remain together.

Similarly the two atoms will remain joined together by each atom attracting the shared pair.

Some atoms can share more than one pair of electrons. The *covalency* of an atom is the number of electron pairs which it shares. Let us now write simple structural formulae of some covalent compounds. The inner energy levels are omitted. If we understand how to use valency line (–) to represent a pair of electrons we will find it very easy to draw the diagrams.

Notice also that atoms of different elements can also form covalent bonds like in carbon(IV) oxide. See Fig. 3.10. Like we saw in ionic bonding, only **outermost** energy level electrons are involved in covalent bonding.



Ammonia, NH_3

Water, H_2O

Hydrogen chloride

Oxygen molecule, O_2

Carbon(IV) oxide, CO_2

Nitrogen molecule, N_2

Fig. 3.10 Illustration of various molecules

Properties of covalent compounds

1. They are non-electrolytes (don't conduct electric current) when in molten or solution form.
2. They have usually **low melting** and boiling points.
3. They are often **insoluble in water.**
4. They are usually **soluble in organic** solvents like benzene and propanone.
5. Most of them are **gaseous** at room temperature or **volatile liquids**.

3.3 Co-ordinate (Dative) bonding

To understand what a co-ordinate bond is, let us look at the following analogy. In school when you are broke, you may have no money to buy a loaf or sugarcane. But sometimes you may have a friend who can afford the loaf or sugarcane. This friend does not expect you to contribute anything. Your friend buys the loaf or sugar cane and shares with you equally as shown in Fig. 3.11. We can regard the buyer as the atom which contributes both electrons as “donor” and the one sharing, without contributing as “acceptor”.

We saw, in the formation of a simple covalent bond that each atom supplies an electron to be shared. A co-ordinate bond (also called a dative bond) is a covalent bond (a shared pair of electrons) in which both electrons come from only one atom. One of the atoms possesses a **lone pair** of electrons, i.e. a pair of electrons not bonded to any atom. This lone pair is shared with an atom or an ion that needs them to build up, or complete an electron octet or duplet and therefore attain stability. For example ammonia molecule which possesses such a lone pair of electrons is shown in Fig 3.12.

(a) A girl buying loaf from school canteen.
(b) Girl greeting the boy.
(c) The girl shares the loaf with the boy equally.

Fig. 3.11 Illustration of co-ordinate bond analogy

This lone pair can be donated to hydrogen ion (H^+) from an acid to form the ammonium ion, NH_4^+. The nitrogen atom is said to be a “donor” and the hydrogen ion is said to be an “ acceptor”.

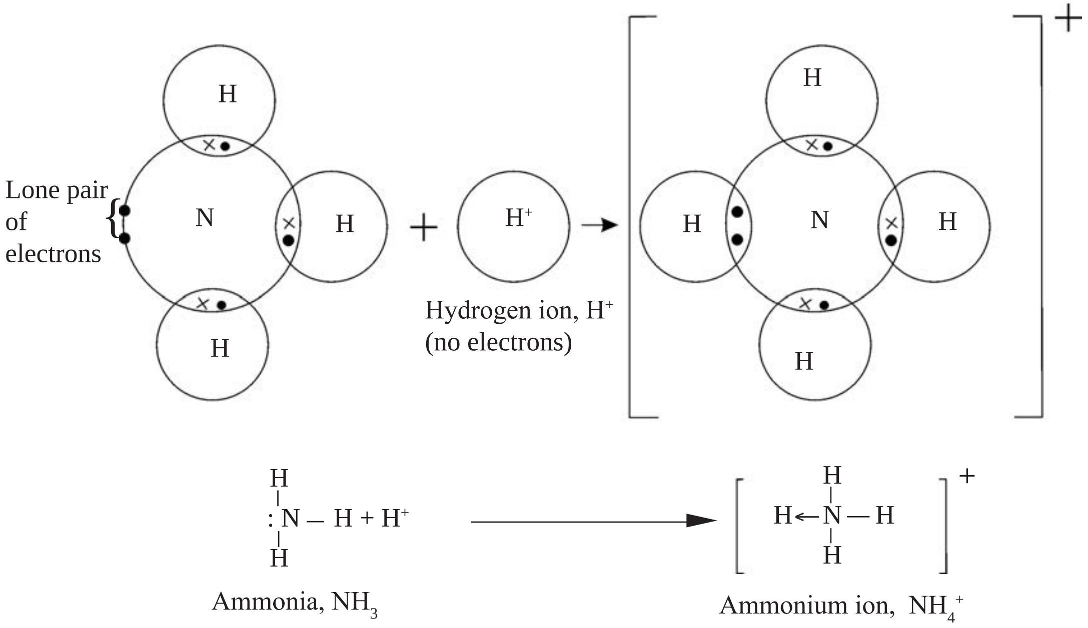


Fig. 3.12 Co-ordinate bond of ammonium ion

The representation of a co-ordinate bond is →. The arrow points from the donor atom to the acceptor atom as shown above. The hydrogen ion contributes the charge on the ion formed.

3.4 Structures of covalent compounds

(i) Giant atomic structures

(a) Diamond

In diamond, the carbon atoms are bonded to each other by covalent bonds in interlacing tetrahedrons throughout the structure. This produces a crystal which is a giant atomic structure. The covalent bonds between the atoms are very strong making diamond very hard.
All four-valence electrons per atom are involved in covalent bond formation with four adjacent carbon atoms.

The carbon atoms in diamond are close to each other and therefore diamond is dense. It has no free electrons (delocalised electrons) and therefore it does not conduct electricity.

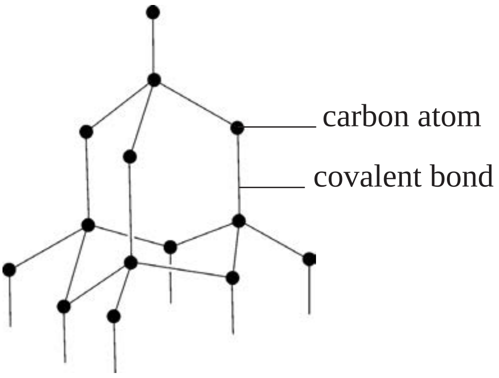


Fig. 3.13 Diamond structure

(b) Graphite

In graphite, the carbon atoms are covalently bonded to each other in hexagonal rings arranged in parallel planes, one on top of the other, with the layers joined by weak van der Waals forces. The layers are able to slip over each other and this makes graphite soft and slippery (See Fig. 6.11). The layers form a giant atomic structure.

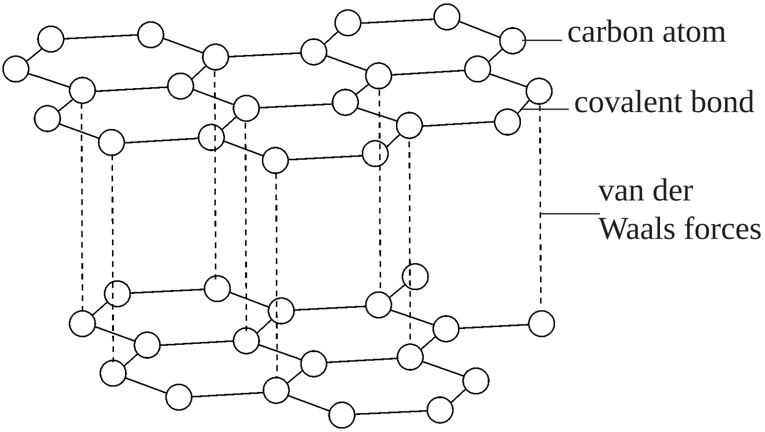
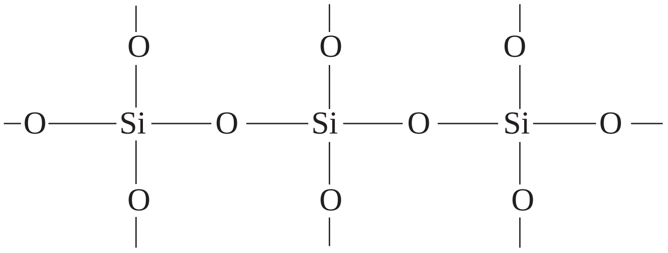


Fig. 3.14 Graphite structure

In the parallel atomic layers (plates) of carbon atoms in graphite, only three valency electrons per atom are used in bond formation. Therefore, some electrons in graphite are free (delocalised) and allow it to conduct electricity.

(c) Silicon (IV) oxide

Silicon is another good example involving atom to atom covalent bonding to form a giant atomic structure. Silicon(IV) oxide is insoluble in water and does not conduct electricity. The bonds between the oxygen – silicon atoms are strong and therefore silicon(IV) oxide has a high melting point.



$$-O-Si-O-Si-O-Si-O-$$

Giant atomic structure of silicon (IV) oxide

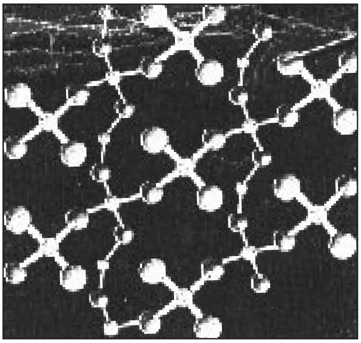


Fig. 3.15 Silicon(IV) oxide structure

(ii) Simple molecular structures

Substances such as iodine, hydrogen, chlorine, nitrogen and carbon(IV) oxide exist as molecules. The atoms in the molecule are held by strong covalent bonds. However, the molecules are held together by weak van der Waals forces. Therefore when heated they easily vapourise or melt. An example is iodine structure.

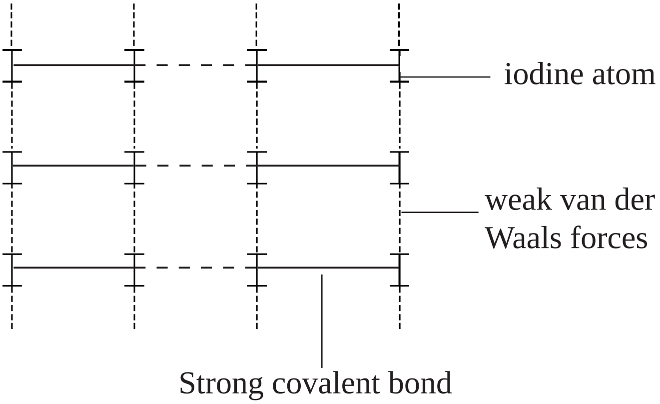


Fig. 3.16 The Structure of iodine

On heating iodine, it sublimes because the weak van der Waals forces are easily broken.

Note: Molecules have discrete units with a definite number of atoms which are covalently bonded. Molecular structures (involving atoms or molecules) have very large indefinite number of atoms in them.

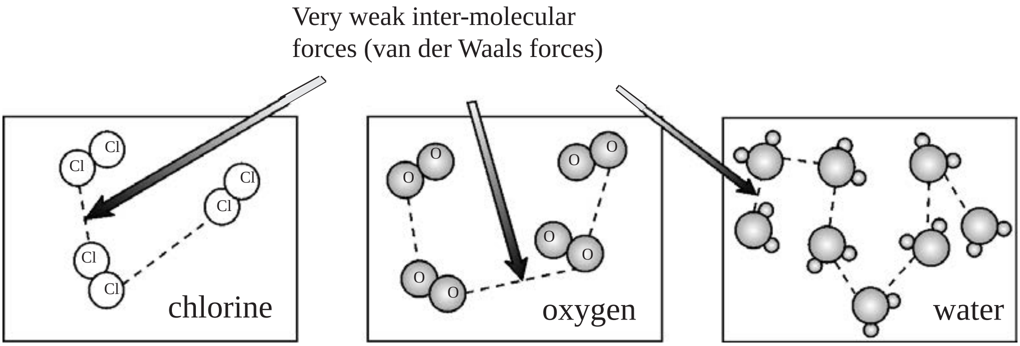


Fig 3.17 Molecules held together by van der Waals forces

3.5 Metallic bonding

The outermost energy level electrons in metals are relatively few. When the atoms of metals are closely packed, each metal atom loses its outer electron(s) which form a "sea" of free electrons (delocalised mobile electrons). The resulting metal positive ions are embedded in the "sea" of electrons. There is attraction between the ions and electrons. This kind of electrostatic attraction between the positive metal ions and the delocalised electrons form the metallic bond.

The ions arrange themselves into a giant metallic structure.

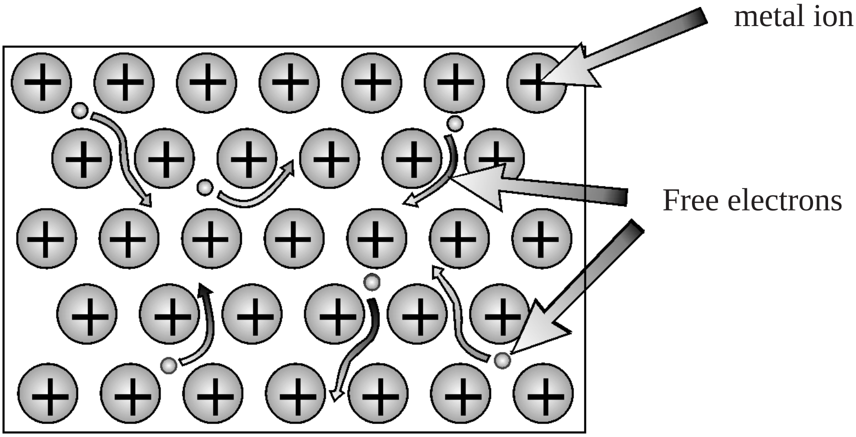


Fig. 3.18 Metallic bond

These delocalised electrons can move on application of an electric current or heat. This explains why metals are good conductors of electricity and heat. Metallic bonding is very strong in some metals like copper, iron, but weak in others like sodium and potassium, which can be cut with a knife. Moving across a period (e.g. period 3) of the periodic table, the number of valency delocalised electrons increases and therefore the strength of the metallic bonds increases. Hence the melting points and boiling points of metals increase across the period. The thermal and electrical conductivity also increases across the period because of the increase in the number of valency electrons (delocalised electrons).

Table 3.1

Element	Na	Mg	Al
Melting point ˚C	98	650	660
Boiling point ˚C	890	1120	2450
Electrical conductivity x10^3 Ohm^{-1} at 250 ˚C	10	16	38
Thermal conductivity kJmol^{-1} S^{-1} at 250˚C	1.34	1.6	2.1
Density g/cm^3	1.34	1.6	2.1

The melting point increase from sodium to aluminium. The change in melting point from magnesium to aluminium is not very big. This is probably because not all the three electrons in aluminium are involved in metallic bonding. As the strength of metallic bond increases across the period, the pull of positive ions towards each other increases thus also increasing the density of the metals.

It might be expected that increase in temperature would speed the movement of free electrons, with a consequent increase in electrical conductivity. In general, however, the electrical conductivity of metals decreases with increase in temperature. This is because increasing temperature produces increased thermal vibration within the metal structure. This upsets the regularity within the crystal and interferes with the ease of movements of electrons within the crystal. It is just like comparing soldiers matching on parade and others matching through a city crowd.

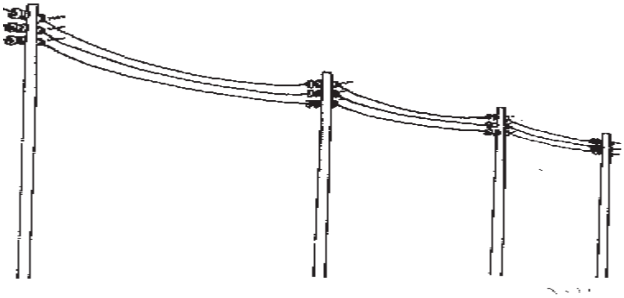


Fig 3.19 Aluminium used for overhead electric cables

3.6 Hydrogen bonding

These bonds can be thought of as two or more dogs that have different desire for bone. The bigger dog has more strength and therefore possesses a larger portion of the bone. Sharing still takes place but **unevenly**. In the case of atoms, the electrons spend more time in the molecule near the atom with the greater attraction of electrons. This makes it develop a partial negative charge represented by (δ^-). The other atom develops a partial positive charge (δ^+). The molecule is said to be **polar.**

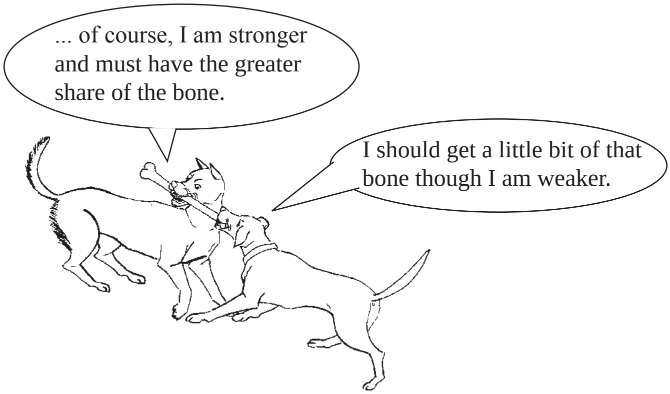
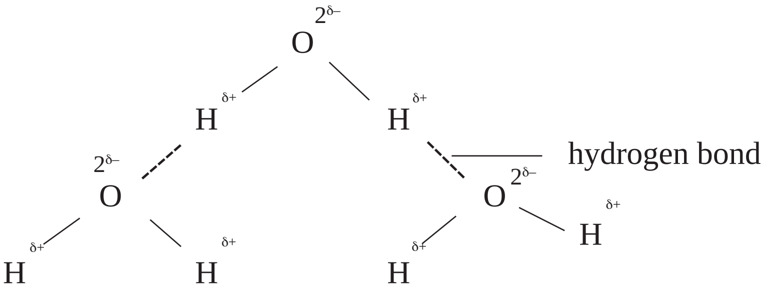


Fig. 3.20 Illustration of hydrogen bonding

For example in a water molecule(H_2O), oxygen atom attracts electrons more than hydrogen. Therefore the water molecule is represented like this:

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

The water molecules can combine together using the polar ends. The attraction between the polar end of the hydrogen (δ+) and the polar end of the oxygen (δ -) is called **hydrogen bond.**



Therefore water has hydrogen bonds between molecules. Had it not been the presence of hydrogen bonds, water would be in gaseous state at room temperature and pressure.

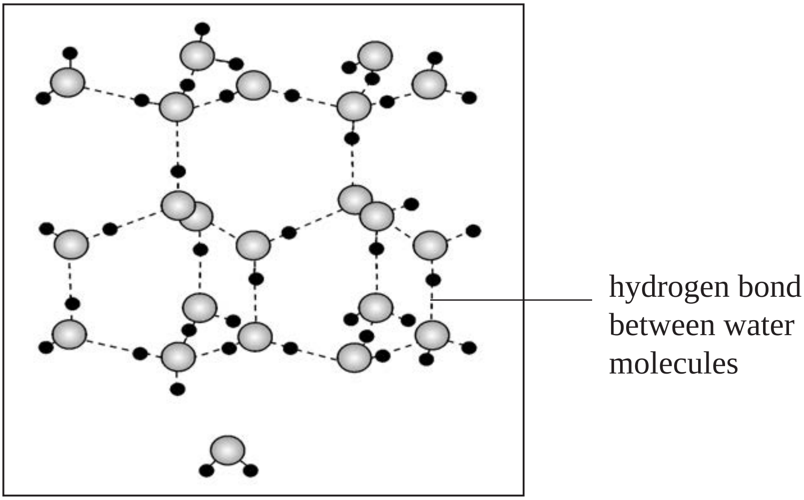


Fig. 3.21 Hydrogen bonding in liquid water

3.7 Types of bonds across period three

(i) *Oxides of elements in period three*

Oxides of sodium, magnesium and aluminium form giant ionic structures. They have therefore high melting points. Sodium oxide is soluble in water, while magnesium oxide is only slightly soluble. When these two oxides dissolve they form basic solutions. They are therefore basic oxides.

Sodium oxide reacts explosively with dilute mineral acids while magnesium oxide reacts at a reasonable rate to form salt and water only. Aluminium oxide is insoluble in water. It however reacts with both dilute mineral acid (explosively) and dilute alkalis to form salts. This oxide is said to be **amphoteric,** that is, it has both acidic and basic properties. Silicon(IV) oxide forms giant atomic structure, while phosphorus(V) oxide, sulphur(IV) oxide, chlorine(VII) oxide (Cl_2O_7) form molecular structures. The oxides of phosphorus and sulphur react with water to form phosphoric(V) acid and sulphurous acid respectively. They are therefore acidic oxides. Silicon(IV) oxide does not react with water or dilute alkali; but it reacts with concentrated alkalis to form salt and water only. Therefore it is regarded as an acidic oxide. The oxides of silicon, phosphorus, sulphur and chlorine have no effect on acids.

Table 3.6

Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Melting point (˚C)	1193	3075	2045	1728	563	30	$^{-}91$
Bonding and structure	Ionic/electrovalent Giant ionic latice			Giant atomic latice (covalent)	Covalent molecular		
Nature of oxide	Decreasingly basic →		Amphoteric	Increasingly acidic →			

Chloride of elements in period three

Chlorides of sodium and magnesium form giant ionic structures which have high melting and boiling points. These two chlorides are soluble in water. The chlorides of aluminium, silicon, phosphorus and sulphur are simple molecular structures. These four chlorides react with water with evolution of heat to form acidic solutions and also give off fumes of hydrogen chloride gas. Some of the hydrogen chloride gas dissolves in water to form hydrochloric acid. This type of reaction is called **hydrolysis** and involves all molecular chlorides.

Table 3.7

Chloride	$NaCl$	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_3	SCl_2	-	-
Melting point (°C)	801	714	sublimes	$^{-}68$	$^{-}91$	$^{-}78$	-	-
Bonding and structure	Ionic/electrovalent Giant ionic lattice		Covalent molecules				-	-
With water	Dissolve		Hydrolysed				-	-

3.8 Applications

Diamond

(i) Jewellery, it has a shinny lustre when polished
(ii) Glass cutter and drilling because it is very hard.

Graphite

(i) As a lubricant, because it is soft.
(ii) Reinforcement of metals and broken bones.
(iii) As electrodes because it conducts electricity.

Aluminium

(i) Overhead electrical cables.
(ii) In making of cooking utensils.
(iii) Wrapping material.
(iv) Silvering of mirrors and different types of reflectors.
(v) Aluminium paint.

All these applications take advantage of the physical and chemical properties of these three elements. These properties are determined by the type of bonding between the atoms of the elements.

For example, let us consider aluminium:

(i) It is a conductor of electricity. It is lighter than copper.
(ii) It is a good conductor of heat. It resists corrosion and has a low density.
(iii) It is soft and malleable.
(iv) It does not tarnish and is a very good reflector of light.
(v) It is a light metal.
(vi) The paint of aluminium is normally applied on iron or steel to prevent corrosion or rusting.

Summary

1. Elements combine or bond in three main forms, namely, *ionic* (electrovalent) bonding, *covalent* bonding and *metallic* bonding.
2. *Ionic bonding* is a chemical bonding formed between oppositely charged ions.
3. Ionic bonding is as a result of donation (loss) and acceptance (gain) of one or more electrons, i.e. transfer of electron(s).
4. The electrostatic attraction resulting from opposite charges of ions is the ionic bond.
5. The charges on ions are equal to the electrons lost or gained. They are also equal to the valency of the element.
6. *Covalent bonding* is as a result of sharing two electrons (a pair) or more than one pair of electrons by atoms.
7. In *co-ordinate or dative bonding*, a lone pair of electrons from one atom is donated to an atom needing them to form an electron octet or duplet.
8. The greater the number of delocalised electrons (valency electrons in metal atoms) the stronger the metallic bond. The stronger the metallic bond the higher the melting and boiling points of the metal.
9. The strength of metallic bond increases from left to right in a period of the periodic table.
10. The greater the number of delocalised electrons the greater the electrical and thermal conductivity. Therefore electrical and thermal

conductivity increases from left to right of the period of the periodic table. Non-metals do not conduct electricty or heat.

11. Increase in temperature interferes with the free movement of the delocalised valency electrons thereby lowering the electrical conductivity of metals and graphite.
12. Densities of metals increase from left to right of the period in the periodic table because the metal positive ions are very close to each other.
13. Giant ionic structures (e.g. sodium chloride), giant atomic structures (e.g. diamond), giant metallic structures (e.g. copper) have high melting and boiling points. A lot of heat energy is required to break the bonds holding the atoms together and break the structure.
14. Substances with low melting points or are gaseous at room temperatures have molecular structures, e.g phosphorus, sulphur, iodine, chlorine. Their atoms have strong covalent bonds between them. Their molecules are held together by weak van der Waals forces.
15. Atoms of argon do not bond and at room temperature, they exist as widely separated gaseous atoms.

Projects

1. Pick about 8 solids, 2 of them should have giant ionic structure, 2 giant atomic structure, 2 giant metallic structure and 2 simple molecular structure. Compare their melting points, thermal and electrical conductivity. What are your findings?
2. Visit a geology laboratory and examine a diamond under a microscope or hand lens. How is diamond cut and polished? Also examine natural graphite. Relate the physical properties of these form of carbon with their uses.

Revision Exercise 3

1. Explain the following:
   (i) Although potassium is a very reactive metal and chlorine is a reactive non-metal, potassium chloride is unreactive.
   (ii) Electrical conductivity of metals decreases as the temperature rises.
2. Draw electronic diagrams for the following species (use dot • and cross x) and show only the outermost energy level electrons): hydroxium ion (H_3O^+), ammonium ion (NH_4^+), water molecule (H_2O and ammonia molecule (NH_3).
3. (a) Study the table below showing properties of substances A, B, C and D and answer the questions that follow.

Substances	**Reaction with oxygen at 25°C**	**Melting point**	**Electrical conductivity**	
			Solid	**Molten**
A	Unreactive	High	Poor	Good
B	Unreactive	High	Poor	Poor
C	Unreactive	High	Good	Good
D	Reactive	Low	Good	Good

Select the substance which is likely to be.

- (i) Copper
- (ii) Magnesium chloride
- (iii) Sodium
- (iv) Diamond

(b) Name the structures in (i) to (iv)

4. Two elements A and B whose atomic numbers are 6 and 8 respectively react to form a compound C.
   - (i) What is the formula of C in terms of A and B
   - (ii) Use dots (•) and crosses (x) to show the bonding in compound C.
5. Both sodium and aluminium are in the same period. The melting points of the metals are 97.8°C and 660°C respectively.
   - (i) Write the electron arrangement of sodium and aluminium.
   - (ii) Why is the melting point of aluminium much higher than that of sodium? (Na = 11, Al = 13)
6. Study the table below which gives some properties of compounds P,Q,R and S,and answer the questions that follow.

	Melting point °C	**Boiling point °C**	**Electrical conductivity in water**
P	-23	77	Does not conduct
Q	-19	74	Does not conduct
R	-85	-61	Does not conduct
S	407	714	Conducts

- (i) Which one of the compounds in the table is ionic? Explain.
- (ii) Select the compound(s) that is/are liquid(s) at room temperature.
- (iii) Select the compound that is gaseous at room temperature.

7. Study the table below and answer the questions that follow:

Elements and their chlorides	**X**	**XCl_2**	**Y**	**YCl_3**	**Z**	**ZCl_4**
Melting point (°C)	650	710	660	sublimes at 180°C	1410	-70
Boiling point (°C)	1120	1420	2450		2680	60
Electrical conductivity in liquid state	Good	Good	Good	Poor	Poor	Poor

Elements X, Y and Z are in period 3.

(i) Which element has a giant covalent structure (giant atomic structure)? Explain.

(ii) Explain the difference in mode of electrical conductivity in substances XCl_2 and Y.

8. Explain why and how the following conduct electricity

(i) Copper (ii) Graphite

9. (i) What is the meaning of giant ionic structure?

(ii) Why do ionic compounds have high melting points?

(iii) Explain why sodium chloride does not conduct electricity when in solid form?

(iv) Why are delocalised electrons free to move within the metal structure?

10. (i) What do we mean by the terms giant atomic structure?

(ii) Why is giant atomic structure sometimes called giant covalent structure?

11. (i) Why is aluminium and copper used in making cooking pans?

(ii) The melting point of silicon is 1410°C while that of chlorine is - –101°C. Explain the difference.

12. Draw electronic diagrams for the following species, use dot (•) and cross (x). Show only the outermost energy level electrons.

(i) Carbon(IV) oxide

(ii) Carbon(II) oxide

(iii) Methane (CH_4)

(iv) Which one has a dative bond?

13. With reference to chlorine:

(i) Explain why most covalent compounds are gaseous at room temperature.

(ii) Why do silicon(IV) oxide, diamond and graphite exist as solids at room temperature while chlorine exists as a gas?

UNIT 4: Salts

Introduction

Most of us assume that the word salt only refers to 'common salt' or sodium chloride which we add to our food every time we eat. Sea water also contains a lot of sodium chloride and for many years sodium chloride has been used in food preservation. However , in Chemistry, the term salt is widely used; many of the chemicals we study in elementary courses in chemistry are salts. We shall see in this unit that the term salt has a more general meaning.

We learnt in Book 1 that acids are compounds which contain hydrogen which may be replaced by a metal to give a salt and hydrogen. This reaction may be represented as follows:

Metal + acid → salt + hydrogen

Let us consider what happens when magnesium reacts with dilute hydrochloric acid as illustrated in Fig. 4.1.

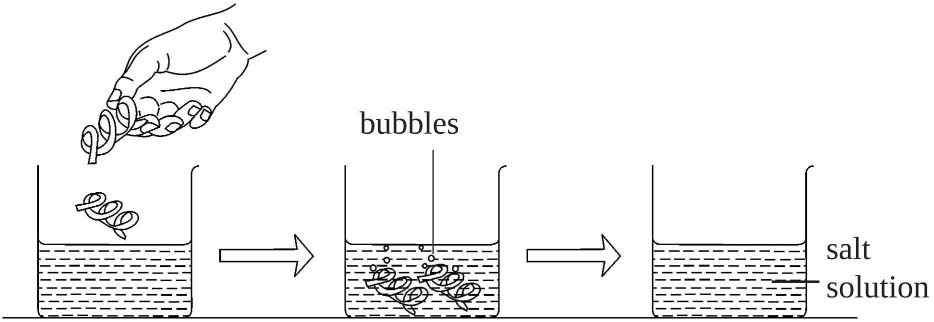


Fig. 4.1: Magnesium reacting with dilute hydrochloric acid to form a salt

Bubbles indicate that a gas (hydrogen) is given off and a clear colourless solution is formed. The salt formed dissolved in the water used to dilute the acid. We can get the salt only if we evaporate the water. The salt is formed when the metal replaces the hydrogen in the acid. So the salt contains magnesium and the acid radical e.g. reaction of magnesium and dilute hydrochloric acid.

Magnesium + hydrochloric acid → magnesium chloride + hydrogen

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

From the above equation, magnesium replaces hydrogen in the hydrochloric acid to form magnesium chloride. In this case the metal itself reacts with acid, and the replacement of the hydrogen is said to be **direct**. Sometimes

an acid reacts with a **base** and the replacement of the hydrogen is **indirect** because the metal comes from a compound of the metal.
For example,

Alkali + acid → salt + water
$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

An ammonium radical can also indirectly replace the hydrogen of the acid to form a salt. Ammonium radical is found in aqueous solution of ammonia.

Ammonia + water → ammonium hydroxide
$NH_3(g) + H_2O(l) \rightarrow NH_4OH(aq)$

Ammonium hydroxide + nitric acid → ammonium nitrate + water
$NH_4OH(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq) + H_2O(l)$

Definition

A *salt* is the product formed when the hydrogen of an acid is replaced directly or indirectly by a metal (or ammonium radical).

Names of salts are derived from the metal or ammonium radical and the acid radical from which they are formed.

When naming salts, the name starts with the name of the metal or ammonium radical in the salt followed by the respective acid radical.

Examples

Name	**Formula**
Sodium chloride	$NaCl$
Ammonium chloride	NH_4Cl
Potassium sulphate	K_2SO_4
Zinc nitrate	$Zn(NO_3)_2$

From the name of the salt you can tell the acid that was used because:

- Hydrochloric acid always gives chlorides
- Nitric acid always gives nitrates
- Sulphuric acid always gives sulphates
- Carbonic acid always gives carbonates
- Ethanoic acid always gives ethanoates
- Phosphoric(V) acid always gives phosphates

4.1 Methods of preparing salts

Salts that dissolve in water are said to be **soluble salts** while those that do not dissolve in water are **insoluble salts**. The method chosen for preparing a particular salt depends on solubility of the salt in water. There are several methods of preparing salts. Soluble salts can be prepared by any of the following methods.

1. Action of an acid on a metal.
2. Action of an acid on an insoluble base (metal oxide or hydroxide).
3. Action of an acid on an insoluble carbonate.
4. Action of an acid on an alkali.
5. Direct combination of elements (synthesis).

We are now going to prepare soluble salts by doing the following experiments.

Preparation of soluble salts

Reaction of a dilute acid with a metal.

It is not safe to react a very reactive metal such as sodium or calcium with an acid.

Experiment 4.1

To prepare zinc sulphate.

Apparatus and chemicals

- beaker (100 cm^3)
- filter funnel
- filter papers
- evaporating dish
- conical flask
- bunsen burner
- dilute sulphuric acid
- zinc

Procedure

1. Pour 50 cm^3 dilute sulphuric acid into a beaker
2. Add zinc granules. What do you observe?
3. If the reaction is too slow, add a little copper(II) sulphate solution and warm gently, but do not boil. What do you observe? Can you give a reason why copper(II) sulphate was added?
4. When all zinc has reacted, add more and wait until it has all reacted. Add more again until some remains unreacted. Give a reason why we must keep on adding zinc until no more can react.
5. Filter off the excess zinc and solid impurities and collect the filtrate in a conical flask. What is the colour of the filtrate?
6. Evaporate the filtrate to dryness as illustrated in Fig 4.2 What do you observe in the evaporating dish?

7. Record your observations in your notebook
8. What do you conclude? Is it possible to prepare copper(II) sulphate crystals using copper metal and sulphuric acid? Explain.

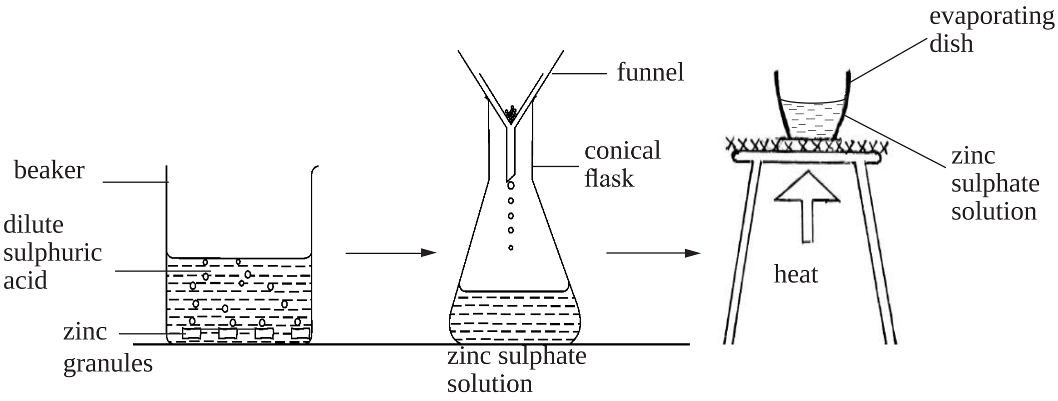


Fig. 4.2 Preparation of zinc sulphate crystals

Obtaining crystals

- To obtain crystals of zinc sulphate, boil the filtrate to evaporate excess water.
- Test the boiling filtrate from time to time to find out if the solution is saturated. You can do this by doing (a) or (b).

(a) Dipping a glass rod into the hot solution to remove a little of the liquid. The solution cools quickly on the glass rod. If crystals form on the glass rod stop heating and let the solution in the dish cool slowly. Cover the solution with a filter paper to keep off dust particles. Leave it until the next lesson.

(b) Pour a little of the hot solution into a test tube and cool with cold water. If small crystals form, then let it cool slowly until the next lesson.

- When good crystals form, pour off the liquid. The liquid we decant is called “mother liquor”.
- Rinse the crystals with a little distilled water.
- Dry the crystals between two filter papers or in sunshine.
- Observe the crystals. Why are the crystals not dried by heating?

When the acid comes into contact with zinc, we observe bubbles.

When copper (II) sulphate solution is added to the reacting mixture it speeds up the reaction. Whenever we add zinc and bubbles are observed, it means that the acid is not used up. If more zinc is added and some remains unreacted, it means all the acid has reacted and a colourless solution of zinc sulphate formed.

When excess water is evaporated and we let the saturated solution cool slowly, big crystals form. Crystals incorporate water as they form. This water is known as **water of crystallisation** and gives shape to the crystals. This explains why the salt formed when we evaporate the filtrate to dryness is a white powder, while the crystals are white "needles" of zinc sulphate or $ZnSO_4{\cdot}7H_2O$. All hydrated salts contain water of crystallisation. But the number of water molecules differ in different crystals.

The reactions that occurred are as follows

(i) Formation of the salt.

$$\text{Zinc} + \text{sulphuric acid} \rightarrow \text{zinc sulphate} + \text{hydrogen}$$
$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

(ii) Formation of crystals from the salt

$$\text{Zinc sulphate} + \text{water} \rightarrow \text{zinc sulphate crystals}$$
$$ZnSO_4(aq) + 7H_2O(l) \rightarrow ZnSO_4{\cdot}7H_2O(s)$$

Iron (II) sulphate and zinc nitrate, are soluble salts which can also be prepared by this method. The equations for the reactions that will take place are as follows.

$$\text{Iron} + \text{sulphuric acid} \rightarrow \text{iron(II) sulphate} + \text{hydrogen}$$
$$Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$$

$$\text{Iron(II) sulphate} + \text{water} \rightarrow \text{iron(II) sulphate crystals}$$
$$FeSO_4(aq) + 7H_2O(l) \rightarrow FeSO_4{\bullet}7H_2O(s)$$

$$\text{Zinc} + \text{Nitric acid} \rightarrow \text{zinc nitrate} + \text{water}$$
$$Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l)$$

Metals, such as sodium, potassium and calcium react vigorously with acids. They should not be used to prepare salts by the method described above. Other metals such as copper, silver and mercury do not react with acids and hence cannot be used to prepare salts by this method. Lead metal when reacted with dilute hydrochloric acid or dilute sulphuric acid forms an insoluble salts, lead(II) chloride or lead sulphate which coats the metal preventing further reaction. Therefore, lead metal is not suitable for preparing salts by this method.

Reaction of a dilute acid with an insoluble metal oxide or hydroxide

Experiment 4.2

To prepare copper(II) sulphate from copper (II) oxide.

Apparatus and chemicals

- conical flask
- beaker
- spatula
- tripod stand
- wire gauze
- bunsen burner
- glass rod stirrer
- filter papers
- filter funnel
- lighter
- evaporating dish
- dilute sulphuric acid
- copper(II) oxide

Procedure

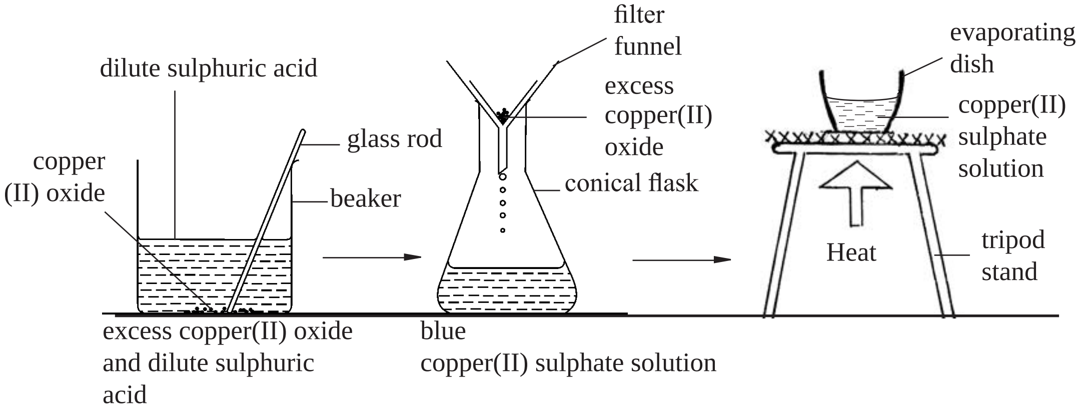


Fig 4.3 Preparation of copper (II) sulphate crystals

1. Pour 50 cm^3 of dilute sulphuric acid into a beaker.
2. Place the beaker on the tripod stand and warm gently but do not boil.
3. Add copper(II) oxide to the hot acid a little at a time. Stir and observe what happens. Test the solution with blue litmus paper.
4. Continue to add the oxide a little at a time stirring until no more oxide reacts with the acid. Give a reason why excess oxide is added. Test the solution with litmus paper.

5. Filter off the excess oxide and collect the filtrate in a conical flask.
6. Evaporate a little of the blue filtrate in an evaporating dish to dryness and note what happens.
7. Heat the remaining filtrate to evaporate excess water and test now and then to see whether the solution is saturated. Leave it until next lesson.
8. When good crystals form, pour off the mother liquor.
9. Dry the crystals between two filter papers or in sunshine.
10. Observe the crystals and note the difference with the salt obtained when the salt was evaporated to dryness in 6 above.
11. Put a few crystals in a test tube and heat strongly. Observe what happens. Place the residue on watch glass or tile and add a few drops of water.
12. What do you conclude? Why is it not possible to prepare copper (II) sulphate by reacting copper metal with dilute sulphuric acid? Explain.

When copper(II) oxide is added to the warm acid, a blue solution is formed. Litmus paper turns red because there is a lot of acid. As more copper(II) oxide reacts the intensity of the blue colour increases. Excess copper(II) oxide is added to make sure that all the acid has reacted. The solution does not change the colour of litmus paper red at this point because there is no acid present in the solution.

When the filtrate is evaporated to dryness, a white powder remains. This is copper(II) sulphate salt without water of crystallisation. It is known as **anhydrous copper(II) sulphate**. But when excess water is evaporated and the filtrate cooled slowly, large blue salt crystals form. Such salt have water of crystallisation and are known as **hydrated salts**. In this case **hydrated copper(II) sulphate** is formed. The salt has 5 water molecules.

When the hydrated copper(II) sulphate is heated it loses water of crystallisation to form **white anhydrous copper(II) sulphate** which is a white powder. When water is added to white anhydrous copper(II) sulphate it turns blue.

The reactions that occurred are as follows:

(i) When the oxide reacted with acid

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

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

(ii) During crystal formation

Copper(II) sulphate + Water → hydrated copper(II) sulphate

$CuSO_4(aq) + 5H_2O(l) \rightarrow CuSO_4.5H_2O(s)$

(iii) On heating the blue hydrated copper(II) sulphate

$$\text{Hydrated copper(II) sulphate} \xrightarrow{\text{heat}} \text{anhydrous copper(II) sulphate} + \text{water}$$

$$CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$$

(iv) On adding water to anhydrous copper(II) sulphate

Anhydrous copper(II) sulphate + water → hydrated copper(II) sulphate

$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_2O(aq)$

This is a reversible chemical reaction and written as follows:

$$\underset{\text{(blue)}}{CuSO_4.5H_2O(s)} \rightleftharpoons \underset{\text{(white)}}{CuSO_4(s) + 5H_2O(l)}$$

The following salts can be prepared in the same way, by adding an excess of the metal oxide or hydroxide to the appropriate acid.

(a) Magnesium sulphate from magnesium hydroxide

Magnesium hydroxide + sulphuric acid → magnesium sulphate + water

$Mg(OH)_2(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + 2H_2O(l)$

(b) Lead(II) nitrate from lead(II) oxide

Lead(II) oxide + nitric acid → lead(II) nitrate + water

$PbO(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + H_2O(l)$

Examples of some other hydrated salts

Name	**Formula**
Magnesium sulphate -7-water	$MgSO_4.7H_2O$
Sodium carbonate -10-water	$Na_2CO_3.10H_2O$
Calcium sulphate - 2- water	$CaSO_4.2H_2O$
Cobalt (II) chloride -6-water	$CoCl_2.6H_2O$

3. Reaction of a dilute acid with metal carbonate or metal hydrogencarbonate

Experiment 4.3

To prepare lead nitrate from lead(II) carbonate.

Apparatus and chemicals

- conical flask
- beaker
- spatula
- tripod stand
- wire gauze
- bunsen burner
- glass rod
- filter paper
- funnel
- lighter
- lead carbonate
- 2M dilute nitric acid

Procedure

1. Pour 50cm^3 of dilute nitric acid into beaker.
2. Add lead carbonate, a little at a time and stir. What do you observe?
3. Test the solution with blue litmus paper to find out whether all the acid has reacted. Filter off excess carbonate. Proceed to crystallize as in Experiment 4.2.
4. Record observations in your notebook
   - Why is it not appropriate to use sulphuric acid or hydrochloric acid when preparing lead sulphate or chloride from lead carbonate?

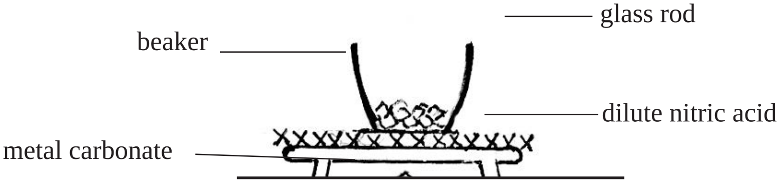


Fig. 4.4 Reacting a metal carbonate with an acid

When the metal carbonate is added to the acid, effervescence occurs showing that a gas is being given off. The gas is carbon(IV) oxide. All metal carbonates react with dilute acids to produce carbon(IV) oxide and a solution of the metal salt so long as the salt formed is soluble.

Carbonate + acid → salt + carbon(IV) oxide + water

Example:

Lead carbonate + nitric acid → lead nitrate + carbon(IV) oxide + water

$PbCO_3(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

The excess carbonate is added to ensure that all the nitric acid has reacted. Other soluble salts that can be prepared in a similar way from the carbonates or hydrogencarbonates include;

(a) (i) Copper(II) sulphate, Copper nitrate, Magnesium sulphate and Zinc sulphate.

Copper carbonate + sulphuric acid → copper(II) sulphate + water + carbon(IV) oxide

$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$

(ii) Copper carbonate + nitric acid → copper nitrate + water + carbon(IV) oxide

$CuCO_3(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2 + H_2O(l) + CO_2(g)$

(iii) Zinc carbonate + sulphuric acid → zinc sulphate + water + carbon(IV) oxide

$ZnCO_3(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l) + CO_2(g)$

(iv) Ammonium carbonate + sulphuric acid → ammonium sulphate + water + carbon(IV) oxide

$(NH_4)_2CO_3(s) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq) + H_2O(l) + CO_2(g)$

(b) Hydrogencarbonates can also used as shown:

(i) Potassium hydrogencarbonate + nitric acid → potassium nitrate + water + carbon(IV) oxide

$KHCO_3(s) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l) + CO_2(g)$

(ii) Sodium hydrogencarbonate + hydrochloric acid → sodium chloride + water + carbon(IV) oxide

$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

However, if the following carbonates are used with the dilute acids as paired, the reaction will stop quickly before all the carbonate has reacted. Explain why.

(i) $CaCO_3(s)$ and $H_2SO_4(aq)$
(ii) $PbCO_3(s)$ and $H_2SO_4(aq)$
(iii) $PbCO_3(s)$ and $HCl(aq)$

(i) Calcium carbonate + sulphuric acid → calcium sulphate + water + carbon(IV) oxide

$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$

(ii) Lead carbonate + sulphuric acid → lead sulphate + water + carbon(IV) oxide

$PbCO_3(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + H_2O(l) + CO_2(g)$

(iii) Lead carbonate + hydrochloric acid → lead(II) chloride + water + carbon(IV) oxide

$PbCO_3(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2O(l) + CO_2(g)$

The salts formed are insoluble and will form a coating on the carbonate preventing further reaction. Calcium sulphate is only slightly soluble. Lead(II) chloride is soluble in hot water.

Reaction of a dilute acid with an alkali

Experiment 4.4

To prepare sodium chloride from sodium hydroxide (a soluble base).

Apparatus and chemicals

- burette
- pipette
- conical flask
- beaker
- funnel
- evaporating dish
- burner
- dilute sodium hydroxide
- dilute hydrochloric acid
- phenolphthalein indicator

Procedure

1. Using a pipette transfer 25cm^3 of the alkali into a conical flask. Then add 2-3 drops of phenolphthalein indicator as shown in Fig. 4.5a. The indicator turns pink. Why is it advisable to use a pipette filler to fill the pipette rather than the mouth?
2. Add the acid from a burette, a little at a time. Then swirl the flask in a controlled way, to allow the acid and alkali to mix see Fig. 4.5b.
3. When all the alkali has been used up, the indicator suddenly turns **colourless,** showing that the solution is neutral. This is the end point. There is no need of adding more acid. Look at Fig. 4.5c.
4. You can tell how much acid was added, using the scale on the burette. Look at Fig. 4.5d. This is the amount of acid that is needed to neutralize 25cm^3 of the alkali.
5. Repeat steps 1 to 4, but this time there is no need for an indicator. 25cm^3 of alkali is put in the flask, and the correct amount of acid added. See Fig. 4.5e.
6. The solution from the flask is heated, to let the water evaporate. See Fig. 4.5f. You will find that dry crystals of sodium chloride are left behind.
7. Record observations in your note book. What do you conclude.
    - What do you think would happen if we evaporated the solution obtained in step 3 to dryness?

- Why was the whole process repeated i.e. steps 1 to 4?
- Using the same procedure describe how you can prepare potassium sulphate. Choose the appropriate reagents.

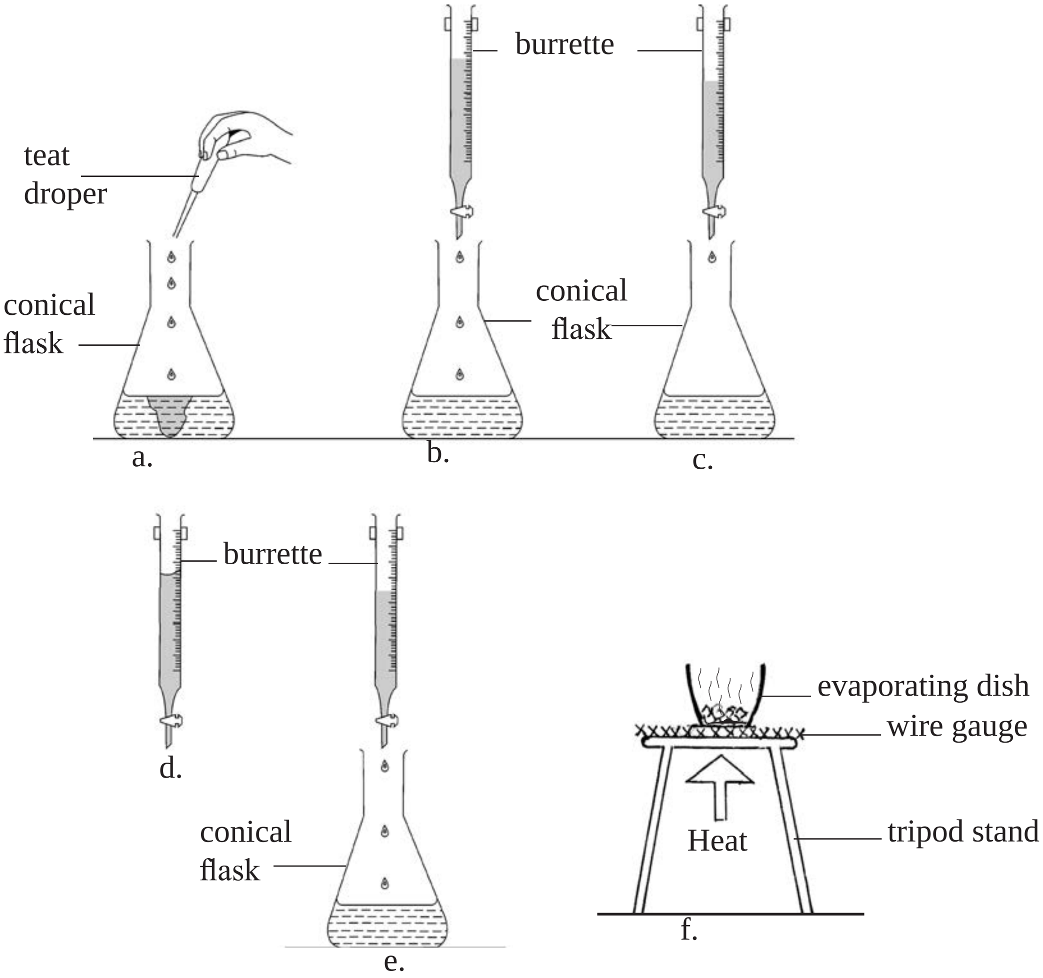


Fig 4.5 Preparation of sodium chloride

When we want to fill a pipette, it is advisable to use a pipette filler to avoid sucking the liquid into the mouth accidentally. This method of preparing a salt, involves the **neutralisation** of an acid with an alkali, for example sodium hydroxide or a soluble carbonate such as sodium carbonate. It is necessary to use an indicator because both the reactants and products are colourless and it would be difficult to know the neutralisation point or end point. If we evaporate the solution with the indicator, our salt would be coloured.

If we do not want to run the burette twice, we can boil the solution with the coloured indicator for 5 minutes with animal charcoal and then filter off the charcoal. The charcoal absorbs the colour of the indicator and the filtrate is colourless. Then we evaporate to get a colourless salt. This method of preparing a soluble salt using a soluble base is known as **neutralisation**. **Neutralisation** is a chemical reaction between a base and an acid to produce a **salt** and **water** only. The reaction that occurred is as follows:

Sodium hydroxide + hydrochloric acid → sodium chloride + water

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

The three alkalis commonly used in this way are; sodium hydroxide, potassium hydroxide and aqueous ammonia (ammonium hydroxide). The following equations represent reactions between the mineral acids and the alkalis to form soluble salts.

Hydrochloric acid + potassium hydroxide → potassium chloride + water

$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$

Sulphuric acid + sodium hydroxide → sodium sulphate + water

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

Nitric acid + ammonium hydroxide → ammonium nitrate + water

$HNO_3(aq) + NH_4OH(aq) \rightarrow NH_4NO_3(aq) + H_2O(l)$

Ionic equation

Hydrogen ion + hydroxide ion → water

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Note: (i) Potassium and sodium hydroxide exist in solid form but 'ammonium hydroxide'exists in solution form.

(ii) Using the mineral acids we can also form sodium, potassium and ammonium salts from their carbonates.

Sodium carbonate + hydrochloric acid → sodium chloride + water + carbon(IV) oxide

$Na_2CO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

Ammonium carbonate + nitric acid → sodium chloride + water + carbon(IV) oxide

$(NH_4)_2CO_3(aq) + 2HNO_3(aq) \rightarrow 2NH_4NO_3(aq) + H_2O(l) + CO_2(g)$

Ionic equation

Carbonate ion + hydrogen ion → water + carbon(IV) oxide

$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

Preparation of insoluble salts by precipitation

The methods we have discussed so far are used to prepare soluble salts. However, some salts are insoluble in water. For example lead(II) sulphate, silver chloride, and barium(II) sulphate among others. Such salts are prepared by a method known as **precipitation or double decomposition.**

Experiment 4.5

To prepare lead(II) sulphate

Apparatus and chemicals

- Beaker or test tube
- Lead nitrate solution
- Sulphuric acid (or any soluble sulphate e.g. sodium sulphate)

Procedure

1. Fill one test tube to about a third with lead nitrate solution and the other with dilute sulphuric acid (or a soluble sulphate).
2. Add dilute sulphuric acid or soluble sulphate to lead nitrate. What happens?
3. Remove the colourless solution by filtration or allow the salt to settle and decant the solution.
4. Wash the salt with hot water to remove impurities.
5. Allow the precipitate to dry on the filter paper, to become a white powder.
6. Record observations in your notebook. What do you conclude?
   - Suggest another substance which you can react with lead(II) nitrate solution to get lead sulphate in place of sulphuric acid.
     Write an equation for the reaction

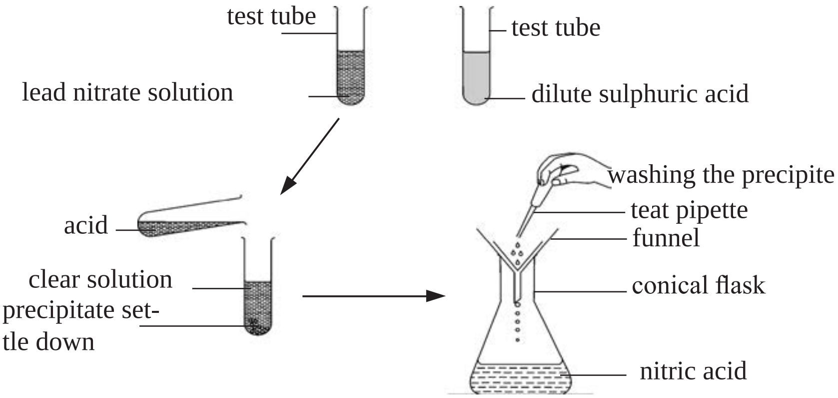


Fig. 4.6 Preparation of lead sulphate

When dilute sulphuric acid or any other soluble sulphate is added to the lead nitrate solution, a white insoluble salt is formed. This insoluble product is also called a **precipitate**. This method of preparing salts is called

precipitation or **double decomposition**. In this method, we start with solutions of substances; therefore both must be soluble in water. When we mix the two solutions, they exchange ions as shown in the following equation.

Lead nitrate + hydrogen sulphate → lead sulphate + hydrogen nitrate
(solution) (solution) (precipitate) (remains in solution)

$Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2HNO_3(aq)$

Note: A sulphuric acid is called hydrogen sulphate and nitric acid hydrogen nitrate. This is to show how hydrogen ions are exchanged. A solution of any sulphate could have been used instead of sulphuric acid for this preparation; for example sodium sulphate.

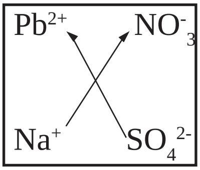
Lead nitrate + sodium sulphate → lead sulphate + sodium nitrate

$Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$

The ions present in the solutions are

From $Pb(NO_3)_2(aq)$ →

From $Na_2SO_4(aq)$ →



ions exchanged to give products $PbSO_4$ and $NaNO_3$

Ionic equation

When we write an ionic equation, we select the ions that form what we see; in this case the precipitate. We leave out the ions that remain in solution. These are known as **spectator ions**. Just as in a match, a spectator does not participate in the game.

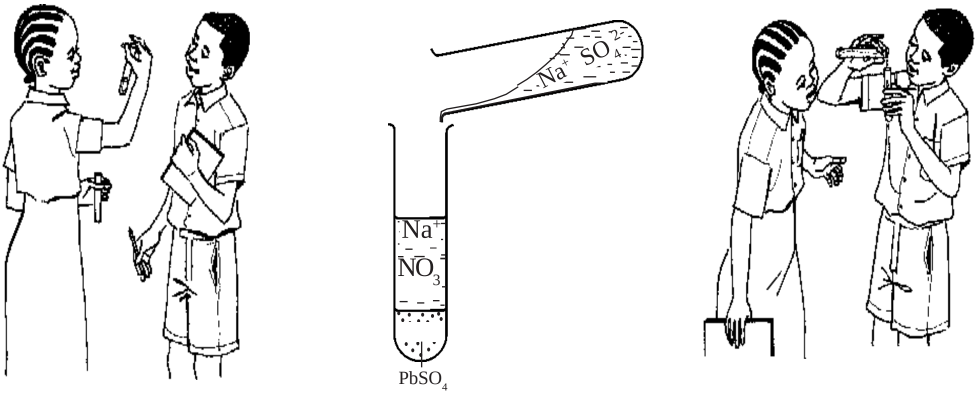


Fig. 4.7 Mixing two solutions

For instance, in the reaction above, Fig 4.7, the precipitate is lead sulphate. The ionic equation for this reaction would be written as:

$$\text{Lead(II) ions} + \text{sulphate ions} \rightarrow \text{lead sulphate}$$
$$Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$$

Other insoluble salts which can be prepared by double decomposition are: barium sulphate and lead chloride. Lead(II) chloride is a precipitate that is soluble in hot water.

$$\text{Barium chloride} + \text{sodium sulphate} \rightarrow \text{barium sulphate} + \text{sodium chloride}$$
$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

$$\text{Barium ion} + \text{sulphate ion} \rightarrow \text{barium sulphate}$$

Ionically: $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

$$\text{Lead(II) nitrate} + \text{sodium chloride} \rightarrow \text{lead(II) chloride} + \text{sodium nitrate}$$
$$Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$$

$$\text{Lead(II) ion} + \text{chloride ion} \rightarrow \text{lead(II) chloride}$$

Ionically: $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

soluble in hot water

Note: In neutralisation reactions when an acid is mixed with an alkali the solution resembles water. Therefore when we write an ionic equation we choose the ions that form water as follows.

$$\text{Hydrogen ion} + \text{hydroxide ion} \rightarrow \text{water}$$
$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Repeat Experiment 4.5 with the following pairs of aqueous solutions in Table 4.1. Note the colour of the precipitates and write balanced equations in each case followed by ionic equations.

Table 4.1 Formation of some coloured precipitates

Solution of salt X	**Solution of substance Y**	**Record colour of precipitate if any**
Lead nitrate	Potassium iodide	
Copper(II) sulphate	Sodium hydroxide	
Iron(II) sulphate	Sodium hydroxide	
Iron(II) chloride	Sodium hydroxide	
Sodium chloride	Silver nitrate	
Lead nitrate	Ammonium chloride	
Ammonia solution	Zinc nitrate	

Precipitation reactions are used in analysis of salts to identify the cations and anions in the salt. We will do this later in the course.

Writing ionic equations

Reactions that take place between ionic compounds dissolved in water can be recorded in a much simpler manner. The ions are free to move and react. When a chemical reaction takes place between ionic compounds, it may happen that only certain ions undergo changes while others do not.

You will remember from Unit 3 that ionic compounds are made of a lattice of ions that are rigidly fixed in solid form. When ionic compounds are dissolved in water, the ions are set free.

Knowledge of solubility of salts in water is very crucial when writing ionic equations. This assists very much so that correct symbols can be identified.

Note: Solubility of salts will be dealt with in detail later in this unit.

However, the following is a summary of the solubility of various salts and bases.

- All potassium (K), sodium (Na) and ammonium (NH_4) salts are soluble.
- All nitrates and hydrogencarbonates (HCO_3) are soluble.
- All sulphates are soluble except lead sulphate ($PbSO_4$) and barium sulphate ($BaSO_4$). Calcium sulphate ($CaSO_4$) is slightly soluble.
- All chlorides are soluble except silver chloride (AgCl). Lead(II) chloride ($PbCl_2$) is only soluble in hot water.
- All carbonate salts are insoluble except potassium, sodium and ammonium carbonates.

Example 1

When sodium carbonate solution is mixed with barium chloride solution, a reaction takes place and a white precipitate is formed. Fig 4.8

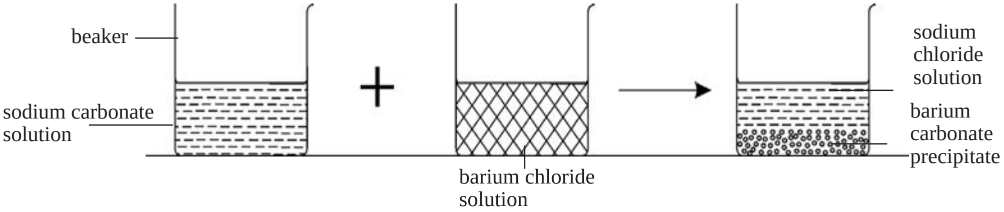


Fig. 4.8 Precipitation reaction

The balanced chemical equation for this reaction is as follows.

Sodium carbonate + barium chloride → sodium chloride + barium carbonate

$$Na_2CO_3(aq) + BaCl_2(aq) \rightarrow 2\,NaCl(aq) + BaCO_3(s)$$

Before the solutions were mixed, the ions present were:
From sodium carbonate solution ($Na_2CO_3(aq)$); $2Na^+$ and CO_3^{2-}
From barium chloride solution ($BaCl_2(aq)$); Ba^{2+} and $2Cl^-$
When the solutions are mixed, Ba^{2+} ions unite with the CO_3^{2-} ions to form $BaCO_3$ which is an insoluble white precipitate. Na^+ and Cl^- ions undergo no changes.

An ionic equation shows only those ions that change in some way during a reaction. From the above example, the ions that do not undergo any change are omitted from the ionic equation.
There are three possibilities during ionic changes:

1. Formation of a **precipitate.**
2. Evolution of a **gas.**
3. Formation of **water.**

Table 4.2 shows some substances that contain free ions and those that do not.

Table 4.2 Substances containing free and no free ions

Free ions	**No free ions**
Metal salts in solution	Insoluble solids including precipitates
Ammonium salts in solution	Covalent compounds (HCl, NH_3, N_2, O_2 etc)
Gases in solution (HCl(aq))	Uncombined elements (O, N, P, S etc)
Dilute acids (HCl, HNO_3, H_2SO_4 etc)	Gases (O_2, N_2, CO_2, SO_2, etc)
Soluble metal alkalis (NaOH, KOH, etc)	

Note: Solids, liquids, gases do not have free ions.
An ionic equation can be written following certain rules. Let us follow these rules to write the ionic equation of our previous example.

Example 1

Write an ionic equation for the reaction of sodium carbonate and barium chloride solutions.

Step 1

Write down well balanced stoichiometric or the normal chemical equation for the reaction

Sodium carbonate + barium chloride → sodium chloride + barium carbonate

$$Na_2CO_3(aq) + BaCl_2(aq) \rightarrow 2NaCl(aq) + BaCO_3(s)$$

Step 2

Decide which substances are ionic and which ions are free in solution.

Omit the spectator ions by crossing the ion that appears on both sides of the equation.

$$Ba^{2+}(aq) + \cancel{2Cl^-}(aq) + \cancel{2Na^+}(aq) + CO_3^{2-}(aq) \rightarrow \cancel{2Na^+}(aq) + \cancel{2Cl^-} + BaCO_3(s)$$

Step 3

Rearrange the ions and rewrite the equation in ionic form.

$$Ba^{2+}(aq) + CO^{2-}_3(aq) \rightarrow BaCO_3(s)$$

Example 2

Write the ionic equation of iron(III) chloride with sodium hydroxide.

Step 1

Write down balanced stoichiometric or the normal chemical equation for the reaction

Iron(III) chloride + sodium hydroxide → sodium chloride + iron(III) hydroxide

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow 3NaCl(aq) + Fe(OH)_3$$

Step 2

Decide which substances are ionic and which ions are free in solution. Omit the spectator ions by crossing the ions that appears on both sides of the equation.

Note: Remember that spectator ions are those ions that remain unchanged during the chemical reaction. They appear on both sides of the equation.

$$Fe^{3+}(aq) + \cancel{3Cl^-(aq)} + \cancel{3Na^+(aq)} + 3OH^-(aq) \rightarrow \cancel{3Na+(aq)} + \cancel{3Cl^-(aq)} + Fe(OH)_3(s)$$

Step 3

Rearrange the ions and rewrite the equation in ionic form

$$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$$

Preparation of salts by direct combination of elements

Some salts can be made by directly reacting together the two elements that make them. This process is called **direct combination** or **synthesis.** The two elements combine to produce one substance. Usually energy in form of heat is required to make the two elements combine chemically.

Experiment 4.6

Preparation of sodium chloride by direct combination of sodium and chlorine

Apparatus and chemicals

- deflagrating spoon
- gas jar full of dry chlorine
- knife/scalpel
- tongs
- bunsen burner
- sodium metal

Procedure

1. Remove sodium from the storage container using a pair of tongs.
2. Cut a small piece of sodium metal and place it in a deflagrating spoon.
3. Heat the piece of metal over the Bunsen flame.
4. When the metal catches fire, lower the burning sodium into a gas jar of dry chlorine.
5. Record your observation.

The reaction is violent. A white powder of sodium chloride is formed.

Sodium + chlorine → sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Some other metals also combine with chlorine to form anhydrous salts. For example, when chlorine is passed over heated iron wire, iron combine with chlorine to form iron (III) chloride.

Iron + chlorine → iron(III) chloride

$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$

See reactions of halogens with metals

Types of salts

We have mentioned that a salt is formed when the hydrogen of an acid is replaced directly or indirectly by a metal or ammonium radical.

The basicity of acids

Hydrochloric acid, (HCl), and nitric acid, (HNO_3), contain one replaceable hydrogen atom per molecule. They are said to be *monobasic acids*. Sulphuric acid, (H_2SO_4) and carbonic acid (H_2CO_3), have two hydrogen atoms that can be replaced by a metal. They are said to be **dibasic acids**. Phosphoric acid, (H_3PO_4) is a *tribasic acid*. Why?

When acids are dissolved in water, they produce hydrogen ions, (H^+). The hydrogen ions are also called protons. An acid can be defined as a **proton donor**. Therefore a monobasic acid is also known as **monoprotic** acid. A dibasic acid diprotic and a tribasic acid can be called **triprotic,** See Table 4.3:

We represent hydrogen ions in aqueous acid solution as follows.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$
$$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$
$$H_3PO_4(aq) \rightarrow 3H^+(aq) + PO^{3-}{}_4(aq)$$

Table 4.3 The basicity of some common acids

Name	Formula	Number of protons H^+ given in aqueous solution	Basicity of acid
Hydrochloric acid	HCl	1	Monobasic (monoprotic)
Nitric acid	HNO_3	1	
Sulphuric acid	H_2SO_4	2	Dibasic (diprotic)
Sulphurous acid	H_2SO_3	2	
Carbonic acid	H_2CO_3	2	
Phosphoric(V) acid	H_3PO_4	3	Tribasic (triprotic)

Dibasic acids and acid salts

When an acid can produce more than one hydrogen ion per molecule, as in the case of sulphuric acid, it is possible for the replacement of H^+ by a metallic ion to occur in stages as follows:

(a) Sulphuric acid + sodium hydroxide → sodium hydrogensulphate + water

$$H_2SO_4(aq) + NaOH(aq) \rightarrow NaHSO_4(aq) + H_2O(l)$$

(b) Sodium hydrogensulphate + sodium Hydroxide → sodium sulphate + water

$$NaHSO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

In the first reaction only one H^+ is replaced by the ion Na^+ to give sodium hydrogensulphate. In the second equation similar replacement gives sodium sulphate (Na_2SO_4).

Salts like sodium hydrogensulphate ($NaHSO_4$) when dissolved in water can donate a proton, (H^+), just as acids do, hence are called **acid salts**.

Sodium hydrogensulphate → sodium ion + hydrogen ion + sulphate ion

$$NaHSO_4 \rightarrow Na^+(aq) + H^+(aq) + SO_4^{2-}(aq)$$

Note: Some acid salts like sodium hydrogencarbonate, ($NaHCO_3$), dissolve in water to give an alkaline solution. It still qualifies to be called an acid salt because of the presence of the hydrogen atom which can be replaced by a metal.

Salts like sodium sulphate, do not contain hydrogen. Consequently when dissolved in water cannot give H^+ ions. Such salts are known as **normal salts**.

Only dibasic and tribasic acids can form acid salts. Look at Table 4.4. Monobasic acids like hydrochloric acid and nitric acid form normal salts only.

Table 4.4 The basicity of some common acids and salts they form

Name	Formula	Basicity	Normal salts	General names of salts	Acid salts	General names of salts
Hydrochloric acid	HCl	Monobasic	$NaCl$	Chlorides	-	-
Nitric acid	HNO_3	Monobasic	$NaNO_3$	Nitrates	-	-
Sulphurous acid	H_2SO_3	Dibasic	Na_2SO_3	Sulphites	$NaHSO_3$	Hydrogensulphites
Sulphuric acid	H_2SO_4	Dibasic	Na_2SO_4	Sulphates	$NaHSO_4$	Hydrogensulphates
Carbonic acid	H_2CO_3	Dibasic	Na_2CO_3	Carbonates	$NaHCO_3$	Hydrogencarbonates
Phosphoric(V) acid	H_3PO_4	Tribasic	Na_3PO_4	Phosphates	NaH_2PO_4 and Na_2HPO_4	Hydrogenphosphates

Double salts

Some salts exist as double salts, for example **trona** which occurs naturally in lake Magadi. Trona contains both sodium carbonate, (Na_2CO_3) and sodium hydrogencarbonate, ($NaHCO_3$). It is a hydrated salt with the following formula.

$$Na_2CO_3.NaHCO_3.2H_2O$$

Experiment 4.7

To find out what happens when certain salts are left exposed in the air.

Apparatus and chemicals

- watch glasses/white tile
- anhydrous copper(II) sulphate
- anhydrous calcium chloride
- iron(III) chloride
- sodium nitrate
- hydrated sodium carbonate
- hydrated sodium sulphate

Procedure

1. Place a small amount of each of the above salts on separate watch glasses.

2. Copy Table 4.5. Observe and record the appearance of each salt.
3. Leave the samples exposed in air for about 3-7 days.
4. Observe and record any changes that may occur after a few minutes, one day, one week or more.
5. Suggest a reason for the change in appearance, if any.

Table 4.5: Record observations made when salts are exposed in air

Salt	**Appearance before staying overnight**	**Appearance after staying overnight**	**Reason for change**
(a) Anhydrous copper(II) sulphate			
(b) Anhydrous calcium chloride			
(c) Iron(III) chloride			
(d) Hydrated sodium carbonate			

From your results you will notice that some changes have occurred in the salts. Some salts have absorbed water from the air while others have lost water into the air. Many substances absorb water from the air. For example ordinary impure sodium chloride (common salt) becomes damp when exposed to air. Impure salt contains magnesium chloride which absorbs water from the air. Pure salt does not absorb water. White anhydrous copper(II) sulphate turns blue when exposed to air.

Substances that absorb water from the air but not enough to form a solution are said to be **hygroscopic**. The process of absorbing water is known as **hygroscopy.** Examples of hygroscopic substances are listed in Table 4.6. Some salts e.g. calcium chloride absorb so much water from the air that it dissolves in it. Substances like this are said to be **deliquescent**. The process is known as **deliquescence**.

Note: All deliquescent substances are hygroscopic, but all hygroscopic substances are not necessarily deliquescent. This means that not all substances that absorb water from the air dissolve in it.

Substances that lose water to the air are said to **efflorescent**. The process is called **efflorescence**. For example, a crystal of hydrated sodium carbonate, when left exposed to air, slowly forms a white powder on its surface and finally the crystal changes to powder. In other words it loses 9 of its 10 water molecules of crystallization to the air as shown in the following equation.

Sodium carbonate-10-water → Sodium carbonate-1-water + 9-water

$$Na_2CO_3.10H_2O(s) \rightarrow Na_2CO_3.H_2O(s) + 9H_2O(g)$$

Table 4.6: Some examples of hygroscopic, deliquescent and efflorescent salts

Hygroscopic salts	Deliquescent salts	Efflorescent salts
Anhydrous copper(II) sulphate	Sodium nitrate	Sodium carbonate-10-water
	Calcium chloride Magnesium chloride	Sodium sulphate-10-water
All deliquescent substances	Iron(III) chloride	Iron(II) sulphate-7-water

4.2 Solubility of salts

When we dissolve substances such as sugar and common salt in water, we cannot see them in the colourless liquid. But we know they are there because we can taste the sugar or the salt in water. The solid has been broken into tiny particles which we cannot see with the naked eye, nor with the most powerful microscope. The particles will pass through the pores of a filter paper. These particles and the water form a homogeneous mixture called a **solution**.

Substances such as salt and sugar which dissolve in water are described as **soluble**. Solids such as sand and soil which do not dissolve are said to be **insoluble**. We say that a solution which contains only a little solute in a given amount of solvent is **dilute**. But if a solution contains a lot of solute in a given amount of solvent we say it is **concentrated**. If we evaporated more and more solvent, a solution becomes more and more concentrated. Eventually, the solution reaches a point that it cannot become any more concentrated while remaining a solution. If any more solvent is evaporated, some solid solute will come out of solution and form crystals. This is because the solution is **saturated**.

Experiment 4.8

To make a saturated solution

Apparatus and chemicals

- beaker
- stirrer/glass rod
- hydrated copper (II) sulphate or potassium nitrate or sodium chloride.

Procedure

1. Place about $20cm^3$ of water in a beaker.
2. Measure a spatula endful of copper(II) sulphate, add to the water and stir well with the glass rod.
3. Observe what happens.
4. Add other portions of the solid and each time stir well.
5. What happens as more and more solid is added?
6. Warm the solution as you stir. What do you observe?

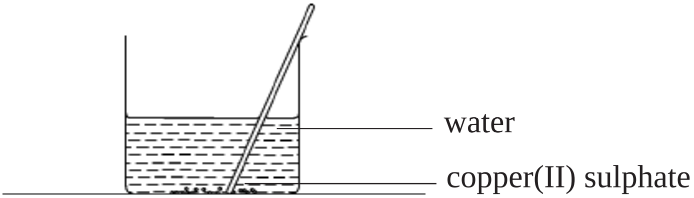


Fig 4.9 Making a saturated solution of copper (II) sulphate

The first portion dissolves and a faint blue solution is formed. As more and more portions are added and dissolved, the blue colour deepens. As further portions are added, a point is reached when no more copper(II) sulphate will go into solution.

In other words the solution cannot dissolve anymore solid. We say that the solution has become **saturated.** The only way we can tell a solution is saturated, is when we see some undissolved solute. If the solution is warmed the undissolved solid dissolves. The hot solution dissolves more solute. This indicates that the solubility of solid increases with rise in temperature. A *saturated solution* is therefore a solution which contains as much solute as can be dissolved at a **particular temperature,** and has some **undissolved solute**.

The amount of substance (in grams) which dissolves in a specified amount (mass) of solvent e.g water, at a particular temperature to form a saturated solution is known as its **solubility.**

Experiment 4.9

To investigate solubility of some salts

Apparatus and chemicals

- test tube rack
- test tubes
- potassium nitrate
- sodium carbonate
- barium sulphate
- zinc carbonate
- sodium chloride
- sodium nitrate
- ammonium sulphate
- lead nitrate
- copper nitrate
- zinc nitrate

- copper carbonate
- copper(II) hydroxide
- sodium hydroxide
- barium chloride
- calcium carbonate
- potassium nitrate
- magnesium sulphate
- copper(II) sulphate
- silver nitrate
- silver chloride
- lead carbonate

Procedure

1. Take a spatula endful of each of the salts listed above and put it in a test tube.
2. Add cold water to about half the test tube. (Do not warm).
3. Close the test tube with your thumb and shake well. What do you observe?
4. Place the test tube on the rack and observe what happens after a shortwhile.
5. Record observations in Table 4.7. What do you conclude?

Copy Table 4.7 in your notebook and put a tick(√) where salt is soluble and a cross (x) where the salt is insoluble.

Table 4.7 Some soluble and insoluble salts

Name of salt	Soluble	Insoluble
Potassium chloride Sodium nitrate Ammonium sulphate Sodium carbonate		
Lead nitrate Copper nitrate Zinc nitrate Magnesium sulphate Calcium sulphate Barium sulphate Silver chloride Barium chloride Calcium carbonate Lead carbonate Zinc carbonate		

From your results you can see that some of the salts are soluble in water whereas others are insoluble. Table 4.8 shows the solubility patterns for various types of salts.

We must learn and remember the soluble and insoluble salts because this will assist us in indicating the state symbols when writing equations especially during precipitation. As mentioned earlier, precipitation is a method used to prepare insoluble salts from solutions of two soluble salts. For example:-

Sodium sulphate + lead nitrate → lead sulphate + sodium nitrate

$$Na_2SO_4(aq) + Pb(NO_3)_2(aq) \rightarrow PbSO_4(s) + 2NaNO_3(aq)$$

If you look on the products side, lead sulphate, $PbSO_4(s)$ is given the solid state because in Table 4.8 it is among the insoluble sulphates. Sodium nitrate, $NaNO_3(aq)$ is given state (aq) because all nitrates and sodium salts are soluble. We do not prepare nitrates and other soluble salts by precipitation.

Table 4.8 Solubility patterns of various types of salts

Salt	**Soluble**	**Insoluble**
Potasium salts	All are soluble	None is insoluble
Sodium salts	All are soluble	None is insoluble
Ammonium salts	All are soluble	None is insoluble
Chlorides	All are soluble hot water)	Except silver chloride, Lead(II) chloride (soluble in
Nitrates	All are soluble	None is insoluble
Carbonates	Only K, Na and NH are soluble	Others are insoluble
Sulphates	All are soluble	Except lead sulphate, barium sulphate, but calcium sulphate is slightly soluble

See summary of the salts on page 142.

4.3 Action of heat on salts

As we have seen salts are compounds which contain metal or ammonium ions which, are positively charged (cations) and acid radicals which are negatively charged (anions). These oppositely charged ions attract each other strongly. These strong electrostatic forces of attraction hold the ions together. In some salts the forces of attraction are so strong that application of heat cannot break them. However, in some cases strong heating decomposes some salts and we get new products.

Action of heat on carbonates and hydrogencarbonates

Carbonates and hydrogencarbonates are salts of carbonic acid, H_2CO_3. There are two solid hydrogencarbonates, namely, potassium hydrogencarbonate ($KHCO_3$) and sodium hydrogencarbonate ($NaHCO_3$).

The other known hydrogencarbonates exist only in solution. These are:

- calcium hydrogencarbonate, $Ca(HCO_3)_2$.
- magnesium hydrogencarbonate, $Mg(HCO_3)_2$

All metals form solid metal carbonates. However, aluminium and iron(III) carbonates do not exist.

The effect of heat on metal compounds can be linked to the reactivity series like the reaction of metals themselves. We are now going to investigate the action of heat on some carbonates and hydrogencarbonates.

Experiment 4.10

To investigate the action of heat on carbonates and hydrogencarbonate.

Apparatus and chemicals

- Bunsen burner
- hard glass test tubes
- spatula
- carbonates and hydrogencarbonates (as listed in the procedure)
- test tube holder
- calcium hydroxide solution (lime water)

Procedure

1. Put into different dry test tubes about one spatula end of each of the following salts: Potassium carbonate, sodium carbonate, magnesium carbonate, zinc carbonate, copper carbonate, iron(II) carbonate, lead carbonate, potassium hydrogencarbonate and sodium hydrogencarbonate.
2. Note the colour of the salt before heating.
3. Heat the sample and test the gas produced with calcium hydroxide solution as shown in Fig 4.10. Copy Table 4.9 and record your observation.
4. Finally note the colour of the residue when hot and when it has cooled.
5. List the compounds which bring a change in calcium hydroxide solution when they are heated. These compounds will have decomposed.

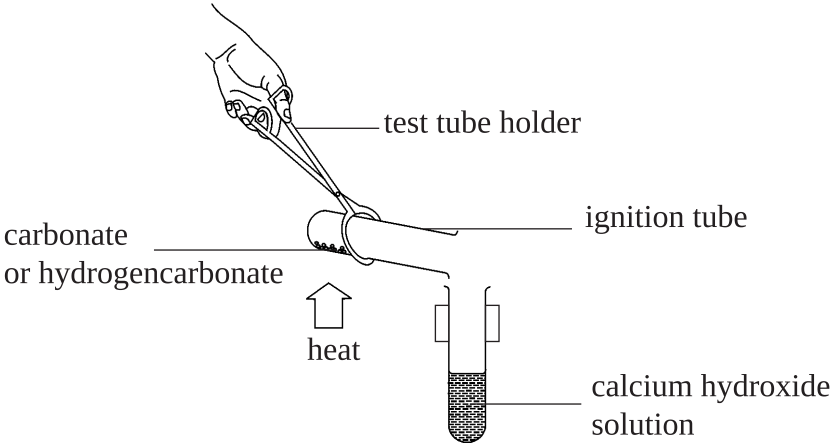


Fig 4.10: Testing decomposition of a carbonate/hydrogencarbonate on heating

The compounds that are coloured are copper carbonate and iron(II) carbonate. They are green solids. All the others are white solids.

Table 4.9 Decompositon of carbonates and hydrogencarbonates

Carbonate/hydrogen carbonate	Colour before heating	Colour of the residue	Effect on calcium hydroxide solution
Potassium carbonate			
Sodium carbonate			
Zinc carbonate			
Lead(II) carbonate			

From results we recorded in Table 4.9, we notice that some carbonates are not decomposed by heat (there is no effect on calcium hydroxide solution. Potassium carbonate and sodium carbonate are not decomposed by heat).

Other metal carbonates decompose on heating to respective metal oxides and carbon(IV) oxide.

The carbon(IV) oxide produced reacts with calcium hydroxide solution forming a white precipitate of calcium carbonate as follows:

Calcium hydroxide + carbon(IV) oxide → calcium carbonate + water

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

The tiny particles of the precipitate are suspended in the liquid.

Generally:

$$\text{Metal carbonate} \xrightarrow{\text{heat}} \text{metal oxide} + \text{carbon(IV) oxide}$$

Examples

Zinc carbonate → zinc oxide + carbon(IV) oxide
$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$
(white) (yellow hot
white cold)

Lead carbonate → lead(II) oxide + carbon(IV) oxide
$PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$
(white) (red-brown hot
yellow cold)

Copper carbonate → copper(II) oxide + carbon(IV) oxide
$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
(green) (black)

Magnesium carbonate → magnesium oxide + carbon(IV) oxide
$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$
(white) (white)

Three products are formed when hydrogencarbonates are decomposed by heat.

For instance, potassium hydrogencarbonate and sodium hydrogencarbonate decompose to **metal carbonate, water** and **carbon(IV) oxide**.

Potasium hydrogencarbonate → potasium carbonate + water + carbon(IV) oxide
$2KHCO_3(s) \rightarrow K_2CO_3(s) + H_2O(g) + CO_2(g)$

Sodium hydrogencarbonate → sodium carbonate + water + carbon(IV) oxide
$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$
(white) (white)

However, ammonium carbonate when heated decomposes into two gases namely ammonia, carbon(IV) oxide and water vapour. The water vapour condenses on the cold part of the test tube.
It can be tested with anhydrous cobalt(II) chloride paper or anhydrous copper(II) sulphate.

Ammonium carbonate → ammonia + carbon(IV) oxide + water
$(NH_4)_2CO_3(s) \rightarrow 2NH_3(g) + CO_2(g) + H_2O(l)$

Note: We mentioned earlier that sodium carbonate is not decomposed by heat.

The hydrated sodium carbonate when heated loses water of crystallisation as shown in the following equation.

$$\text{Hydrated sodium carbonate} \rightarrow \text{sodium carbonate} + \text{water}$$
$$Na_2CO_3.10H_2O(s) \rightarrow Na_2CO_3(s) + 10H_2O(g)$$

The reactivity series helps to explain the action of heat on carbonates. Potassium and sodium are very reactive metals (top of the reactivity series). Hence as mentioned earlier their carbonates do not decompose on heating. Calcium and magnesium follow in the series. Their carbonates decompose but at a higher temperature. The other carbonates lower in the series decompose easily. For example copper metal is near the bottom of the series; copper carbonate decomposes easily on heating.

Experiment 4.11

To investigate the effect of heat on nitrates

Apparatus and chemicals

- ignition tubes
- burner
- wood splint
- test tube holder
- sodium nitrate
- potassium nitrate
- calcium nitrate
- zinc nitrate
- lead nitrate
- magnesium nitrate
- anhydrous copper(II) sulphate or dry cobalt chloride paper

Procedure

1. Place a spatula end-full of sodium nitrate into an ignition tube.
2. Note the colour of the nitrate and heat the sample strongly.
3. Insert a glowing splint into the test tube and note what happens.
4. Copy Table 4.10 in your notebook and record your observations.
5. If any liquid condenses in the test tube, test with anhydrous copper (II) sulphate or dry cobalt chloride paper.
6. Repeat steps 1 - 4 using the rest of the nitrates listed above.

Table 4.10 Action of heat on nitrates

Name of nitrate used	Colour before	Change of colour during heating (if any)	Colour of gas heating	Effect on glowing splint produced	Effect on dry cobalt chloride paper
Sodium nitrate					
Potassium nitrate					
Calcium nitrate					
Zinc nitrate					
Lead nitrate					
Magnesium nitrate					

- List the nitrates which behaved the same in this experiment.
- Which nitrates evolved a gas that relights a glowing splint?
- Which nitrates produced residues that showed colour changes?
- Which nitrates were easily decomposed?
- Which nitrate makes a lot of noise when heated?

From the observations in Table 4.10 we notice that when nitrates are heated, they decompose forming different products. Some nitrates produce a colourless gas which rekindles a glowing splint. Others produce a red-brown gas and another which re-kindles a glowing splint.

Let us classify the nitrates according to products formed when they are heated.

1. Nitrates that form two products only
   - Sodium nitrate
   - Potassium nitrate
   - Ammonium nitrate

Sodium nitrate and potassium nitrate melt to form colourless liquid. On further heating, they slowly decompose forming the nitrite and oxygen as shown in the following equation.

e.g. Metal nitrate $\xrightarrow{\text{heat}}$ metal nitrite + oxygen

Sodium nitrate → sodium nitrite + oxygen

$$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$$

Potassium nitrate → potassium nitrite + oxygen

$$2KNO_3(s) \rightarrow 2KNO_2(s) + O_2(g)$$

Note: Ammonium nitrate melts to a colourless liquid, which decomposes to form nitrogen(I) oxide and water.

Ammonium nitrate → nitrogen(I) oxide + Water

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$$

Caution: Do not heat ammonium nitrate in the laboratory because when a little of the nitrate is left, it explodes on strong heating.

2. Nitrates that decompose to form three products
These nitrates decompose to the metal oxide, nitrogen(IV) oxide, (red-brown gas) and oxygen.
- Calcium nitrate
- Magnesium nitrate
- Zinc nitrate
- Copper nitrate
- Lead nitrate

Note:
- All the above nitrates are hydrated except lead nitrate.
- All the crystals are white except copper nitrate.

When the above metal nitrates are heated, they decompose as shown in the following word equation.

Metal nitrate → Metal oxide + nitrogen(IV) oxide + oxygen

Examples

Calcium nitrate → Calcium oxide + nitrogen(IV) oxide + oxygen

$$2Ca(NO_3)_2(s) \rightarrow \underset{\text{(white)}}{2CaO(s)} + 4NO_2(g) + O_2(g)$$

Magnesium nitrate → magnesium oxide + nitrogen(IV) oxide + oxygen

$$2Mg(NO_3)_2(s) \rightarrow \underset{\text{(white)}}{2MgO(s)} + 4NO_2(g) + O_2(g)$$

Zinc nitrate → zinc oxide + nitrogen(IV) oxide + oxygen

$$\underset{\text{(white)}}{2Zn(NO_3)_2(s)} \rightarrow \underset{\text{(yellow when hot; white when cold)}}{2ZnO(s)} + 4NO_2(g) + O_2(g)$$

Copper nitrate → Copper(II) oxide + Nitrogen(IV) oxide + Oxygen
$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$
(blue) (black)

Lead nitrate

When lead nitrate is heated, gas forms inside the crystals and splits them. This **crackling** noise is called **decrepitation**. Otherwise lead nitrate forms the same products like the ones above.

Lead nitrate → Lead(II) oxide + Nitrogen(IV) oxide + Oxygen
$2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$

Note: We have mentioned that lead nitrate is formed in anhydrous form. The other nitrates are hydrated. When we heat the hydrates we drive off water and on further heating they decomposes just like lead nitrate.

e.g. hydrated metal nitrate → metal nitrate + water
$Cu(NO_3)_2.3H_2O(s) \rightarrow Cu(NO_3)_2(s) + 3H_2O(l)$

3. Nitrates that form three products, namely a free metal, nitrogen (IV) oxide, oxygen are:
   - Silver nitrate
   - Mercury nitrate

Why do these two nitrates form free metals when heated?

Silver nitrate and mercury nitrate form metal oxides, nitrogen(IV) oxide and oxygen just like the other heavy metal nitrates. But silver oxide, and mercury oxide when heated usually decompose to free metal and oxygen. So the oxide formed during decomposition of the nitrate will decompose to the metal and oxygen.

Silver nitrate $\xrightarrow{\text{heat}}$ silver + nitrogen(IV) oxide + oxygen
$2AgNO_3(s) \rightarrow 2Ag(s) + 2NO_2(g) + O_2(g)$
(silver)

Mercury nitrate → mercury + nitrogen(IV) oxide + oxygen
$Hg(NO_3)_2(s) \rightarrow Hg(l) + 2NO_2(g) + O_2(g)$
(mercury)

The action of heat on nitrates vary according to the position of the metal in the reactivity series.

K Na	On heating nitrates of these metals we get the nitrite and oxygen.
Ca Mg Al Fe Pb Cu	On heating nitrates of these metals we get the metal oxide, nitrogen(IV) oxide and oxygen.
Hg Ag	Nitrates of these metals decompose to the metal, nitrogen(IV) oxide and oxygen.

Action of heat on sulphates

Experiment 4.12

To investigate the effect of heat on sulphates

Apparatus and chemicals

- Ignition tubes
- Bunsen burner
- Litmus papers
- Test tube holder
- Hydrated iron(II) sulphate
- Hydrated copper(II) sulphate

Procedure

1. Place a little of the sulphate at a time in the ignition tube.
2. Heat gently first and record your observation.
3. Then heat strongly and test any gases produced with litmus papers. Record your observations.
4. Record the initial and final colour of the residue.
5. Repeat the experiment with the other sulphates.

Hydrated salts first lose their water of crystalisation. Then the anhydrous salt decomposes on very strong heating.

For example

Iron(II) sulphate-7-water → iron(II) sulphate + water

$$FeSO_4.7H_2O(s) \rightarrow FeSO_4(s) + 7H_2O(g)$$

Then

Anhydrous iron(II) sulphate → iron(III) oxide + sulphur(IV) oxide + sulphur(VI) oxide

$$2FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$$

Blue hydrated copper(II) sulphate will also first lose water of crystallisation. It will form white anhydrous copper(II) sulphate.

Copper(II) sulphate-5-water → anhydrous copper(II) sulphate + 5-water

$$\underset{\text{(blue)}}{CuSO_4.5H_2O(s)} \rightarrow \underset{\text{(white)}}{CuSO_4(s)} + 5H_2O(g)$$

On further heating, anhydrous copper(II) sulphate decomposes to copper(II) oxide and sulphur(VI) oxide.

Anhydrous copper(II) sulphate → copper(II) oxide + sulphur(VI) oxide

$$\underset{\text{(white)}}{CuSO_4(s)} \rightarrow \underset{\text{(black)}}{CuO(s)} + SO_3(g)$$

Uses of some salts

1. *In controlling soil pH and plant growth.*
   Plant growth is affected by the alkalinity or acidity of the soil. The soil pH is affected by the use of fertilisers and the acid rain. If the soil is too acidic, the effect can be neutralised by calcium carbonate or calcium hydroxide. Calcium oxide is also used to neutralise soil acidity
2. *Fertilizers*
   The following salts are used as fertilisers.
   - Ammonium nitrate
   - Ammonium sulphate
   - Ammonium phosphate
   - Sodium nitrate
3. *Indigestion* can be relieved by salts.
   Examples are calcium carbonate and sodium hydrogencarbonate.
4. *Baking*
   Sodium hydrogencarbonate is used in baking. When heated it produces carbon (IV) oxide which causes bread to rise.
5. *Softening of hard water*
   Sodium carbonate is used to soften hard water.

6. *Melting ice on roads*
   Sodium chloride and other salts are poured on roads to lower the freezing point of ice in countries that experience winter.
7. *Extraction of metals*
   Calcium chloride is used in the extraction of sodium metal.

Summary

1. A salt is formed when the hydrogen of an acid is directly or indirectly replaced by a metal or ammonium radical.
2. The name of a salt starts with the name of the metal or ammonium radical.
3. Soluble salts can be prepared by reacting metals with dilute acids, alkalis with dilute acids, bases with acids and carbonates or hydrogencarbonates with acids; and direct synthesis.
4. Insoluble salts are prepared by precipitation.
5. Salts which contain water of crystallization are said to be **hydrated** while those without water of crystallization are said to be **anyhydrous**.
6. A salt that absorbs water from the air but not enough to form a solution, is said to be **hygroscopic** e.g. impure sodium chloride contains calcium chloride or magnesium chloride which are hygroscopic.
7. A salt that absorbs enough water until it dissolves in it is said to be **deliquescent**.
8. A salt that gives its water of crystallization to the atmosphere is said to be **efflorescent**.
9. Basicity of an acid refers to the number of ionisable hydrogen atoms per molecule e.g. the dibasic acid, H_2SO_4 contains two hydrogen atoms.
10. Monobasic acids form normal salts only.
11. Dibasic and tribasic acids form normal and acid salts.
12. A normal salt is formed when all hydrogens of dibasic or tribasic acids are replaced by a metal or an ammonium radical.
13. An acid salt is formed when only part of the hydrogen of dibasic or tribasic acids is replaced by a metal or an ammonium radical.

Project

- Analysis of anti-acid drugs.

Revision Exercise 4

1. What is a salt?
2. Briefly explain the difference between the following:
   (a) acid salts and normal salts
   (b) hydrated salts and anhydrous salts.
3. Write down four properties of an acid.

4. Give three characteristic properties of a base.
5. What do you understand by the term double decomposition?
6. Briefly state the difference between the following:
   (a) deliquescent and efflorescent salts
   (b) hygroscopic and deliquescent substances.
7. State three different methods of making normal salts. Illustrate each method by writing an equation for the reaction.
8. Describe in detail how you would prepare in the laboratory, crystals of zinc sulphate from zinc.
9. (a) Explain the meaning of the term 'basicity' of an acid.
   (b) Give an example of a dibasic acid and state how you can form an acid salt using the named acid.
10. The salts known as chlorides can be prepared by the following methods:
    (a) Direct combination
    (b) Precipitation
    (c) Neutralisation

    Select suitable, but different methods and briefly state how you can prepare each of the following salts.
    (i) Sodium chloride
    (ii) Silver chloride
    (iii) Lead(II) chloride
    (iv) Iron(III) chloride
11. A student wanted to prepare lead sulphate. He chose dilute sulphuric acid and lead(II) carbonate. He put the acid in a beaker and added lead(II) carbonate. But after a short while the reaction stopped. Explain why the reaction stopped?
12. Copy the following table and put a tick (√) if a precipitate is formed when the solutions are mixed and a cross (x) if no precipitate is formed.

Solutions	**Silver nitrate**	**Sodium sulphate**
Sodium chloride		
Barium chloride		
Lead nitrate		
Calcium chloride		
Magnesium sulphate		

13. Write balanced ionic equations for the following reactions:
    (a) Copper sulphate solution and zinc to product zinc sulphate solution and copper.

(b) Potassium bromide solution and chlorine gas to produce potassium chloride solution and liquid bromine.
(c) Sodium carbonate solution with copper(II) sulphate solution to produce copper carbonate precipitate and sodium sulphate solution.
(d) Sodium hydroxide solution with iron(II) chloride solution to produce iron(II) hydroxide precipitate and sodium chloride solution.
(e) Ammonium hydroxide with aluminium sulphate solution to produce ammonium sulphate solution and aluminium hydroxide precipitate.

UNIT 5: Effect of an electric current on substances

Introduction

The flow of electrons is known as an *electric current* or *electricity.* Electricity is generated in our country from water falls (hydroelectricity), geothermal power and from the sun i.e. solar energy. It is also stored in dry cells and lead acid accumulators (car batteries).

Electricity is transmitted along power cables by Kenya Power and Lighting Company (KPLC). This type of current is called *alternating current* (A.C). Car batteries and dry cells supply direct current (D.C.). It is the one commonly used in laboratory experiment.

In urban areas such as towns and cities, electricity is used on a large scale. People in these areas are greatly dependent on electricity supplies for household and industrial use.

5.1 Conduction of electricity

Conductors and non-conductors

Experiment 5.1

To investigate the conduction of electric current by solids.

Apparatus and chemicals

- torch cells or battery (a source of D.C)
- crocodile clips
- bulb (6V)
- connecting wires
- charcoal
- graphite
- sulphur
- wood
- rubber
- glass
- paper
- plastic
- metals – copper, Zinc, aluminium, tin, lead
- iodine crystal
- compounds – sodium chloride, sugar, lead (II) bromide

Procedure

1. Set up the apparatus as shown in Fig. 5.2 or 5.1.
2. Bring the crocodile clips close together without letting them come into contact. Observe the bulb.
3. Let the crocodile clips touch each other and observe the bulb. What happens?
4. Now clip a graphite (carbon) rod and observe the bulb.
5. Clip a piece of metal e.g. copper, iron etc and observe the bulb.
6. Copy Table 5.1 and record your observations.
7. Repeat the experiment using the other materials listed and each time record your observations.

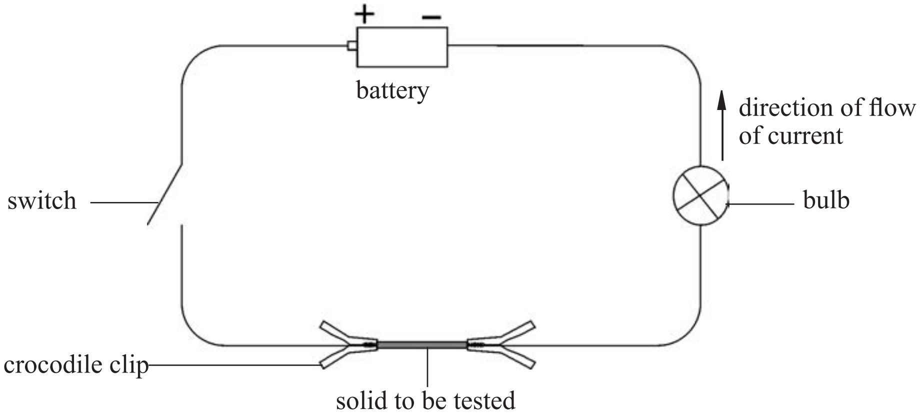


Fig 5.1. Testing solid material to see if it conducts electricity

A full battery or torch cells are usually drawn as indicated in the circuit diagram shown as Fig. 5.2. The long vertical strokes represent the +ve terminal while the short ones represent the –ve terminal.

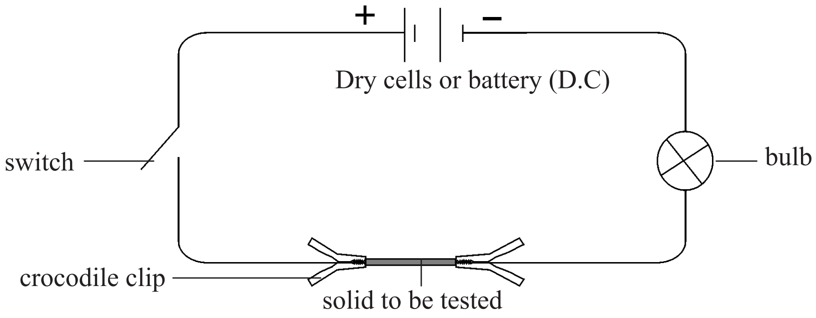


Fig. 5.2 Showing how a direct current source is represented in a circuit

Table 5.1 Conductors and non-conductors

Name of substance	Observation on the bulb
Graphite	
Copper	
Rubber	
Sulphur	
Aluminium	
Iron	
Paper	
Wood	
Magnesium	

- What is the purpose of the battery?
- What is the purpose of the bulb in the circuit?
- Why do some solids allow electricity to pass through while others do not?
- Are all the solids that conduct electric current the same or are there some exceptions?
- Do the solids that conduct electric current change at the end of the experiment?

It is observed that when the crocodile clips are apart the bulb does not light. The reason is that no electric current flows in the circuit as the circuit is incomplete. This is because air is a non-conductor and does not allow electricity to pass through it.

Any substance with which the bulb lit up allowed electricity to pass through it and it is an **electrical conductor.** Any substance that did not allow electricity to pass through is a **non-conductor**.

Some solids conduct electricity because they have **free mobile electrons**. These electrons are introduced into the circuit from the –ve and +ve terminal of the battery. All the electrons flow in one direction and when they reach the bulb it lights up.

Results of Experiment 5.1 show that all metals such as copper, zinc, aluminium, tin and lead are conductors. This is because they have free mobile electrons in their structures. The only non-metal that conducts electricity is **graphite.** This is because it has free electrons in its structure. Graphite is a form of carbon. It is used in the manufacture of dry cells. All other solid materials used such as rubber, paper, plastics and glass are non-conductors. Compounds like sodium chloride, sugar and lead(II) bromide among others are non-conductors.

The physical state of metals, whether solid or molten, does not change when current is passed through them. This also applies to graphite. Non-metals generally do not conduct electricity even in molten state.

5.2 Electrolysis

It is important first to understand the different types of liquids which we shall do investigation in the subsequent experiments.

(a) Test for conduction of electricity by aqueous solutions

Aqueous solutions

Water is the most common solvent. A solution of a substance in water is an **aqueous** solution. Experiment 5.2 is used to test whether substances dissolved in water conduct electricity.

Experiment 5.2

To investigate the conductivity of aqueous solutions.

Apparatus and chemicals

- battery/dry cells
- connecting wires
- bulb (6V) or ammeter
- beaker 250ml
- graphite rods
- distilled (pure) water
- acidified water
- aqueous solutions of: sugar, urea, copper(II) chloride, sodium chloride, mineral acids (hydrochloric, nitric and sulphuric acids).

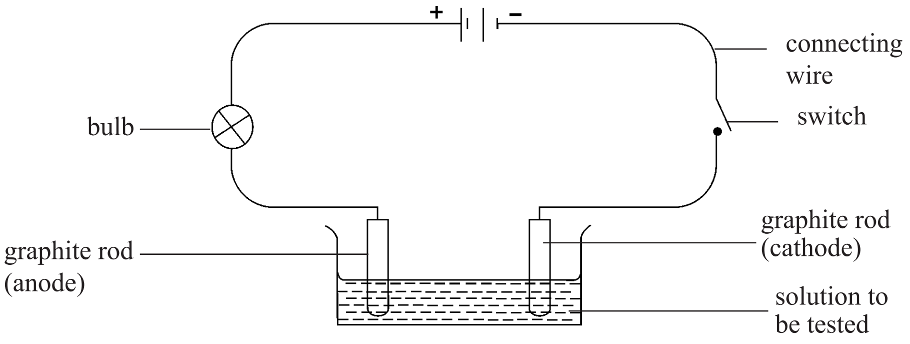


Fig. 5.3: The apparatus for testing the electrical conductivity of an aqueous solution

Procedure

1. Arrange the apparatus as in Fig. 5.3.
2. Switch on the current.
3. Before dipping the graphite rods into the solution, let the two electrodes come into contact and check whether the bulb lights up.
4. Place the graphite rods apart in the beaker.
5. Switch off the current.
6. Pour pure water into the beaker.
7. Switch on the current. Does the bulb glow?
8. Switch off the current. Add a few drops of concentrated sulphuric acid to the water in the beaker.
9. Switch on the current. Does the bulb now light?
10. What do you observe on the graphite rods? What does this indicate?
11. Repeat the above experiment using same graphite rods and the solutions of the following compounds : sugar, urea, copper(II) chloride, sodium chloride, mineral acids.
12. Note the changes (if any) occurring around the graphite rods and in the contents of the beaker. Rinse the beaker and the rods thoroughly with distilled water before each new solution is used.
13. Copy Table 5.2 and fill your results, include all the other substances being investigated.

Table 5.2 Electrolytes and non-electrolytes

Solution/substance	**Does the bulb light**		**Any observation made on the anode (+ve), cathode (–ve)**	
	Yes	**No**		
			Anode (+ve)	Cathode (-ve)
Copper(II) chloride	Yes		Bubbles form	Electrode is coated with a brown solid.
Sugar				
Sodium chloride				
Mineral acids				
Sodium chloride				
Copper (II) chloride				
Pure water				
Acidified water				

Different results are obtained from the apparatus in Fig. 5.3 when used to test different solutions. Liquids such as acidified water, sodium chloride solution, mineral acids and copper(II) chloride solution conduct electricity. Aqueous solutions which conduct electric current are called **electrolytes.**

The current decomposes the electrolyte as it passes through it. Aqueous solutions such as sugar and urea do not conduct electricity. These are known as **non-electrolytes**.

Electrolytic cell

Electrolysis is the decomposition of a compound by passing an electric current through it. The set-up of apparatus shown in Fig 5.4 is known as an **electrolytic cell**. The graphite rods which carry the current into and out of the electrolytes are called **electrodes.** Graphite is chosen because it does not react with the electrolyte or with the products formed during electrolysis. Unreactive metal electrodes such as platinum and copper can also be used as electrodes. The electrode connected to the positive terminal of the battery is called **anode**. The electrode connected to the negative terminal is called **cathode.**

(b) Test for conduction of electricity by melts (molten substances)

Melts

We have seen in the Experiment 5.2 that when some compounds are dissolved in water they conduct electricity whereas others do not. In our next experiment we shall investigate whether some compounds and elements conduct electricity when in molten form.

Experiment 5.3

To investigate the conductivity of molten substances.

Apparatus and chemicals

- crucible
- graphite rods
- connecting wire fitted with crocodile clips
- battery/dry cells with cell holder
- bunsen burner
- tripod stand
- pipe-clay triangle
- lead(II) bromide or lead(II) iodide
- sulphur
- paraffin wax

Procedure

Note: Lead(II) bromide or lead(II) iodide produce poisonous fumes. The experiment should be performed in a fume cupboard or in an open space, like a field.

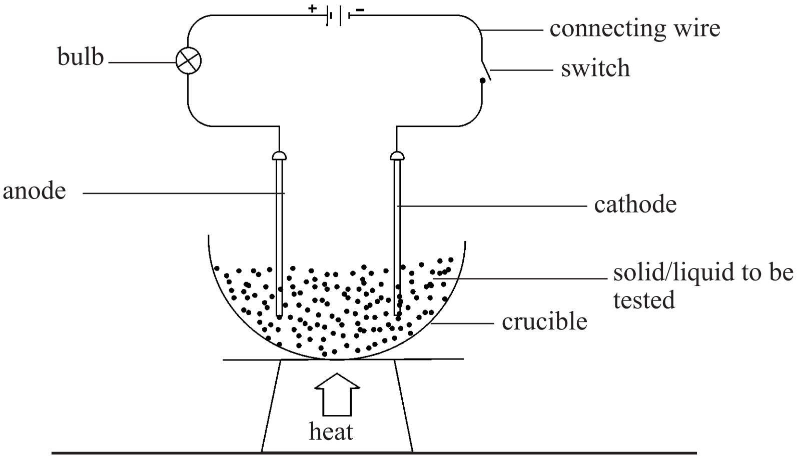


Fig. 5.4: Electrical conductivity of a molten substance

1. Set up the apparatus as in Fig 5.4.
2. Fill the crucible with either lead(II) bromide or lead(II) iodide to approximately two thirds.
3. Place the crucible and its contents on a pipe clay triangle supported by a tripod stand.
4. Insert the graphite electrodes into lead(II) bromide and close the switch. Does the bulb light?
5. Open the switch.
6. Gently heat the crucible until lead(II) bromide melts.
7. Adjust the burner so that the substance remains in molten state as the experiment progresses. Close the switch.
   Does the bulb light?
8. Observe whether there is any chemical reaction around the electrodes.
9. Record your observations as shown in Table 5.3.
10. Repeat the experiment with other substances such as paraffin, wax naphthalene, calcium chloride, sugar. Use a fresh crucible for each substance and clean the electrodes thoroughly before using it with the next substance.

Table 5.3 Electrolysis of molten substances

Substance	Does bulb light?		Observations at the electrodes	
	Solid	Molten	Anode	Cathode
Lead(II) bromide	No	Yes	Red brown fumes of bromine	Grey bead of lead metal
Sulphur				
Wax				
Lead (II) iodide etc.				

When we insert the electrodes in solid lead(II) bromide, the bulb does not light. But the bulb lights when lead(II) bromide is heated until it melts. This indicates that molten lead(II) bromide conducts electricity. It is an electrolyte. Metal compounds in molten state are electrolytes. It is also observed that other substances like sulphur, paraffin wax do not conduct electricity when in solid or molten form. These are non-electrolytes.

Other examples of electrolytes and non-electrolytes are given in Table 5.4.

Table 5.4 Other examples electrolytes and non-electrolytes

Electrolytes	Non-electrolytes
Sodium chloride solution	Pure water
Copper(II) chloride solution	Sugar solution
Sodium hydroxide solution	Paraffin wax
Molten lead(II) bromide	Molten sulphur
Molten lead(II) iodide	Solid lead(II) bromide
Hydrochloric acid	Solid sodium chloride
Sulphuric acid	Ethanol (alcohol)
Ethanoic acid	Urea
Most salts in molten or aqueous form	

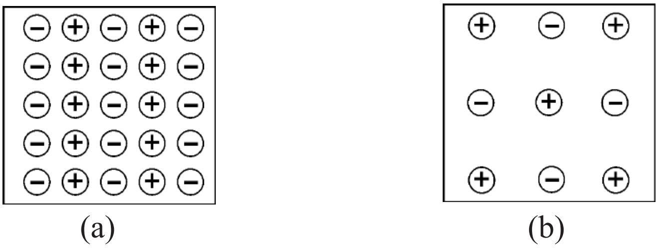
Why electrolytes conduct electricity

In Experiments 5.2 and 5.3 we tested the ability of a range of liquid compounds and solutions to conduct an electric current. We have seen that some compounds when dissolved in water or when in molten state are electrolytes. These include sodium chloride and lead(II) bromide and most other salts. Compounds such as sugar do not conduct electricity when in solution or molten form.

The reason for this behaviour is bonding in these compounds. When bonding is covalent, there are no free electrons or other charged particles that flow through the melt or solution.

Electrolytes are ionic compounds. They contain charged particles called *ions*. In solid form the ions are not free to move, therefore a solid ionic compound does not conduct electricity. The ions move if the solid is **melted** or **dissolved in water.** Study Figs. 5.5 and 5.6 (a & b). The ions carry the charge through the liquid and the current flows.

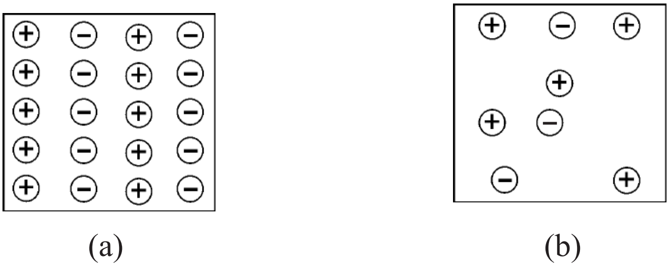
The positively charged ions are called **cations** and the negatively charged ions are **anions**. During electrolysis the cations move to the negative electrode (cathode) and the anions move to the positive electrode (anode).



Solid lead(II) bromide
Ions not fee to move

Molten lead(II) bromide
Ions are free to move

Fig. 5.5 State of lead(II) bromide in solid and molten form



Solid sodium chloride
Ions not free to move

Aqueous sodium chloride
Ions free to move

Fig. 5.6 State of sodium chloride in solid and solution form

Metals form cations which move to the –ve cathode when an electricity is passed though. Non-metallic elements form anions and these move to the +ve anode. Positive radicals and hydrogen ions move to the cathode. Negative radicals move to the anode.

Why non-electrolytes do not conduct electricity

Bonding in non-electrolytes is covalent. When dissolved, there are no charged particles (ions) to carry an electric current. We learnt in unit 3 that a covalent bond is formed when atoms share a pair of electrons. In this way they form discrete molecules. Therefore non-electrolytes exist in the form of molecules.

Examples of covalent compounds that do not conduct

Sugar	$C_{12}H_{22}O_{11}$
Urea	$(NH_2)_2CO$
Ethanol	C_2H_5OH
Distilled water	H_2O

5.2 Electrolysis of molten compounds

An electrolytic cell, like the one shown in Fig. 5.4 is used to pass an electric current through a molten ionic compound. After a compound has been melted, heat must be supplied to keep the salt in molten state. Fig. 5.7 and Fig. 5.8., show how ions occur in a binary compound and how they move after melting the compound.

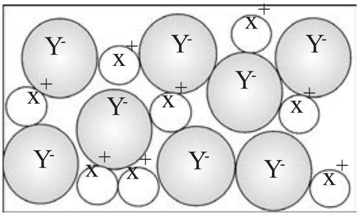
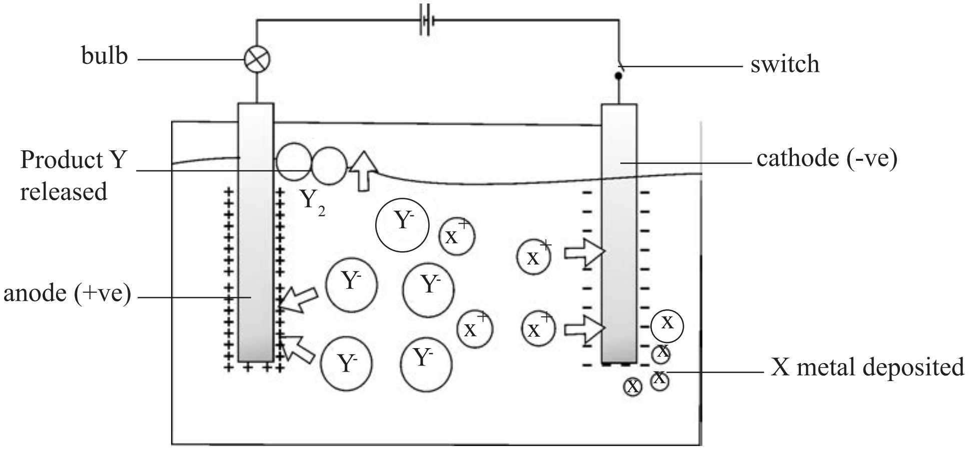


Fig. 5.7 Ions in a binary compound XY

When an ionic compound e.g. X^+Y^- is melted it produces X^+ ions and Y^-ions as follows:

$$X^+Y^-(s) \xrightarrow{\text{heat}} X^+(l) + Y^-(l)$$

When the switch is closed to complete the circuit, all +ve ions move to the cathode and all -ve ions move to the anode. As the current flow, a chemical reaction occurs at the electrodes. The ions lose the charge (they are discharged). Y^- ions lose electron(s) while X^+ ions gain electron(s). The respective products are deposited on the surface of each electrode, as shown in Fig. 5.8.



NON-METAL
- Negatively charged (-ve) ions are called anions and are attracted to the anode.

METALS
- Positively charged (+ve) ions are called cations and are attracted to the cathode.

Fig. 5.8 Discharging of anions at anode and cations at the cathode

Discharging ions

$X^+(l) + e^- \longrightarrow X(s)$

$Y^- (l) - e \longrightarrow Y(g)$

Let us now look at a particular salt and we compare it with the XY salt.

Electrolysis of molten Lead(II) bromide

Caution: Perform this experiment in a fume cupboard or in the open.

Lead(II) bromide is an ionic compound. When melted it, the electrolyte consists of two types of ions: - lead(II) ions and bromide ions.

$$\text{Lead(II) bromide} \xrightarrow{\text{Heat}} \text{lead(II) ions} + \text{bromide ions}$$

$$PbBr_2(s) \longrightarrow Pb^{2+}(l) + 2Br^-(l)$$

Lead(II) bromide is a binary salt. In molten state it produces two types of ions to be considered for discharging. It is easy to predict the products at the electrodes.

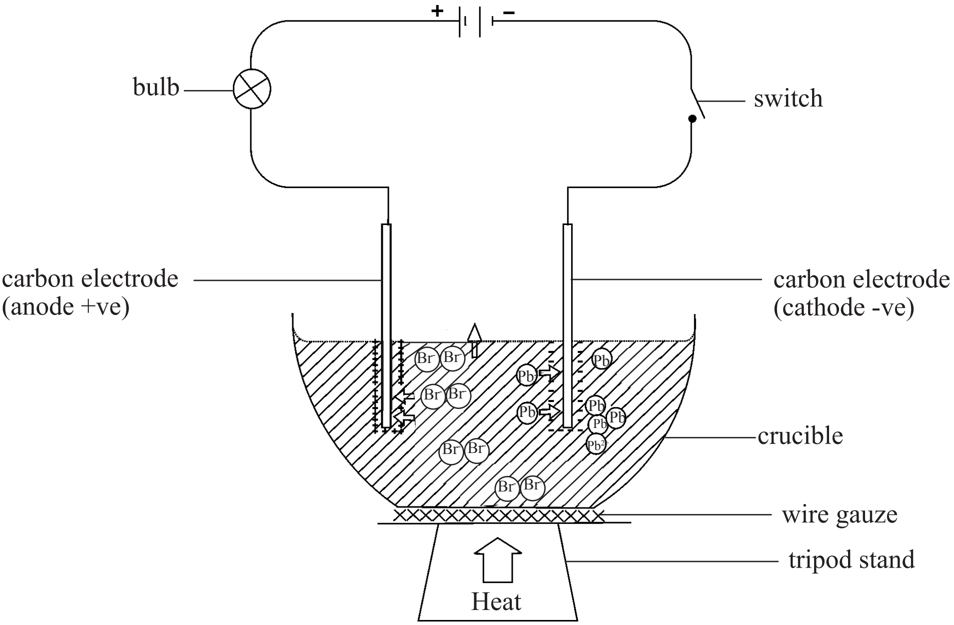


Fig. 5.9 The movement of ions during the electrolysis of molten lead(II) bromide.

Procedure

1. Set up the apparatus as in Fig. 5.9.
2. Fill a crucible to about two-thirds with lead(II) bromide.
3. Heat the salt until it melts. Keep on supplying heat to keep the salt in molten state.
4. Observe the bulb to see whether it lights.
5. Let the current pass for about 10 minutes.
6. Pay particular attention to any changes taking place around the electrodes. Record any observation made.
7. Remove the electrodes and examine them carefully. Record any observation.
8. After the crucible cools observe its contents. What do you observe around each electrode? Write equations to explain your observations (refer to the general ionic compound $X^+ Y^-$, Fig. 5.7).
9. A student passed an electric current through molten lead(II) iodide.
   Write equations to show products formed at the +ve and –ve electrodes.

The bulb lights brightly because molten lead(II) bromide conducts an electric current i.e. it is an electrolyte.

Brown vapours are observed at the anode. These is bromine fumes produced during electrolysis. Beads of a grey solid are deposited on the cathode and at the bottom of the crucible. This is lead metal.

We can conclude that as the electric current passes through molten lead(II) bromide, it decomposes it to form lead metal and bromine. The following equations show what happens at the electrodes.

At the anode

$Br^-(l)$ ions

- Move to the anode and are discharged by losing an electron and become a bromine atom.

  $Br^-(l) \rightarrow Br(g) + e$

- Bromine atoms combine in pairs to form molecules

  $Br(g) + Br(g) \rightarrow Br_2(l)$

It is common to write one equation once we understand that there are very many Br^- ions in the melt as follows:
$2Br^-(l) \rightarrow Br_2(l) + 2e$

At the cathode

$Pb^{2+}(l)$ ions

- Move to the cathode and discharge by gaining electrons to become lead atoms i.e.grey lead metal is formed.
  The metal is deposited on the cathode.

  $Pb^{2+}(l) + 2e \rightarrow Pb(s)$

The compound is thus decomposed into the elements lead and bromine by the electric current:

$$\text{Lead(II) bromide} \xrightarrow{\text{electric current}} \text{Lead} + \text{Bromine}$$

$$PbBr_2(l) \longrightarrow Pb(s) + Br_2(l)$$

Application of electrolysis

1. Electroplating

This is a method used in industries to coat (plate) a metal with another in order to improve its appearance by making it more attractive or in order to prevent corrosion.

Examples

- Gold-plated watches
- Silver–plated utensils
- Chromium plated electric kettles, car bumpers.
  Items to be plated with nickel are first plated with chromium because nickel does not adhere well to steel.

The items e.g. spoons to be electroplated (coated with another metal) are always used as the cathode, while the coating metal is the anode. The electrolyte is a salt of the metal to be used as the coat. For example, to electroplate spoons with silver, use silver nitrate solution.

2. Purification of metals

Impure copper can be purified by placing it at the anode. The cathode is a sheet of pure copper and the electrolyte copper(II) sulphate solution.

3. Extraction of metals

Copper metal is extracted from a rock by using the rock as the anode and pure copper the cathode. The electrolyte is copper(II) sulphate solution.

Reactive metals like sodium and aluminium are extracted from their compounds by electrolysis. Sodium is extracted from molten sodium chloride, while aluminium is extracted from its molten oxide.

Summary

1. A *conductor* (a metal or graphite) conducts an electric current but is not decomposed. Free or delocalised electrons are responsible for the conduction of an electric current.
2. *Electrolytes* are compounds which conduct an electric current in molten or aqueous state and are decomposed by the current into component elements in solid, or gaseous state.
3. Solutions of salts in water, solutions of acids and solutions of bases are electrolytes.
4. In an electrolyte ions carry the electric current as they move to the electrodes.
5. The positive electrode is called *anode*; the negative electrode is called *cathode*.
6. During electrolysis positive ions (metal ions or hydrogen ions) that move towards the cathode are called *cations*; negative ions that move towards the anode are called **anions**.
7. During electrolysis, anions are discharged by losing electrons at the anode. Therefore electrons leave the electrolyte via the anode. Electrons enter the electrolyte via the cathode where the cations gain these electrons.
8. *Graphite* is the only non-metal that conducts an electric current. It is mostly used because it does not react with the electrolyte or the products formed.
9. When electroplating (coating) an item (spoon, fork, knife) note the following:
   - The item to be coated should be the cathode e.g. spoon.
   - The metal to coat should be the anode e.g. copper
   - The electrolyte should be a salt of the metal to be used for coating.
10. An article is coated to prevent corrosion and make it attractive.

Revision Exercise 5

1. A student wants to electroplate a nickel jug with silver metal. Draw a labelled diagram showing the basic requirements to do this experiment.
2. Explain why solid sodium chloride does not conduct electricity but aqueous or molten sodium chloride does.
3. Explain the difference in electrical conductivity of sodium metal and solid sodium chloride.
4. What is meant by the following terms:
    (i) Electrolyte
    (ii) Conductor
    (iii) Non-electrolyte
    (iv) Non-conductor
    (v) Electrolysis
    (vi) Anode
    (vii) Cathode

UNIT 6: CARBON AND ITS COMPOUNDS

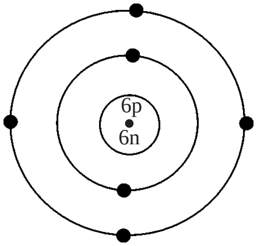
Introduction

All living things contain carbon compounds. These include carbohydrates, proteins, fats among others. Petrol, oil and wood are carbon compounds. Carbon is incorporated in plants from carbon(IV) oxide during photosynthesis. Carbon is one of the most useful and important elements. It is located in period two, group IV of the periodic table. Fig 6.1

Group	I	II		III	IV	V	VI	VII	O
Period 1	H 1								He 2
2	Li 3	Be 4		B 5	$^{6}_{12}C$ **Carbon**	N 7	O 8	F 9	Ne 10
3	Na 11	Mg 12		Al 13	Si 14	P 15	S 16	Cl 17	Ar 18
4	K 23	Ca 20							

Fig 6.1 Position of carbon in the periodic table

Carbon atom has atomic number 6 and electron arrangement of 2.4



C, 2.4

Fig 6.2 Atomic structure of carbon

There are four electrons in the outermost energy level in a carbon atoms. It is very difficult for a carbon atom to either lose or gain electrons. Therefore, carbon can only bond by sharing of electrons to gain stability.

6.1 Forms of carbon

Carbon occurs in the pure state in two allotropic forms; *diamond* and *graphite*. Allotropy is the existence of an element in more than one form but in the same state. It also occurs in many impure forms (charcoal). Charcoal is also made up of the graphite lattice.

Diamond

Structure

In diamond, each carbon atom is joined to four other carbon atoms. Each carbon atom shares electrons with each of its neighbours forming strong covalent bonds. Therefore all 4 valence electrons are involved in bonding, See Fig. 6.3

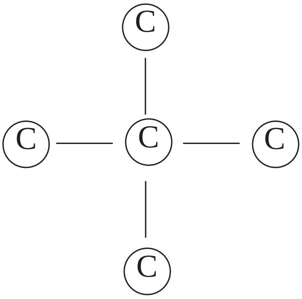


Fig. 6.3 Carbon atom bonded to four other carbon atoms

Each carbon atom at the centre is surrounded by four other carbon atoms at the corners to form a tetrahedron. The covalent bonds extend through all the molecules, forming a three-dimensional structure, as shown in Fig 6.4.

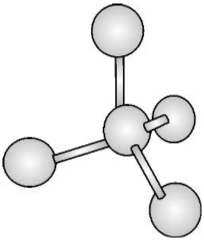


Fig. 6.4 Carbon forming three dimensional structure

In diamond, a giant atomic structure is formed by three dimensional structure as shown below. Each carbon atom is surrounded tetrahedrally by four others as illustrated in Fig. 6.5. Each carbon atom is bonded by very strong covalent bond to the other carbon atom.

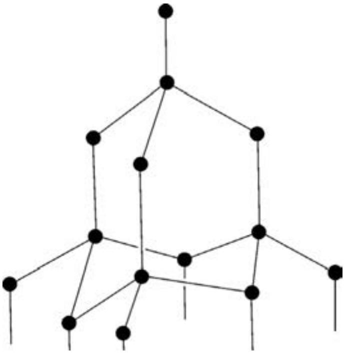


Fig. 6.5: Diamond giant atomic structure

Properties of diamond

- *Diamond is very hard*
  Carbon atoms in diamond are linked by very strong covalent bonds. The atoms in diamond are not arranged in layers, therefore they cannot slide over one another like the atoms in metals or graphite. This makes diamond the hardest substance ever known.

- *Diamond has a very high melting and boiling point*
  Since diamond is held by very strong covalent bonds, the atoms cannot vibrate enough to break away the bonds. This results in diamond having very high melting and boiling points.

- *Diamond does not conduct electricity*
  All valency electrons in diamond are involved in bonding. There are no delocalised electrons to conduct electricity like in metals or graphite.

- *Diamond has a very shiny sparkling appearance*

Uses of diamond

1. It is used for making jewellery e.g. necklaces, rings and earrings. Fig. 6.6
2. Diamond is used to cut glass and in drilling equipment as seen in Fig. 6.7 because it is very hard.

Fig 6.6 Couple exchanging rings

Fig. 6.7 Drilling through bedrock

Graphite

Structure

The structure of graphite is very different compared to that of diamond. In graphite, one carbon atom forms covalent bonds with three others to give a ring of six carbon atoms as illustrated in Fig. 6.8. Since there are four electrons available for bonding in a carbon atom, one electron is left free. When an electron is free to move within the structure, it is said to be **delocalised**.

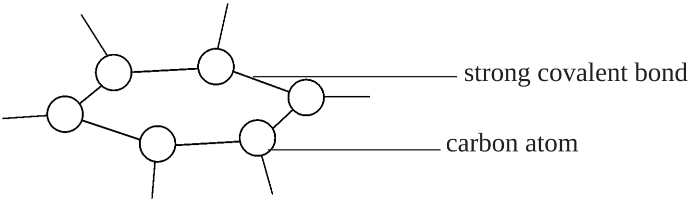


Fig. 6.8 A ring of carbon atoms

Many of these hexagonal rings join together by strong covalent bonds to form a layer so that every layer has a giant atomic structure. These giant atomic structures make flat sheets that lie on top of each other held together by weak van der Waals forces as illustrated in Fig. 6.10.

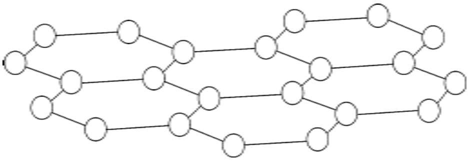


Fig 6.9 Giant atomic structure

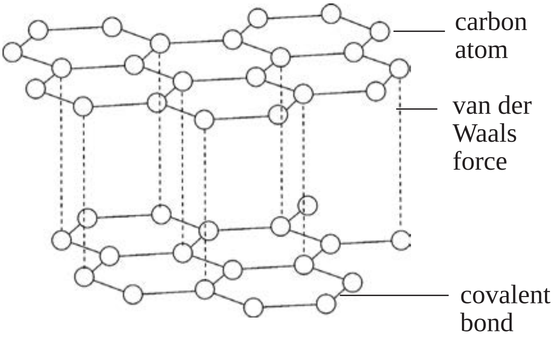


Fig 6.10 Structure of graphite showing the layers

Properties of graphite

- *Soft and slippery*

  Graphite is soft and slippery because the sheets of atoms can slide over each other easily due to the weak Van der waals forces between the parallel sheets of giant atomic structures.

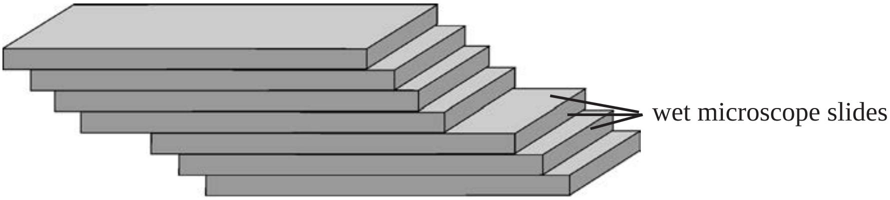


Fig. 6.11 The layers of graphite slide over each other like wet microscope slides.

- *Good conductor of electricity*

  This is because each carbon atom has four outer electrons but forms only three covalent bonds. The fourth electron is free to move through the graphite, carrying a charge. Therefore there are many delocalised electrons to conduct an electric current.

- *Graphite has a very high melting*

  Although the layers of graphite move over each other easily, it is difficult to break the bonds between carbon atoms within one layer. Covalent bonds are strong, hence require a lot of energy to break. Due to this, graphite melts at 3,370°C and boils at 4,830°C.

Uses of graphite

1. Graphite is used as lubricant because it is soft and slippery.
2. It is used to reinforce metals and broken bones. See Fig. 6.13.
3. Graphite is used as the positive terminals in dry cells and as electrodes in industries.

Fig. 6.12 Graphite lubricant

Fig. 6.13 A doctor at work

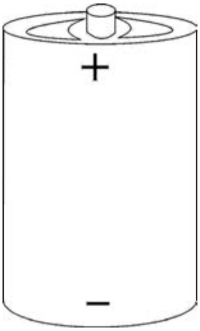


Fig 6.14 Dry cell of a torch

Charcoal

Structure

Charcoal is an amorphous (without shape) form of carbon and exists in many forms. These include:

Animal charcoal **-** formed when bones are heated in limited supply of air.

Wood charcoal **-** which is the solid left when wood is heated in limited supply of air.

Sugar charcoal **-** is formed by dehydrating cane-sugar or glucose with concentrated sulphuric acid or heating the sugar in the absence of air.

Lamp black **-** is formed when petroleum, kerosene, turpentine, natural gas and other hydrocarbons burn in a limited supply of air.

Coke **-** is the solid left when coal is heated in absence of air.

Soot - is formed when there is incomplete combustion of fuels. It is found in chimneys of houses, lantern lamps, and many other places.

Uses

1. Because of its good absorption power, amorphous carbon is used to purify sugar in industries.
2. It can also be used in gas masks for people who work in industries which produce poisonous gases since it absorb gases, see Fig. 6.15.

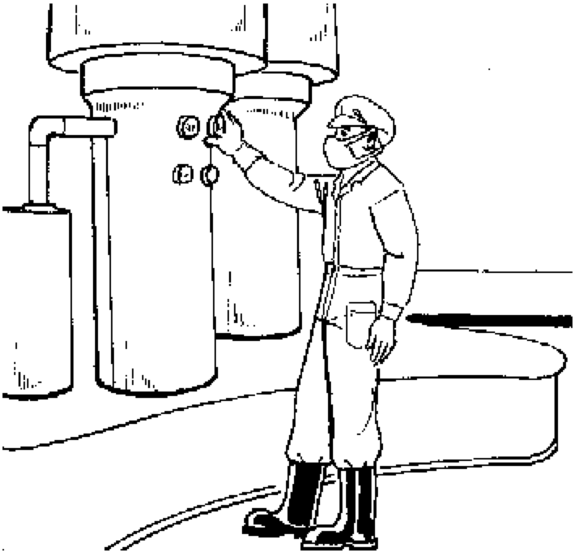


Fig. 6.15 Gas masks containing powdered charcoal

Table 6.1 A summary of properties of allotropes of carbon

	Diamond	**Graphite**	**Amorphous carbon**
Appearance	Colourless, transparent crystals that sparkle in light	Dark-grey, opaque and shining	Black, opaque and very dull
Density	Highest (3.5g/cm^3)	Moderate (2.3g/cm^3)	Low 1.5g/cm^3 when air free
Hardness	Hardest natural substance known	Soft and slippery	Soft
Electrical conductivity	Does not conducts electricity	Conducts electricity	Does not conduct electricity

6.2 Chemical properties of carbon

(a) combustion

Experiment 6.1: Burning of carbon

Apparatus and chemicals

- charcoal
- deflagrating spoon
- source of heat

Procedure

1. Burn charcoal in a deflagrating spoon.
2. Drop the burning charcoal in a gas jar which has calcium hydroxide solution. The charcoal should not touch the solution. What happens to the solution?
   A gas is produced that forms a white precipitate with calcium hydroxide solution.
   All forms of carbon burn in sufficient supply of oxygen to form carbon(IV) oxide and heat.

$$\text{Carbon} + \text{oxygen} \rightarrow \text{carbon(IV) oxide} + \text{heat}$$
$$C(s) + O_2(g) \rightarrow CO_2(g) + \text{heat}$$

Diamond and graphite ignite to red-hot with a lot of difficulty.
If burned in an insufficient supply of air (oxygen), carbon(II) oxide is produced instead.

$$\text{Carbon} + \text{oxygen} \rightarrow \text{carbon(II) oxide} + \text{heat}$$
$$2C(s) + O_2(g) \rightarrow 2CO(g) + \text{heat}$$

(b) Reducing nature of carbon

Experiment 6.2 To investigate the reducing action of carbon on some oxides

Apparatus and chemicals

- crucibles
- blow pipe
- lead(II) oxide
- zinc oxide
- iron (III) oxide
- copper(II) oxide
- source of heat

Procedure

1. Mix charcoal powder with lead(II) oxide in a crucible. See Fig. 6.1 a. What is the initial colour of the mixture?
2. Apply heat on the crucible until it is red-hot. See Fig. 6.16 b. What do you observe? What is the colour of the final product? Write the chemical equations of the reaction.
3. Record your observations in Table 6.2.
4. Repeat the same procedure with zinc oxide, iron(III) oxide and copper(II) oxide. Record your observations as illustrated in Table 6.2.

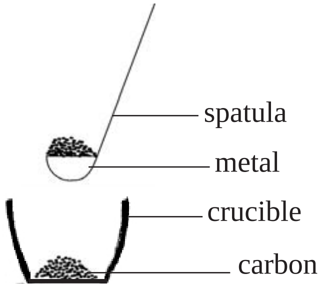


Fig. 6.16 (a) mixing lead (II) oxide with charcoal

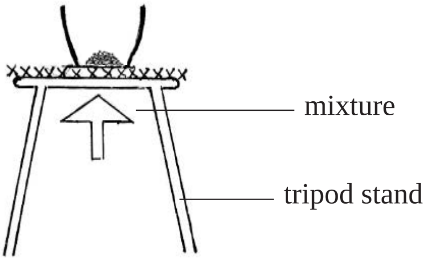


Fig. 6.16 (b) Heating of the mixture of lead (II) oxide and charcoal

Table 6.2 Reducing action of carbon on some metal oxide

Metal oxide	**Colour of mixture before heating**	**Colour of mixture after heating**	**Record any other observation and write chemical equations**
Zinc oxide and carbon			
Lead(II) oxide and carbon			
Iron(III) oxide and carbon			
Copper(II) oxide and carbon			

It is noted that in all cases a red glow started to spread through carbon/metal oxide mixture and continued to spread even in the absence of heat. Silvery beads of lead are formed. For black copper(II) oxide, a red-brown copper was formed. In all cases, carbon reduces these metal oxides to their respective metals.

Lead(II) oxide + carbon → lead + carbon(IV) oxide
$2PbO(s) + C(s) \rightarrow 2Pb(s) + CO_2(g)$
(yellow) (silvery)

Zinc oxide + carbon → zinc + carbon(IV) oxide
$2ZnO(s) + C(s) \rightarrow 2Zn(s) + CO_2(g)$
(white) (grey)

Iron(III) oxide + carbon → iron + carbon(IV) oxide
$2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$
(red-brown) (grey)

Copper(II) oxide + carbon → copper + carbon(IV) oxide
$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$
(black) (red-brown)

(c) Reaction of carbon with acids and steam

Carbon reduces concentrated sulphuric and nitric acids when heated with them.

Carbon + sulphuric acid → carbon(IV) oxide + sulphur(IV) oxide + water
$C(s) + 2H_2SO_4(aq) \rightarrow CO_2(g) + 2SO_2(g) + 2H_2O(l)$

Carbon + nitric acid → carbon(IV) oxide + nitrogen(IV) oxide + water
$C(s) + 4HNO_3(aq) \rightarrow CO_2(g) + 4NO_2(g) + 2H_2O(l)$

These acids donate oxygen to carbon to form carbon(IV) oxide

Carbon + oxygen → carbon(IV) oxide
$C(s) + O_2(g) \rightarrow CO_2(g)$

Water-gas which is a mixture of carbon(II) oxide and hydrogen is produced by passing steam through coke at temperatures higher than 1000˚C.

$$\begin{array}{lcccc} \text{Carbon} & + & \text{steam} & \rightarrow & \text{carbon(II) oxide} + \text{hydrogen} \\ C(s) & + & H_2O(g) & \rightarrow & \underbrace{CO(g) \quad + \quad H_2(g)}_{\text{(water-gas)}} \\ \text{(coke)} & & \text{(steam)} & & \end{array}$$

Carbon(II) oxide are hydrogen and combustible and therefore water-gas is used as a fuel. Equations for the reactions of combustion of water-gas are shown below.

$$\begin{array}{ccccccc} \text{Hydrogen} & + & \text{oxygen} & \rightarrow & \text{steam} & + & \text{heat} \\ 2H_2(g) & + & O_2(g) & \rightarrow & 2H_2O(g) & + & \text{heat} \end{array}$$

$$\begin{array}{ccccccc} \text{Carbon(II) oxide} & + & \text{oxygen} & \rightarrow & \text{carbon(IV) oxide} & + & \text{Heat} \\ 2CO(g) & + & O_2(g) & \rightarrow & 2CO_2(g) & + & \text{Heat} \end{array}$$

6.3 Carbon(IV) oxide

Ordinary air contains about 0.03 per cent of carbon(IV) oxide. Exhaled air contains about 3 per cent. The burning of wood, coal, petrol,oils and other carbon compounds add carbon(IV) oxide to the atmosphere. Green plants use it to make their food. We also use in soft drinks like cocacola.

Experiment 6.3

The laboratory preparation of carbon(IV) oxide (CO_2).

Apparatus and chemicals

- flat-bottomed flask
- thistle funnel
- two wash bottles.
- delivery tubes
- gas jars
- cardboard cover
- calcium carbonate (marble or limestone)
- dilute hydrochloric acid
- distilled water
- concentrated sulphuric acid

Procedure

1. Arrange the apparatus as shown in Fig 6.17, but without the wash bottles.
2. To the limestone or marble chips, add dilute hydrochloric acid drop by drop. What do you observe?
3. Collect the gas by downward delivery (upward displacement of air) or over warm water. Why do we choose upward displacement of air to collect the gas?

- Why is the gas collected over warm water?
- Why is the gas passed through water then through concentrated sulphuric acid?

4. Record your observation in your notebook.
5. Collect a few jars of the gas for the next experiment. The gas is passed through a wash bottle containing water and then through another one containing concentrated sulphuric illustrated in Fig 6.17. The first bottle removes suspended hydrochloric acid sprays of hydrogen chloride gas and the second dries the gas (i.e. remove water vapour).

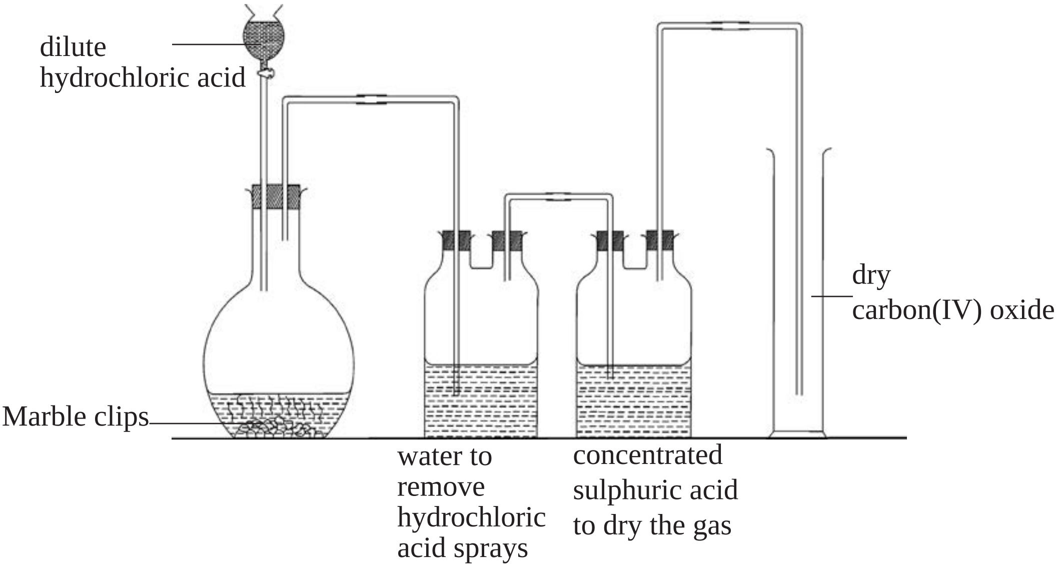


Fig. 6.17 Preparation of carbon(IV) oxide

The gas should now be collected by downward delivery
A steady reaction is observed when marble chips react with dilute hydrochloric acid.

Calcium carbonate + hydrochloric acid → calcium chloride + carbon(IV) oxide + water

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Any mineral acid liberates carbon(IV) oxide with a carbonate salt. The reaction is, however, unsuitable if the salt formed is insoluble or sparingly soluble in water. For example, calcium carbonate and dilute sulphuric acid liberate the gas for a short time; then the reaction stops. Why?

The sparingly soluble calcium sulphate formed by the reaction of calcium carbonate and dilute sulphuric acid forms a coating on the unreacted calcium carbonate and the sulphuric acid will no longer have contact with the calcium carbonate. Solubility of gases decrease with increase in temperature hence the reason for collecting the gas over warm water.

Is the reaction between dilute sulphuric acid and lead carbonate or barium carbonate suitable for the preparation of carbon(IV) oxide? Explain your answer.

Experiment 6.4

To investigate the properties of carbon(IV) oxide

Apparatus and chemical required

- gas-jars full of carbon(IV) oxide
- test-tubes
- magnesium ribbon
- pair of tongs
- candle
- blue litmus paper
- two water troughs
- beehive shelfs
- calcium hydroxide
- sodium hydroxide solution
- dilute nitric acid

Procedure

1. Examine a gas jar full of carbon(IV) oxide.
    - What is the colour of the gas?
    - How does it smell?
2. Pass carbon(IV) oxide from the generator through a test-tube containing calcium hydroxide solution for a short time as shown in Fig. 6.18
3. What do you observe? Continue passing more of the gas into this same test-tube until there is no further change.
    - What do you observe?
    - Write a balanced chemical equation for the reaction.

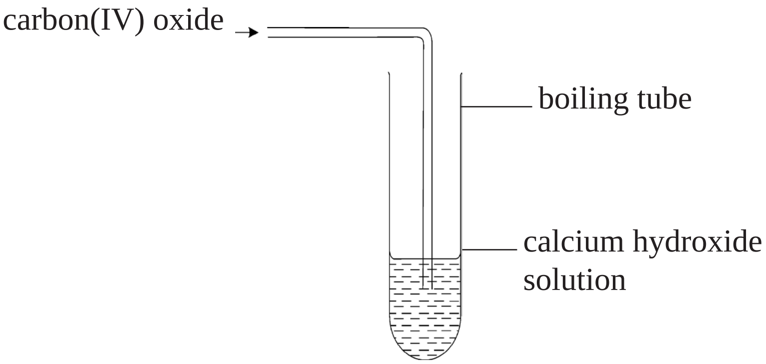


Fig 6.18 Reaction of carbon(IV) oxide with calcium hydroxide solution

4. Pass carbon(IV) oxide into a test-tube containing distilled water and then dip a blue litmus paper into the resulting solution.
    - What colour change occurred with the blue litmus paper? Explain.
    - Write a balanced chemical equation for the reaction between carbon(IV) oxide and water.
5. Pour a gas-jar full of carbon(IV) oxide on a burning candle. What do you observe?

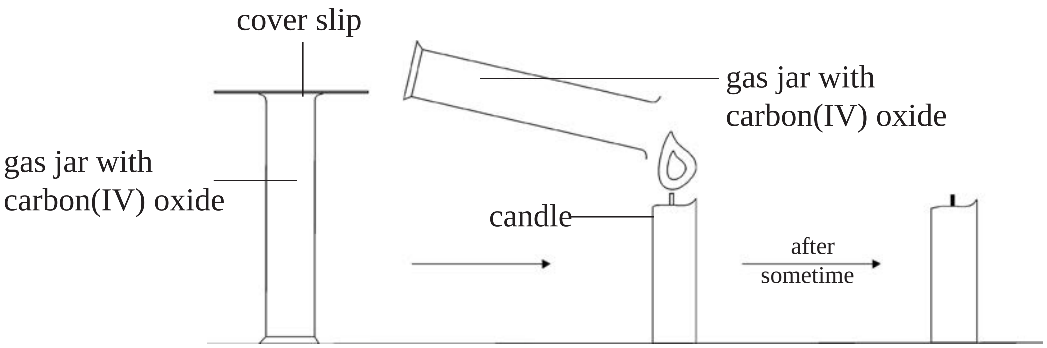


Fig 6.19 Pouring carbon(IV) oxide into a burning candle

6. Plunge a piece of burning magnesium ribbon quickly in a gas-jar full of carbon(IV) oxide as illustrated in Fig. 6.20.

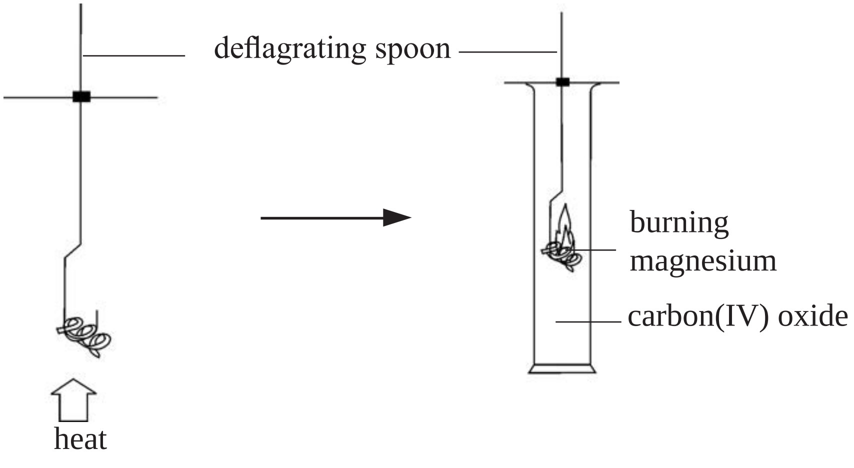


Fig 6.20 Burning magnesium in carbon(IV) oxide

What do you observe on the sides of the gas-jar?
Add about 50cm^3 of dilute nitric acid to the mixture in the gas-jar

- What do you observe?
- Write a chemical equation for the reaction.

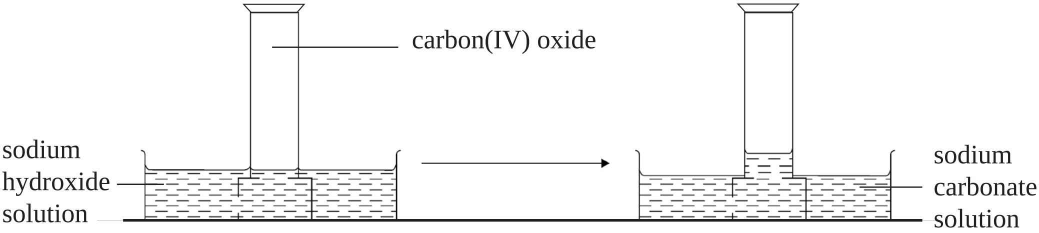


Fig 6.21 Reaction of carbon(IV) oxide with sodium hydroxide solution

7. Using a set-up as in Fig 6.21 invert a gas jar full of carbon(IV) oxide in a trough containing sodium hydroxide solution instead of water. Shake the gas jar. Wait for sometime and note any water level changes inside the gas jar.
   - What happens to the level of sodium hydroxide solution? Explain.
   - Write chemical equations for the reaction.

Physical properties of carbon(IV) oxide

- Carbon(IV) oxide is a colourless gas.
- It has a slight smell.
- It is denser than air. (This is why it can be collected by downward delivery).
- It is slightly soluble in water.

Chemical properties of carbon(IV) oxide

- **Reaction of carbon(IV) oxide with calcium hydroxide**

  It forms a white precipitate with calcium hydroxide solution. This precipitate is called calcium carbonate and it is insoluble in water.

  Calcium hydroxide + carbon(IV) oxide → calcium carbonate + water

  $$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

When carbon(IV) oxide is passed through the calcium carbonate and water for a long time, the insoluble calcium carbonate is converted into soluble calcium hydrogencarbonate solution.

Calcium carbonate + carbon(IV) oxide + water → calcium hydrogencarbonate

$$CaCO_3(s) + CO_2(g) + H_2O(l) \rightarrow Ca(HCO_3)_2(aq)$$

- **Reaction of carbon(IV) oxide with water**

  Carbon(IV) oxide moderately reacts with water. The solution formed is slightly acidic, it does not turn blue litmus paper fully red. Some carbonic acid is formed but in very small quantities.

  Carbon(IV) oxide + water → carbonic acid

  $$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

Since carbon(IV) oxide reacts with water to form carbonic acid, this acid formed reacts with sodium hydroxide to form sodium carbonate and water only. The sodium hydroxide level rises inside the gas jar to occupy the space of the reacted carbon(IV) oxide as illustrated by Fig. 6.21.

Sodium hydroxide + carbon(IV) oxide → sodium carbonate + water

$$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

Since carbonic acid is dibasic it forms a normal and an acid salt. When carbon(IV) oxide is passed for along time in concentrated solution of sodium hydroxide a further reaction occurs forming sodium hydrogencarbonate, some of which precipitates.

Sodium carbonate + carbon(IV) oxide + water → sodium hydrogencarbonate

$$Na_2CO_3(aq) + CO_2(g) + H_2O(l) \rightarrow 2NaHCO_3(aq)$$

Potassium hydroxide solution gives a similar reaction but its hydrogen carbonate is more soluble and does not precipitate. This is why potassium hydroxide solution is preferred as an absorbent for carbon(IV) oxide.

- **Carbon(IV) oxide and combustion**

  Carbon(IV) oxide extinguishes a burning candle showing that it does not support combustion.

- **Reaction of carbon(IV) oxide with magnesium**

  The burning magnesium ribbon continues to burn inside a gas jar full of carbon(IV) oxide. Black carbon particles form on the sides of the jar, together with white ash of magnesium oxide. The addition of dilute nitric acid enables the black carbon particles to be seen more clearly because the acid reacts with magnesium oxide forming magnesium nitrate solution and water. Carbon(IV) oxide does not support combustion but magnesium continues to burn in it. Why is it so? The heat of burning magnesium is sufficient to decompose carbon(IV) oxide into carbon and oxygen.

$$\text{Carbon(IV) oxide} \xrightarrow{\text{heat}} \text{carbon} + \text{oxygen}$$
$$CO_2(g) \rightarrow C(s) + O_2(g)$$

The oxygen produced supports the combustion of magnesium to form magnesium oxide.

$$\text{Magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$$
$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

The general equation can thus be written as follows:

$$\text{Magnesium} + \text{Carbon(IV) oxide} \rightarrow \text{Magnesium oxide} + \text{Carbon}$$
$$Mg(s) + CO_2(g) \rightarrow MgO(s) + C(s)$$

This experiment shows that magnesium displaces oxygen from carbon(IV) oxide. Therefore magnesium is more reactive towards oxygen than carbon. It also shows that carbon(IV) oxide contains carbon and oxygen.

Uses of carbon(IV) oxide

1. Carbon(IV) oxide is used in the manufacture of sodium carbonate, used for baking cakes, bread among other products and sodium hydrogencarbonate which is used in some health salts such as *Eno* and *Andrews liver salt*.

2. Carbonated drinks
In the production of mineral water and carbonated drinks like coca-cola (also called aerated or effervescence drinks). Soda-water is a solution of carbon(IV) oxide in water under pressure. It is later sweetened, flavoured and sometimes coloured. The dissolved carbon(IV) gives it a pleasant taste.

3. As a refrigerant
Solid carbon(IV) oxide commonly known as *dry ice* is preferred to ordinary ice (solid water) because it sublimes at room temperature forming gaseous carbon(IV) oxide and therefore leaves no residue like ordinary ice. Fig. 6.22.

4. Fire extinguishers
Fire extinguishers often contain sodium hydrogencarbonate or sodium carbonate solution and sulphuric acid. When mixed by inversion or pressing a plunge, carbon(IV) oxide is produced due to the reaction of carbonate with sulphuric acid. Fig. 6.23 and 6.24.

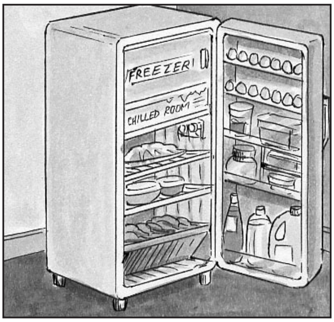


Fig 6.22. A refrigerator for cooling perishables

Sodium hydrogencarbonate + sulphuric acid → sodium sulphate + carbon(IV) oxide + water

$$2NaHCO_3(aq) + H_2SO_4 \rightarrow Na_2SO_4(aq) + 2CO_2(g) + 2H_2O(l)$$

Sodium carbonate + sulphuric acid → sodium sulphate + carbon(IV) oxide + water

$$NaCO_3(aq) + H_2SO_4 \rightarrow Na_2SO_4(aq) + CO_2(g) + H_2O(l)$$

A high pressure is set which forces a froth of carbon(IV) oxide through the jet. It is directed onto the burning substances.

Fig. 6.23 Fire extinguisher

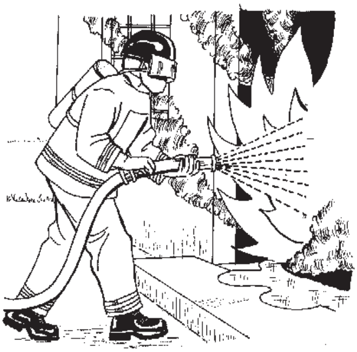


Fig 6.24 A fire fighter

In fire extinguishers that have carbon(IV) oxide stored under pressure, the carbon(IV) oxide is released through a nozzle.

Carbon(IV) oxide does not support combustion and since it is denser than air, it forms a blanket on the burning substance thereby cutting off the supply of oxygen from the air.

Check the types of fire extinguisher you have in your school. The teacher can demonstrate how to operate them.

5. **Making rain during drought or in areas of little rain**
   In making rain, dry ice (solid carbon(IV) oxide) is spread in the clouds to accelerate the condensation process. Small aircrafts are sometimes used to spread dry ice in the sky.

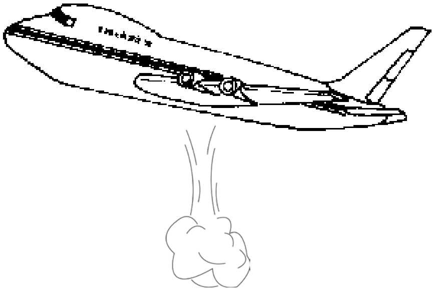


Fig. 6.25 Aircrafts dropping dry ice

6. In Solvay process
   This is a process used in the manufacture of sodium carbonate (soda-ash).
7. It is also used to transfer heat energy from certain types of nuclear reactors.

6.4 Carbon(II) oxide

Note: Preparation and investigation of properties of carbon(II) oxide should not be done practically due to its poisonous nature. It will only be discussed theoretically.

Experiment 6.5

Laboratory preparation of carbon(II) oxide (CO)

Apparatus and chemicals required

- round-bottomed flask
- thistle funnel/dropping funnel
- wash-bottle
- water trough
- gas-jar
- beehive shelf
- sodium methanoate or methanoic acid.
- concentrated sulphuric acid
- potassium hydroxide solution
- distilled water

Procedure

1. Arrange the apparatus as shown in Fig. 6.26.
2. Two or three spatulas of sodium methanoate is put in the round-bottomed flask.
3. Concentrated sulphuric acid is poured drop by drop through the thistle funnel or dropper.
4. If the reaction is slow, the reagents are warmed gently.

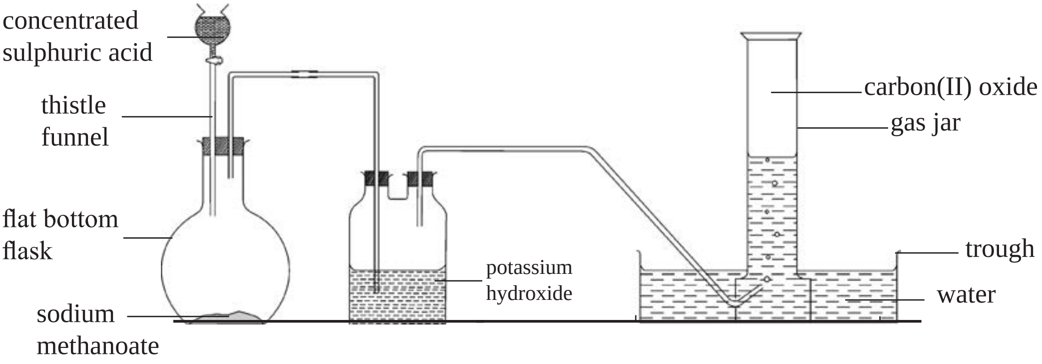


Fig. 6.26 Preparation of carbon(II) oxide from concentrated sulphuric acid and sodium methanoate

Sulphuric acid first liberates methanoic acid from sodium methanoate and then dehydrates it.

Sodium methanoate + sulphuric acid → methanoic acid + sodium hydrogensulphate

$$HCOONa(s) + H_2SO_4(aq) \rightarrow HCOOH(aq) + NaHSO_4(aq)$$

$$\text{Methanoic acid} \xrightarrow{\text{Concentrated Sulphuric acid}} \text{Carbon(IV) oxide} + \text{water}$$

$$HCOOH(aq) \xrightarrow[H_2SO_4(l)]{} CO_2(g) + H_2O(l)$$

Potassium hydroxide solution removes any carbon(IV) oxide or sulphur(IV) oxide which may be present.

The gas is usually collected over water as shown in Fig. 6.26. If it is required dry, it may be passed through concentrated sulphuric acid and then collected by upward delivery (being less dense than air).

Note: Methanoic acid can be used instead of sodium methanoate.

Experiment 6.6

Preparation of Carbon(II) oxide from carbon(IV) oxide.

Apparatus and chemicals required

- combustion tube
- source of heat
- source of dry carbon(IV) oxide
- wash-bottle
- gas jar
- water trough
- beehive shelf

- potassium hydroxide solution
- charcoal
- distilled water

Procedure

1. Charcoal is put in a combustion tube and the apparatus arranged as shown in Fig. 6.27.
2. The combustion tube should be full of charcoal (closely packed). This will increase efficiency and make sure that almost all the carbon(IV) oxide reacts with charcoal.
3. The charcoal is heated to red-hot.
4. Dry carbon(IV) oxide is passed over the red-hot charcoal.

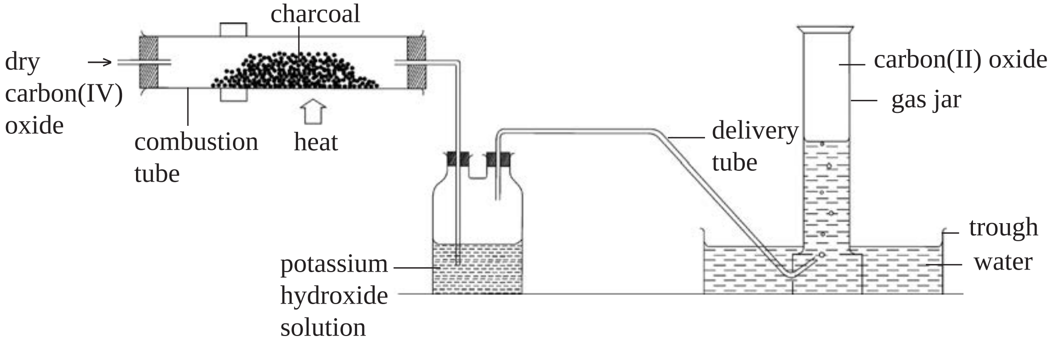


Fig 6.27 Preparation of carbon(II) oxide from carbon(IV) oxide

Carbon(IV) oxide is reduced by the red-hot charcoal to carbon(II) oxide.

Carbon(IV) oxide + carbon → carbon(II) oxide

$$CO_2(g) + C(s) \rightarrow 2CO(g)$$

Any unreacted carbon(IV) oxide is removed by potassium hydroxide solution.

Potassium hydroxide + carbon(IV) oxide → potassium carbonate + water

$$2KOH(aq) + CO_2(g) \rightarrow K_2CO_3(aq) + H_2O(l)$$

Physical properties of carbon(II) oxide (CO)

- Colourless gas
- Odourless
- Tasteless
- Less dense than air
- Insoluble in water

Experiment 6.7

Combustion of carbon(II) oxide

Apparatus and chemicals

- source of carbon(II) oxide
- calcium hydroxide solution
- test-tube

Procedure

1. Ignite carbon(II) oxide. Fig 6.28.
   Record your observation.
2. If the flame was directed into a test-tube containing calcium hydroxide (Fig. 6.29) What would happen to the calcium hydroxide solution?

Fig. 6.28: Burning of carbon(II) oxide

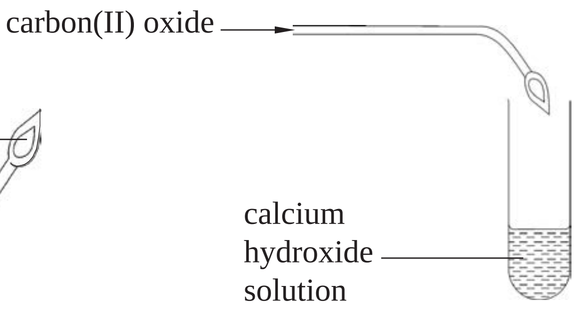


Fig. 6.29: Reaction of gas produced by burning carbon (II) oxide with calcium hydroxide solution

Carbon(II) oxide gas burns with a blue flame forming carbon(IV) oxide gas.

Carbon(II) oxide + oxygen → carbon((IV) oxide

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

Carbon(IV) oxide forms a white precipitate of calcium carbonate with calcium hydroxide solution.

Carbon(II) oxide is said to be a **reducing agent** because it takes oxygen. Oxygen is the **oxidizing agent**.

Experiment 6.8

To investigate the chemical properties of carbon(II) oxide.

Apparatus and chemicals

- source of dry carbon(II) oxide
- combustion tube
- porcelain boats
- test tubes
- zinc oxide
- lead(II) oxide
- iron(III) oxide
- copper(II) oxide
- calcium hydroxide solution

Procedure

1. Put lead(II) oxide and copper(II) oxide in a combustion tube.
2. The rest of the apparatus are arranged as shown in the same Fig. 6.30. What colour are the oxides?
3. Heat the oxides strongly. What is happening to calcium hydroxide solution? What is the colour of the cold residues?
4. Repeat the same with iron(III) oxide and zinc oxide.

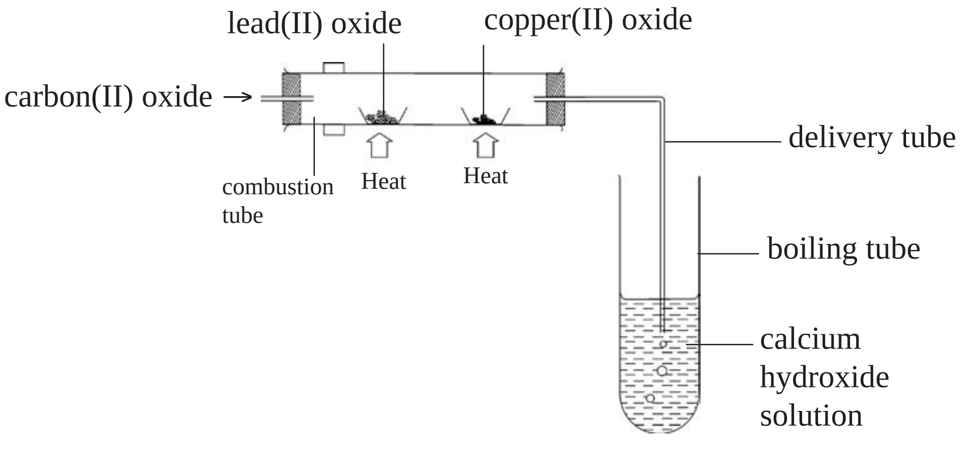


Fig. 6.30 Reducing properties of carbon(II) oxide

Carbon(II) oxide reduces lead(II) oxide to grey metallic lead and copper(II) oxide to red-brown copper. The oxygen removed from these oxides converts carbon(II) oxide to carbon(IV) oxide. Carbon(IV) oxide reacts with calcium hydroxide to form a white precipitate of calcium carbonate.

Lead(II) oxide + carbon(II) oxide → lead + carbon(IV) oxide

$PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$

(yellow) (bluish grey)

Copper(II) oxide + carbon(II) oxide → copper + carbon(IV) oxide

$CuO(s) + CO(g) \rightarrow Cu(s) + CO_2(g)$

(black) (red-brown)

Zinc(II) oxide + carbon(II) oxide → Zinc + carbon(IV) oxide

$ZnO(s) + CO(g) \rightarrow Zn(s) + CO_2(g)$

(white) (grey)

Iron(III) oxide + carbon(II) oxide → Iron + carbon(IV) oxide

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

(red-brown) (grey)

Calcium hydroxide + carbon(IV) oxide → calcium carbonate + Water

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

(colourless) (white precipitate)

Therefore carbon(II) oxide is a **reducing agent.** It reduces zinc oxide to zinc, lead(II) oxide to lead, copper(II) oxide to copper and iron(III) oxide to iron.

It does not reduce magnesium oxide or aluminium oxide because magnesium and aluminium are higher in the activity series.

Combustion of carbon(II) oxide

Carbon(II) oxide does not support combustion. In internal combustion, engines which use petrol or diesel, insufficient supply of oxygen inside the engines produces carbon(II) oxide making engine exhaust gases very poisonous especially in places where ventilation is poor.

A burning jiko usually produce carbon(II) oxide in a badly ventilated room (a room with limited supply of air).

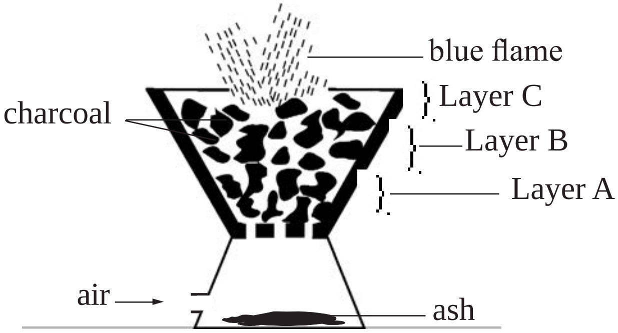


Fig. 6.31: Burning charcoal in a jiko

In layer A there is sufficient supply of air oxygen and therefore carbon burns to carbon(IV) oxide.

Carbon + oxygen → carbon(IV) oxide

$C(s) + O_2(g) \rightarrow CO_2(g)$

The carbon(IV) oxide from layer A is reduced by the red-hot carbon in layer B to form carbon(II) oxide.

Carbon + carbon(IV) oxide → carbon(II) oxide

$C(s) + CO_2(g) \rightarrow 2CO(g)$

Around area C the carbon(II) oxide is oxidised by the oxygen from the air producing a characteristic blue flame.

Carbon(II) oxide + oxygen → carbon(IV) oxide

$2CO(s) + O_2(g) \rightarrow 2CO_2(g)$

If the air (oxygen) supply at C is limited, some of the carbon(II) oxide will pass unoxidised. This makes the room to be filled with this very poisonous gas.

Carbon(II) oxide is poisonous because it combines with haemoglobin in the blood to form a stable compound called *carboxyhaemoglobin*. Carboxyhaemoglobin does not give up oxygen to body cells like oxyhaemoglobin (a combination of haemoglobin and oxygen). Carbon(II) oxide also takes up the space that would otherwise be occupied by oxygen. So a person poisoned by this gas is in effect, suffering from lack of oxygen. It is dangerous because it has no odour or colour and thus cannot be easily detected.

If you come across a victim of carbon(II) oxide poisoning give him or her artificial respiration (Fig. 6.32) or "Kiss of life" in the open air. If possible take the victim to hospital because they might need to be given pure oxygen, or a blood transfusion.

Fig. 6.32: Kiss of life

6.5 Carbonates and hydrogencarbonates

Experiment 6.9

To investigate the action of heat on carbonates and hydrogencarbonates.

Apparatus and chemicals

- test-tubes
- calcium hydroxide solution
- distilled water
- red and blue litmus papers
- copper carbonate
- magnesium carbonate
- zinc carbonate
- lead carbonate
- sodium carbonate
- potassium carbonate
- sodium hydrogencarbonate
- ammonium carbonate

Procedure

1. Place a spatula full of the carbonate or hydrogencarbonate salts mentioned above in different test-tubes then heat.
2. Test any gas formed with calcium hydroxide solution and moist litmus papers (red and blue).

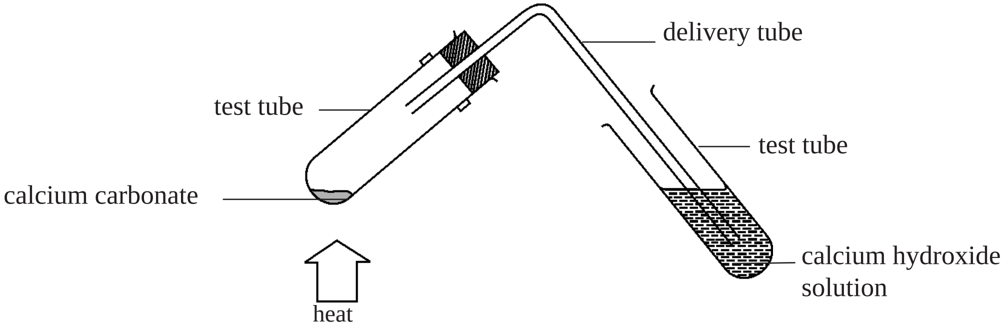


Fig. 6.33 Action of heat on carbonates and hydrogencarbonates

- What is the colour of the residue when hot and when cold?

3. Record all your observations as illustrated by Table 6.3.

Table 6.3 Heating of carbonates and hydrogencarbonate

Name of carbonate or hydrogencarbonate	**Colour before heating**	**Colour after heating**	**Colour on cooling**	**Effect on calcium hydroxide solution**	**Effect on red or blue litmus paper**
Sodium carbonate					
Potassium carbonate					
Magnesium carbonate					
Zinc carbonate					
Lead carbonate					
Copper carbonate					
Ammonium carbonate					
Sodium hydrogencarbonate					

- Write the chemical equations for the reactions.
- List the carbonates which behaved the same.
- Which carbonate evolved an alkaline gas?
- Which carbonate produced a residue that showed colour changes?
- Which carbonates were easily decomposed by heat?

It will be observed that some carbonates decompose when heated and others do not.

Carbonates of:

K
Na } do not decompose on heating.

Mg
Zn
Pb } A metal oxide and carbon(IV) oxide are produced on heating.
Cu

$(NH_4)_2CO_3$ produces carbon(IV) oxide, water and ammonia gas on heating.

Therefore, the carbonates of metals high in the reactivity series do not decompose.

Ease of decomposition of metal carbonates increases as we go down the reactivity series.

The following are equations showing decomposition of carbonates.

Copper carbonate → copper(II) oxide + carbon(IV) oxide
$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
(green) (black)

Magnesium carbonate → magnesium(II) oxide + carbon(IV) oxide
$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$
(white) (white)

Zinc carbonate → zinc oxide + carbon(IV) oxide
$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$
(white) (white)

Lead carbonate → lead(II) oxide + carbon(IV) oxide
$PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$
(white) (yellow)

Calcium carbonate → calcium oxide + carbon(IV) oxide
$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
(white) (white)

Copper(II) oxide is black, while oxides of magnesium and calcium are white. Zinc oxide is yellow when hot and white on cooling. Lead(II) oxide is red-brown when hot and yellow on cooling. Sodium carbonate and potassium do not decompose on heating. Sodium hydrogencarbonate and calcium hydrogencarbonate decompose on heating to form a carbonate, water and carbon(IV) oxide.

Sodium hydrogencarbonate → sodium carbonate + carbon(IV) oxide + water
$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

Calcium hydrogencarbonate → calcium carbonate + carbon(IV) oxide + water
$Ca(HCO_3)_2(s) \rightarrow CaCO_3(s) + CO_2(g) + H_2O(l)$

When calcium hydrogencarbonate is heated strongly calcium carbonate formed decomposes to form calcium oxide and carbon(IV) oxide.

Ammonium carbonate decomposes on heating to form ammonia gas, water and carbon(IV) oxide.

Ammonium carbonate → ammonia + carbon(IV) oxide + water

$$(NH_4)_2CO_3(s) \rightarrow 2NH_3(g) + CO_2(g) + H_2O(l)$$

Ammonia and carbon(IV) oxide gases are liberated at the same time. If moist red and blue litmus papers are put together at the mouth of the test-tube when ammonium carbonate is heated, ammonia gas is detected first (red litmus paper turns blue).

Discuss with the teacher why the ease of decomposition of metal carbonates increases as one goes down the reactivity series.

Note: Sometimes sodium carbonate may be contaminated with sodium hydrogencarbonate and therefore carbon(IV) oxide can be obtained on heating it.

Experiment 6.10

To investigate the action of dilute acids on carbonates and hydrogencarbonates

Apparatus and chemicals

- test tubes
- calcium hydroxide
- calcium carbonate
- magnesium carbonate
- zinc carbonate
- lead carbonate
- copper carbonate
- sodium carbonate
- ammonium carbonate
- sodium hydrogencarbonate
- calcium hydrogencarbonate
- dilute nitric acid
- dilute sulphuric acid
- dilute hydrochloric acid

Procedure

1. In different test-tubes put a spatula full of the above carbonates and hydrogencarbonates.
2. Add dilute hydrochloric acid in each test tube and test the gas evolved with calcium hydroxide solution as shown in Fig. 6.34.
    - What happens when the acid is added to the carbonates or hydrogencarbonates?

- What happens to the calcium hydroxide solution?

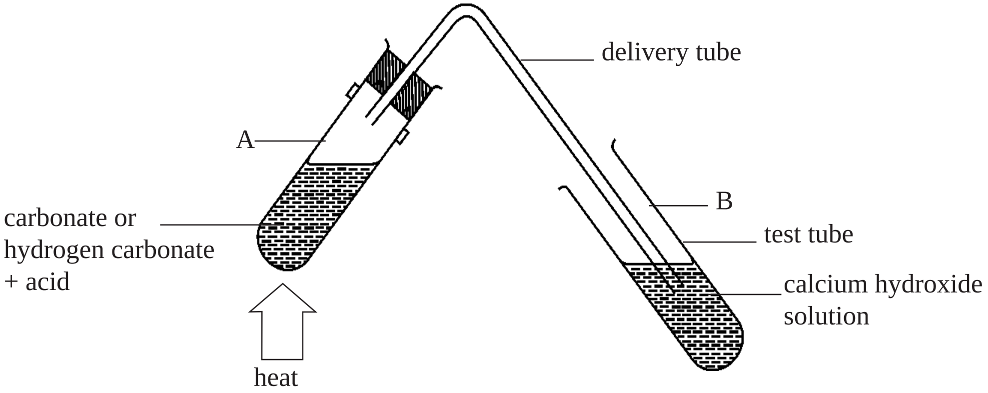


Fig. 6.34 Reaction of carbonates or hydrogencarbonates with acid

3. Repeat the experiment using sulphuric acid and then dilute nitric acid on all the carbonates and hydrogencarbonates.
    - What happens when dilute sulphuric acid is added to calcium hydrogencarbonate or calcium carbonate?

4. Copy Table 6.14 and record your observations.

You will notice that all carbonates liberate carbon(IV) oxide on addition of a dilute acid. They also form a salt and water. For example:

Calcium carbonate + nitric acid → calcium Nitrate + water + carbon(IV) oxide

$$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)(aq) + H_2O(l) + CO_2(g)$$

Calcium carbonate + hydrochloric acid → calcium chloride + water + carbon(IV) oxide

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Write equations for the other reactions.

All hydrogencarbonates also liberate carbon(IV) oxide on addition of an acid. For example:

Sodium hydrogen carbonate + nitric acid → sodium nitrate + carbon(IV) oxide + water

$$NaHCO_3(s) + HNO_3(aq) \rightarrow NaNO_3(aq) + CO_2(g) + H_2O(l)$$

Sodium hydrogencarbonate + hydrochloric acid → sodium chloride + carbon(IV) oxide + water

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Table 6.14

Carbonate /hydrogen carbonate	**Acid**	**Observation in test-tube A**	**Observation in test-tube B**
Sodium hydrogen carbonate	hydrochloric		
	sulphuric		
	nitric		
Sodium carbonate	hydrochloric		
	sulphuric		
	nitric		
Calcium hydrogen carbonate	hydrochloric		
	sulphuric		
	nitric		
Calcium carbonate	hydrochloric		
	sulphuric		
	nitric		
Magnesium carbonate	hydrochloric		
	sulphuric		
	nitric		
Zinc carbonate	hydrochloric		
	sulphuric		
	nitric		
Ammonium carbonate	hydrochloric		
	sulphuric		
	nitric		
Lead carbonate	hydrochloric		
	sulphuric		
	nitric		
Copper carbonate	hydrochloric		
	sulphuric		
	nitric		

What was observed when dilute sulphuric acid was added to calcium carbonate and lead carbonate or dilute hydrochloric acid to lead carbonate? You must have noticed that the reaction started and then stopped after a very short time. Explain the reason why (refer to the laboratory preparation of carbon(IV) oxide).

The products formed from these reactions are insoluble and form a coating on the surface of the carbonate preventing any further reaction. Write equations for these reactions and identify the insoluble salts.

We can conclude that all carbonates and hydrogencarbonates react with acids to form a salt, water and carbon(IV) oxide.

Carbonate + acid → salt + carbon(IV) oxide + water

Hydrogencarbonate + acid → salt + carbon(IV) oxide + water

Note: All carbonates are insoluble in water, except sodium, potassium and ammonium carbonates. All hydrogencarbonates are soluble in water.

6.6 Production and manufacture of sodium carbonate

I. Production of sodium carbonate (soda ash) at Magadi Soda Company

The most efficient means of trona extraction uses *dredging techniques*. Both a *bucket dredge* and a *cutter-suction* dredge are used for this purpose, concentrating on the top 4 metres or so of the solid crust where the purest material is found.

As the solid material is removed from the surface, *liquor* (unwanted solution) from the surrounding trona drains into the resulting cavity, forming a pool or "paddock" in which dredgers can float. The crystals of trona produced by the dredgers are mixed with liquor and pumped as *slurry* back to the ash plant. Here the liquor is discarded and the crystal washed and centrifuged. The dump crystals are then fed into *calciners* (kilns). Here the residual moisture, water crystallisation and carbon(IV) oxide gas are driven off to leave normal sodium carbonate (soda ash).

$$\text{Trona} \xrightarrow{\text{heat}} \text{Sodium carbonate} + \text{Carbon(IV) oxide} + \text{Water}$$

$$2(Na_2CO_3.\ NaHCO_3.\ 2H_2O)(s) \longrightarrow 3Na_2CO_3 + CO_2(g) + H_2O(l)$$

The soda ash from the calciners passes through rotary drum coolers before entering the grinding and screening plant. Oversize material is removed and grounded again. The final product is conveyed to silos where it is packed or loaded directly into specially bulk raid hopper wagons.

The liquor from this process is used to extract sodium chloride. Solar evaporation process is used. Most of the Magadi sodium chloride is sold within Kenya for livestock consumption or for industrial purposes.

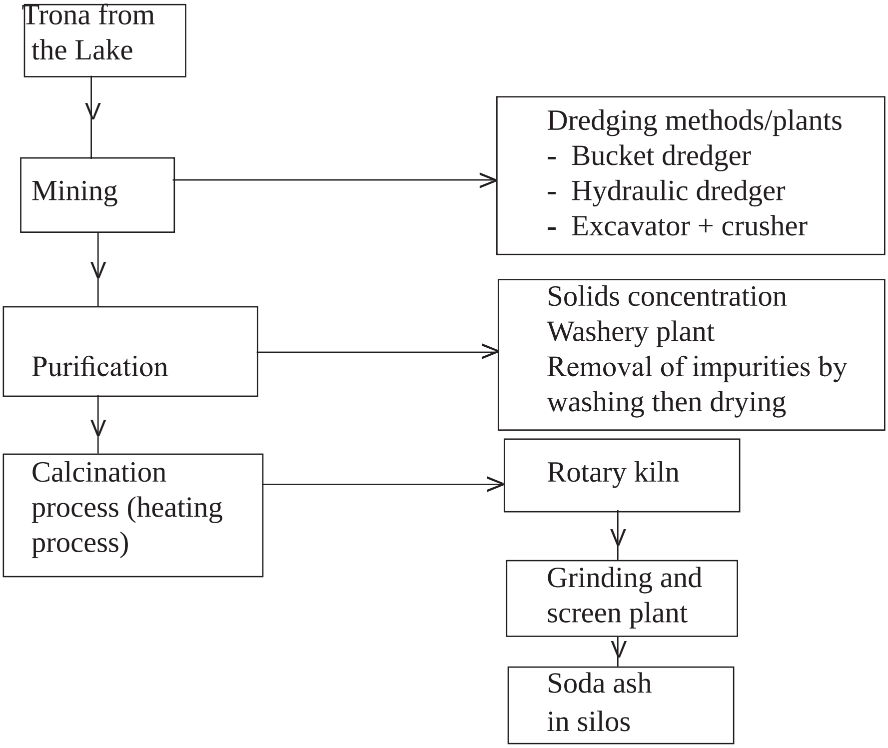


Fig 6.35 A flow chart of soda ash manufacturing process

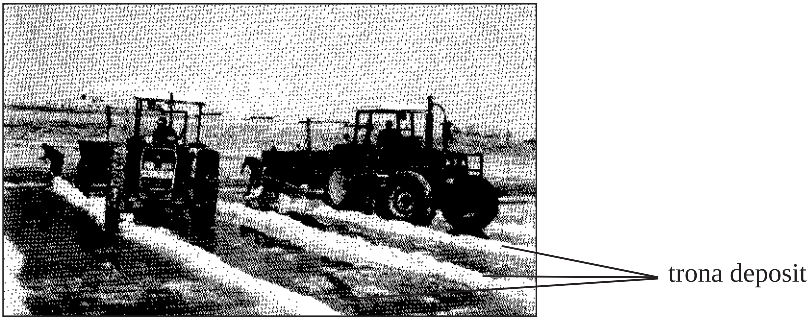


Fig 6.36 White soda ash of Lake Magadi

II. Manufacture of sodium carbonate (soda ash) by Solvay process

The main *raw materials* are **sodium chloride** and **calcium carbonate.**
Fig. 6.36. Concentrated brine (sodium chloride solution) is saturated with ammonia in a tower. The ammoniacal brine formed is run down the Solvay tower. Carbon(IV) oxide is forced into the tower from the bottom. The towers are filled with *baffles*. These baffles make the liquid flow slowly and increase the surface area for reaction. Sodium hydrogen carbonate is formed in Solvay tower. The reaction occurring is as follows.

Sodium chloride + ammonia + carbon(IV) oxide + water → sodium hydrogencarbonate + ammonium chloride

$NaCl(g) + NH_3(aq) + CO_2(g) + H_2O(I) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$

Then sodium hydrogencarbonate is filtered off, and washed to remove ammonium chloride. It is then heated in a furnace roaster to give sodium carbonate and carbon(IV) oxide.

Sodium hydrogencarbonate → sodium carbonate + carbon(IV) oxide + water

$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(l)$

Carbon(IV) oxide formed is recycled to the Solvay tower. The other source of carbon(IV) oxide is from heating calcium carbonate(limestone) in the kiln where it dissociates into calcium oxide and carbon(IV) oxide.

Calcium carbonate → calcium oxide + carbon(IV) oxide

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

The calcium oxide formed from above reaction is ‘slaked’ by addition of water.

Calcium oxide + water → calcium hydroxide

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$

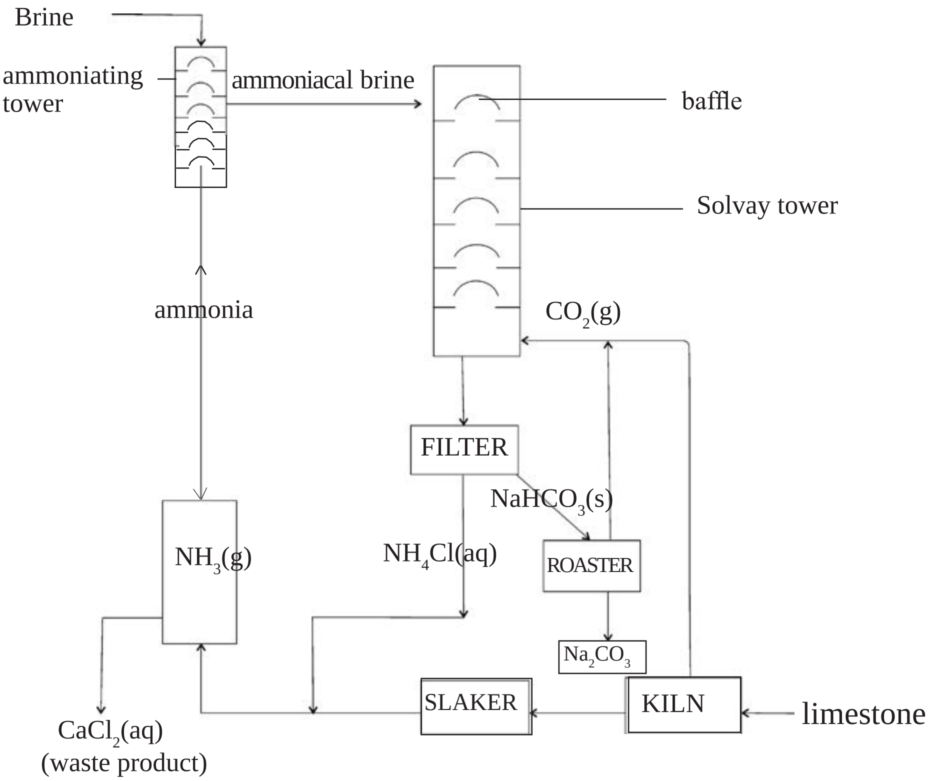


Fig. 6.36 Solvay process

The calcium hydroxide formed is heated with the filtrate, ammonium chloride from the solvay tower, to produce ammonia.

Ammonium chloride + calcium hydroxide → calcium chloride + ammonia + water

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

The ammonia from this reaction is returned to the ammoniating tower.

A careful examination of the flow diagram (Fig. 6.36) shows that the only waste product in this process is calcium chloride. The carbon(IV) oxide and ammonia are recycled.

Therefore, the solvay process is very efficient. The raw materials are cheap and readily available and only one waste product is formed.

Uses of sodium carbonate

1. Manufacture of glass.
2. In domestic water-softening.
3. In the manufacture of chemicals e.g. sodium hydroxide.
4. In the manufacture of laundry detergents.
5. In the paper-making process.
6. Textiles.

6.7 Importance of carbon and its oxides

Carbon cycle

All the processes which involve the addition of carbon(IV) oxide to the atmosphere and its removal from the atmosphere are summed up as the carbon or *carbon(IV) oxides* cycle, illustrated in Fig. 6.37. The number of carbon atoms on Earth is constant. It simply circulates. Carbon(IV) oxide is the “transfer agent” for carbon circulation.

The processes which *add* carbon(IV) oxide to the atmosphere are:

1. **Combustion**: Petroleum, coal, wood, charcoal, wax or any other organic compound.
2. **Respiration:** During this process the sugar in the animal bodies is changed by oxygen to carbon (IV) oxide and water. Carbon(IV) oxide is breathed out to the atmosphere.
3. **Decay of plants and animals:** Any carbon in plants and animals is converted to carbon(IV) oxide in the decaying process if sufficient oxygen supply is available.
4. **Making of beer and wine:** Yeast changes sugar to ethanol and carbon(IV) oxide. The process is called *fermentation.*

$$\text{Sugar} \xrightarrow{\text{Yeast}} \text{ethanol} + \text{carbon(IV) oxide}$$
$$C_6H_{12}O_6 \xrightarrow{\text{Yeast}} 2C_2H_5OH(aq) + 2CO_2(g)$$

The carbon(IV) oxide formed finds its way to the atmosphere.

5. In areas where there are factories manufacturing calcium oxide from calcium carbonate, carbon(IV) oxide is formed as a by-product. If not collected for use it finds its way to the atmosphere.

$$\text{Calcium carbonate} \rightarrow \text{calcium oxide} + \text{carbon(IV) oxide}$$
$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Carbon(IV) oxide is removed from the atmosphere by:

1. **Water**: Carbon(IV) oxide dissolves in rivers and lakes to form carbonic acid. The carbonic acid reacts with calcium carbonate and magnesium carbonate to form calcium hydrogen carbonate and magnesium hydrogencarbonate respectively.

$$\text{Carbon(IV) oxide} + \text{water} \rightarrow \text{carbonic acid}$$
$$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

$$\text{Calcium carbonate} + \text{water} + \text{carbon(IV) oxide} \rightarrow \text{calcium hydrogen carbonate}$$
$$CaCO_3(s) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$$

$$\text{Magnesium carbonate} + \text{Water} + \text{Carbon(IV) oxide} \rightarrow \text{Magnesium hydrogen carbonate}$$
$$MgCO_3(s) + H_2O(l) + CO_2(g) \rightarrow Mg(HCO_3)_2(aq)$$

The hydrogen carbonates cause temporary hardness in water. This hard water is used by animals to make shells.

2. **Photosynthesis:** In this process, green plants use carbon(IV) oxide to make sugar using sunlight as a source of energy.

$$\text{Carbon(IV) oxide} + \text{water} \rightarrow \text{s ugar} + \text{oxygen}$$
$$CO_2(g) + H_2O(l) \rightarrow C_6H_{12}O_6(aq) + O_2(g)$$

3. **Calcium hydroxide**: Natural calcium hydroxide slowly reacts with carbon(IV) oxide from the air to form calcium carbonate

$$\text{Calcium(II) hydroxide} + \text{carbon(IV)oxide} \rightarrow \text{calcium carbonate} + \text{water}$$
$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

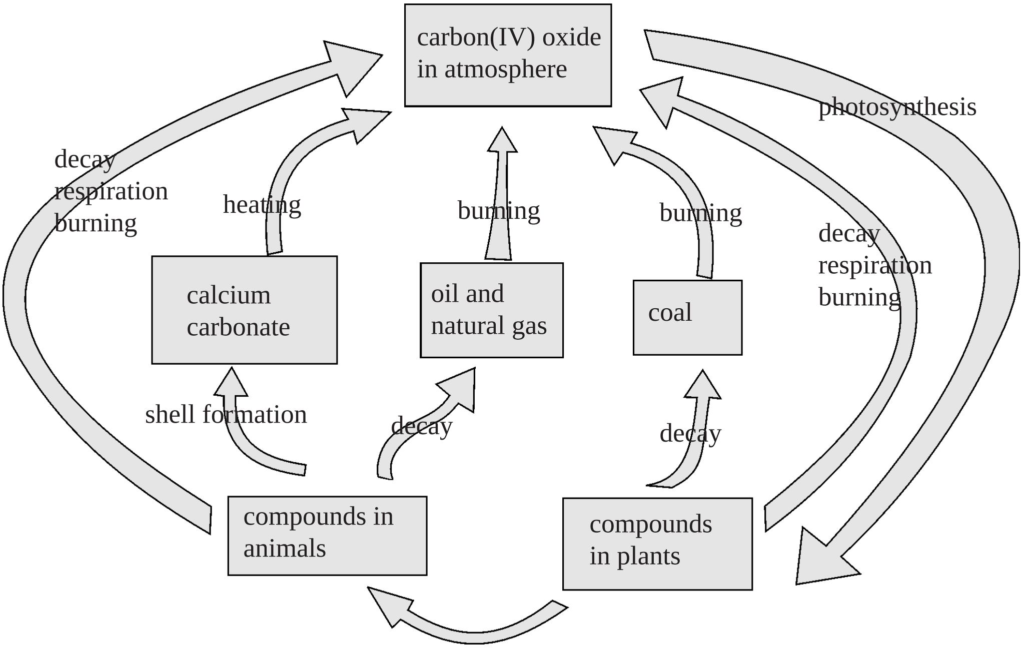


Fig. 6.37 Carbon cycle

Effects of Carbon(IV) oxide and carbon (II) oxide on the environment

Carbon(IV oxide, greenhouse effect and weather

This topic will be introduced by performing the following simple experiment.

Experiment 6.22

To demonstrate the working of a greenhouse

Apparatus and chemicals

- a large glass beaker or glass box.
- two small plastic beakers full of water.
- thermometer.

Procedure

1. Place a small plastic beaker full of water under a big glass beaker or a glass box. Place and another one outside in the open.
2. Leave them in the sunlight.
3. At the end of the day measure the temperature of the water in the two beakers.
   - Which has a higher temperature?

   The temperature of the water inside the glass beaker or box is higher. This is because the heat from the sun passes through glass to the earth surface and is reflected back to the glass. Some of the heat energy is trapped by the glass and sent back to the earth. This is the heat which warms the water which is inside the big glass/box.

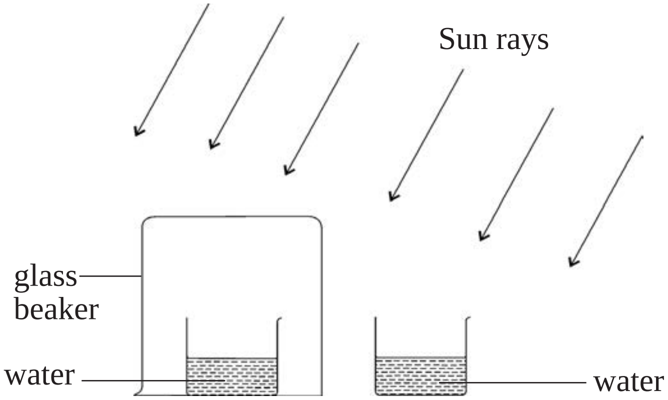


Fig 3.8 Demonstrating greenhouse

Greenhouse effect and weather

Generally when we talk about weather, we are talking about temperature, rain or hail storms, typhoons and so on.

Temperature is the amount of heat found in the atmosphere. This heat in the atmosphere comes from the sun. Each day a small amount of the sun's total heat energy reaches the top layers of earth's atmosphere. The earth then maintains heat energy balance so that not all the heat coming to the earth remains. In time, the heat energy held in the atmosphere and in earth's surface escapes back into space.

Then earth's atmosphere acts like the glass roof of a greenhouse. In a greenhouse the sun's rays pass through the glass windows and warm the plants as seen in Fig 6.39.

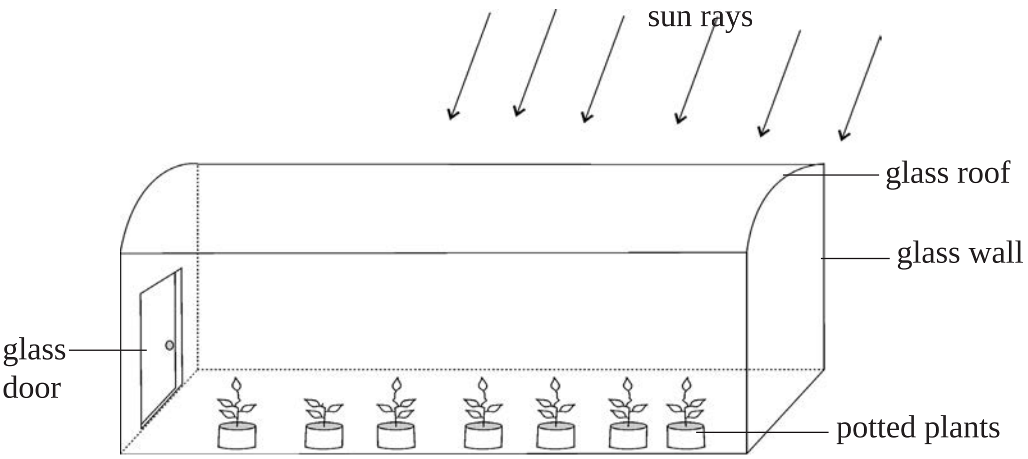


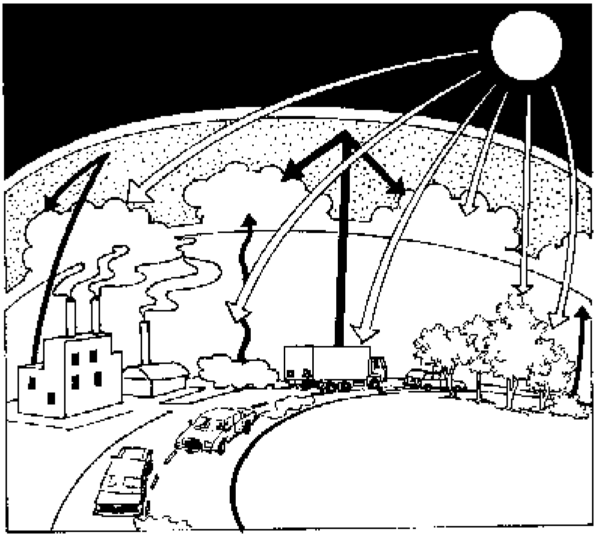
Fig. 6.39 A greenhouse

The sun’s rays pass through the space, and onto the soil, rocks and water bodies such as ocean, lakes and rivers. The rays are reflected back into space in form of infrared radiations which do not easily escape into atmosphere. This is because some gases like carbon(IV) oxide in the lower part of the atmosphere take in some of the heat energy and send it back to earth. As a result the temperature of the surface of the earth is raised thereby preventing excessive cooling on the earth. This is called the **greenhouse effect**.

There is an increasing concern among scientists that large amounts of carbon(IV) oxide in the atmosphere are changing the green house effect. This build up of carbon(IV) oxide has resulted from increased burning of coal, oil, natural gases, firewood, etc and extended cutting of trees.

Trees purify the atmosphere by absorbing carbon(V) oxide and releasing oxygen. When a lot of trees are cut down, the carbon(IV) oxide they would have absorbed is released to the atmosphere causing an increase in the earth’s temperature. This heat has caused melting of ice in parts of the north and south poles and subsequent rise in sea level. As a result it has been observed that ocean waters have risen in some parts of the world causing shoreline areas to flood.

The result is that more areas in the high altitudes can be farming but with more in the low altitude areas being rendered useless for farming because of flooding. Also, changing temperatures have destroyed plant and animal life on sea and land, as some organisms cannot survive in areas of extreme temperatures. Such changes have also interfered with the world climate causing unexpected heavy rains and droughts in many parts of the world.



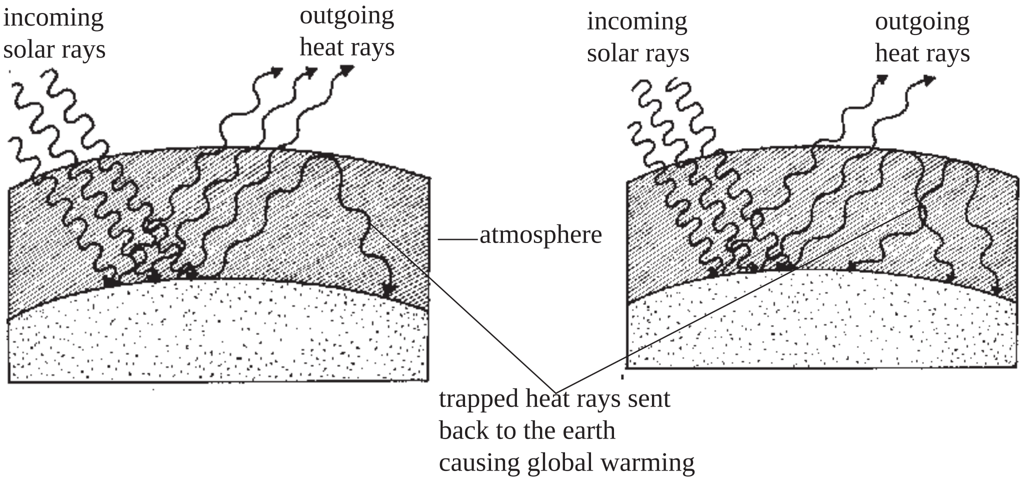


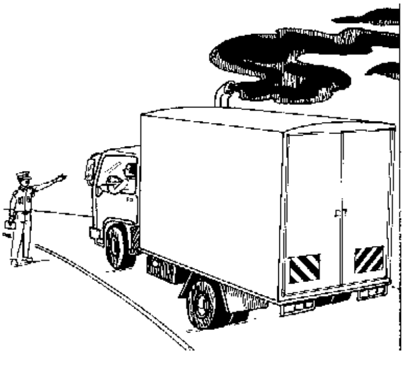
Fig.6.40 Greenhouse effect

Planting of trees and other plants also help in minimizing the amount of carbon(IV) oxide during photosynthesis. Factories whose industrial processes involve emission of carbon(IV) oxide should be encouraged to recycle the gas if necessary or collect it for other uses instead of releasing it to the atmosphere. Fig. 6.41 d

Effects of carbon(II) oxide on the environment

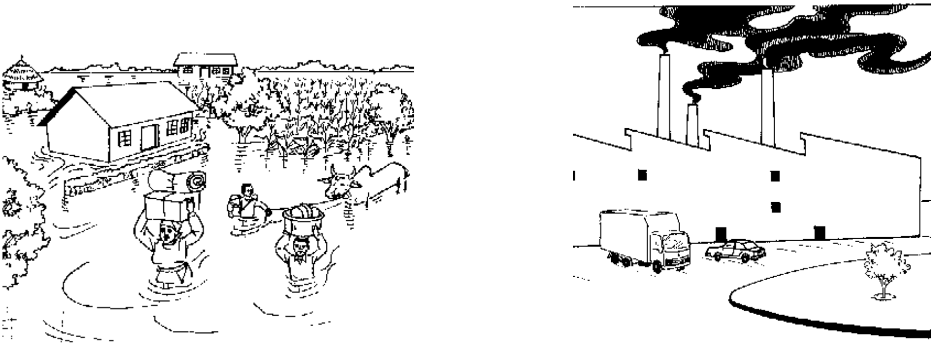
Carbon(II) oxide, as we have discussed earlier, is a very poisonous gas. It has no colour, taste or smell, thus making it hard for anyone to detect its presence. It is formed during respiration by incomplete combustion of charcoal, coal, natural (methane) gas, fires, firewood in poorly ventilated areas. and proper ventillation in rooms, industrial areas, etc, helps in preventing the build up of carbon (IV) oxide

This gas also comes from car exhausts and diesel engines (insufficient supply of air in these engines). Combustion of these fuels should therefore be stored or alternative fuels which do not emit this gas used.



(a) Smoke from vehicles

(b) Smoke from burning tyres



(c) Flooding

(d) Smoke from industry

Fig. 6.41 Causes and effects of carbon(IV) oxide in the environment

Reducing pollution

It is important that we reduce the level of air pollution for a healthier environment. The following are some of the guidelines that will ensure that our environment is kept clean and safe for our future generations.

1. Laws in the cities and other urban areas should restrict the amount of smoke and other waste products released by factories which pollute the air. Recycling should be encouraged as much as possible.
2. Burning of rubbish should be restricted such that it is done only in designated areas.
3. Vehicles should have their exhaust pipes fitted with filters to trap impurities and other substances that contaminate the environment.

Summary

1. Carbon exists naturally mainly in form of two crystalline allotropes namely, *diamond* and graphite.
2. Carbon can be used to reduce some metallic oxides.
3. Carbon(IV) oxide is usually prepared by the reaction between a carbonate and an acid.
4. A carbonate used to prepare carbon(IV) oxide forms a soluble salt of the acid used.
5. Insoluble or sparingly soluble salts formed when a carbonate salt reacts with an acid form a coating/layer on the unreacted carbonate and the reaction ceases.
6. In an insufficient supply of oxygen, carbon burns to form carbon(II) oxide.
7. Carbon(IV) oxide forms a white precipitate of calcium carbonate when it reacts with calcium hydroxide solution.
8. When carbon(IV) oxide is passed for a long time in calcium hydroxide solution,

the insoluble precipitate forms a soluble calcium hydrogen carbonate and therefore the precipitate dissolves.

9. Magnesium continues to burn in carbon(IV) oxide because it is more reactive towards oxygen than carbon. It displaces oxygen from carbon(IV) oxide. This oxygen supports the combustion of magnesium.
10. Carbonic acid is a *dibasic acid* and therefore it can form a normal and an acid salt.
11. Solid carbon(IV) oxide is usually known as *dry ice*. It sublimes at room temperature leaving no residue like ordinary ice.
12. Carbon(IV) oxide can be reduced by red-hot charcoal (carbon) to from carbon(II) oxide.
13. In the laboratory carbon(II) oxide is prepared by dehydration of methanoic acid (formic acid).
14. Carbon(II) oxide is combustible. It burns with a blue flame to form carbon(IV) oxide.
15. Carbon(II) oxide can also be used as a reducing agent. When reducing an oxide it forms carbon(IV) oxide.
16. Water-gas is a mixture of carbon(II) oxide and hydrogen. It is produced by passing steam over red-hot coke.
17. Most metallic carbonates decompose on heating to form a metallic oxide and carbon(IV) oxide.
18. Ammonium carbonate decomposes on heating to form ammonia, water and carbon(IV) oxide.
19. Ammonia gas turns moist red litmus paper blue. It is the only alkaline gas known to man.
20. All carbonate salts are insoluble except sodium, potassium and ammonium carbonates.
21. When heated, hydrogencarbonates form carbonate salts, water and carbon(IV) oxide.
22. All carbonates and hydrogencarbonates react with acids to form a salt of the acid, water and carbon(IV) oxide.
23. In the Solvay process the only raw materials required are common salt, limestone and ammonia gas.
24. The only waste in the Solvay process is calcium chloride solution.
25. Carbon or carbon(IV) oxide cycle includes all the processes which involve addition of carbon(IV) oxide to the atmosphere and the removal of it from the atmosphere.
26. Carbon(IV) oxide contributes to global warming, while carbon(II) oxide is a very poisonous gas.

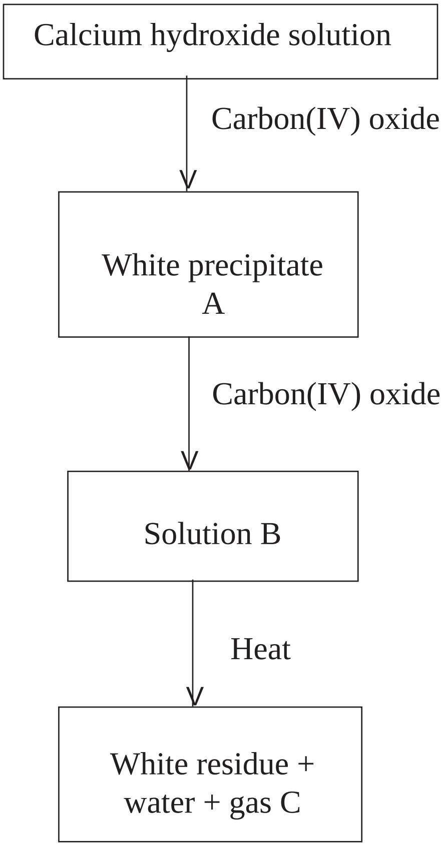
Projects

- Make a collection of carbonate and hydrogencarbonate bearing minerals from your locality. Check with dilute acids whether they all produce carbon(IV) oxide.
- Prepare some baking powder in the laboratory. Involve your chemistry and/or home science teacher.
- Prepare a chart on carbon-cycle. Use it to give a talk to your class on how carbon atoms are kept constant on earth by this cycle.
- Examine, with the help of your teacher, how a fire extinguisher in your laboratory works. Construct a simple fire extinguisher.

Revision Exercise 6

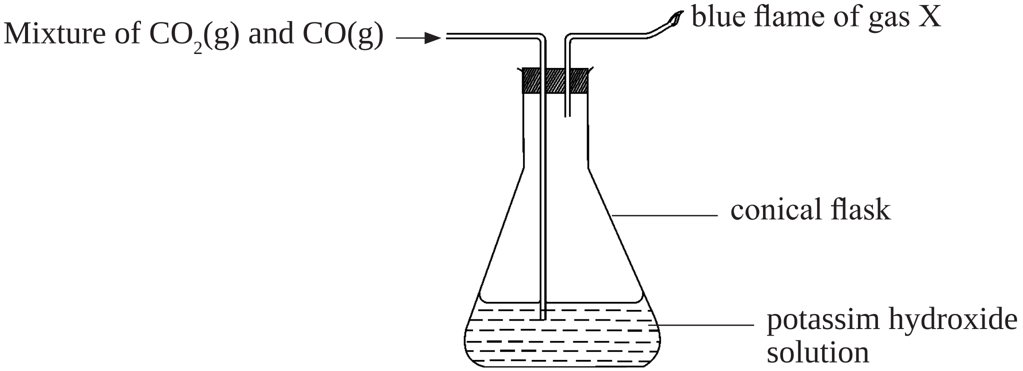
1. You are provided with the following substances: sodium chloride, wax, copper carbonate and copper(II) oxide. Which of the above listed substances will provide the most useful information on the substance heated? Explain.

2. Study the flow chart below and answer the questions that follow:



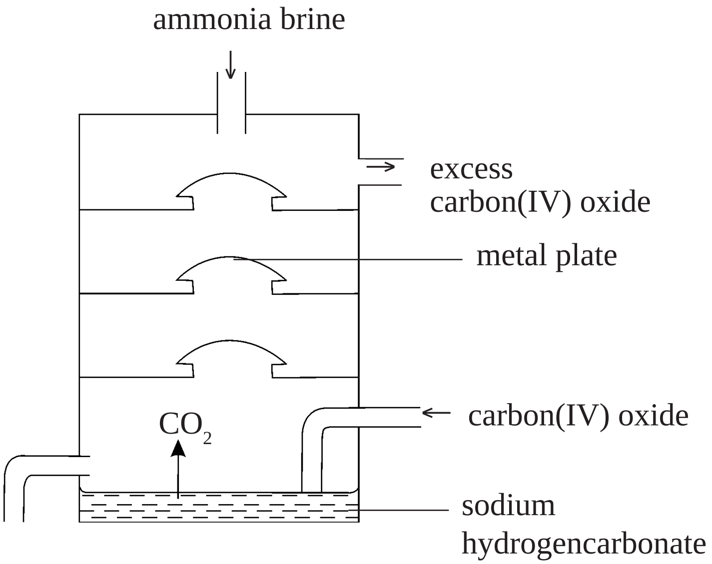
(i) Name the white precipitate A.
(ii) Name solution B.
(iii) Name gas C.

3. A mixture of carbon(IV) oxide and carbon(II) oxide are passed through potassium hydroxide solution as shown in the following set-up.



(i) Name gas X.
(ii) Write the equation of the reaction that takes place:
    (a) In the conical flask in the first 30 seconds.
    (b) In the burning of gas X.

4. The diagram below shows a simple ammonia-soda tower used in the manufacture of sodium carbonate:

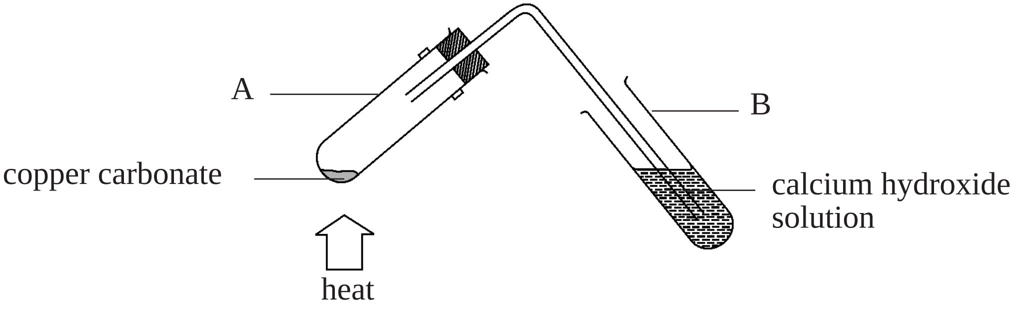


(i) Name the raw materials needed in the above process.

(ii) Write an equation for the reaction taking place in the tower.
(iii) Give a reason for having the circular metal plates in the tower.
(iv) Name the gases recycled in the process illustrated above.
(v) Write equations to show how the two gases mentioned in (iv) above are generated.
(vi) Outline the commercial uses of sodium carbonate.

5. State two properties that make carbon(IV) oxide suitable for use as a fire extinguisher.

6. A student set up the apparatus shown below to investigate the effect of heat on a sample of copper carbonate.



(i) State the observations made in each test-tube.
(ii) Write the equation for the reaction that occurs in tube A.

7. (i) Diamond and graphite are allotropes of carbon.
Explain why graphite is used as a lubricant whereas diamond is used as an abrasive.
(ii) Why is graphite a better lubricant than oil?
(iii) State two uses of carbon(II) oxide.

8. If two unlabelled black powders were placed before you on the laboratory bench and you are told that they were copper(II) oxide and charcoal, what simple test would you perform to distinguish them?

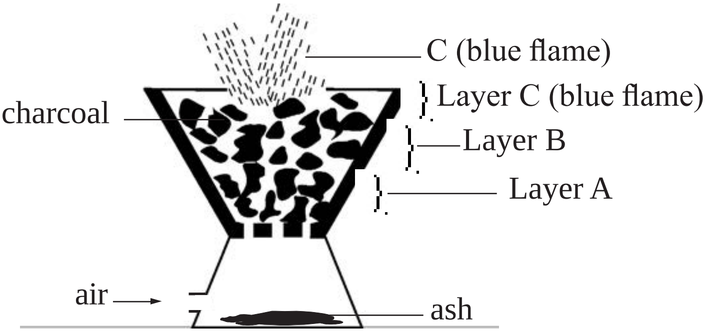
9. (i) When carbon is burned in air, two gases are formed.
Give the conditions which favour the formation of each gas.
(ii) How can you differentiate between carbon(IV) oxide and carbon(II) oxide in the laboratory?
(iii) Why is "dry ice" preferred to ordinary ice in refrigeration?

10. (i) A student from Kilungu High School mixed up the labels of solid sodium

carbonate and solid sodium hydrogen carbonate. How did she go about differentiating the two solids?

(ii) Using magnesium sulphate solution, how can you differentiate between a solution of sodium carbonate from a solution of sodium hydrogen carbonate?

11.



The above figure shows a diagram of a burning jiko in a room which has sufficient supply of air.

Explain what happens around

(i) Layer A
(ii) Layer B
(iii) Area C

Use chemical equations in the three cases to support your answers.

SELF CHECK TEST PAPER 1

1. Balance the following chemical equations

   (a) $Zn(s) + O_2(g) \rightarrow ZnO(s) + SO_2(g)$

   (b) $Pb(NO_3)_2(s) \rightarrow PbO(s) + NO_2(g) + O_2(g)$

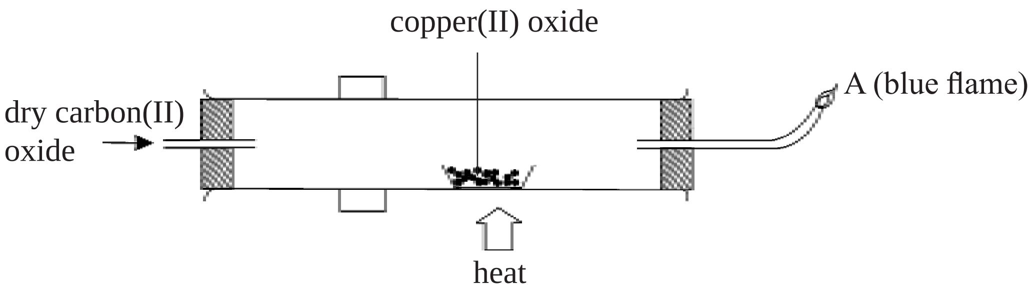
   (c) $FeSO_4(s) \rightarrow Fe_2O_3(s) + SO_2(g) + SO_3(g)$

   (d) $KOH(aq) + Cl_2(g) \rightarrow KCl(aq) + KClO_3(aq) + H_2O(l)$

2. Determine the relative atomic mass (r.a.m. or A_r) of the element whose isotopic mixture occurs in the given proportions.

Mass of isotope	39.0	40.0	41.0
Percentage proportion	92.0	2.0	6.0

3. Excess carbon(II) oxide gas was passed over a heated sample of copper(II) oxide as shown below for five minutes.



   (a) What observation is made in the combustion tube after five minutes?

   (b) Write a balanced chemical equation of the reaction which took place in the combustion tube?

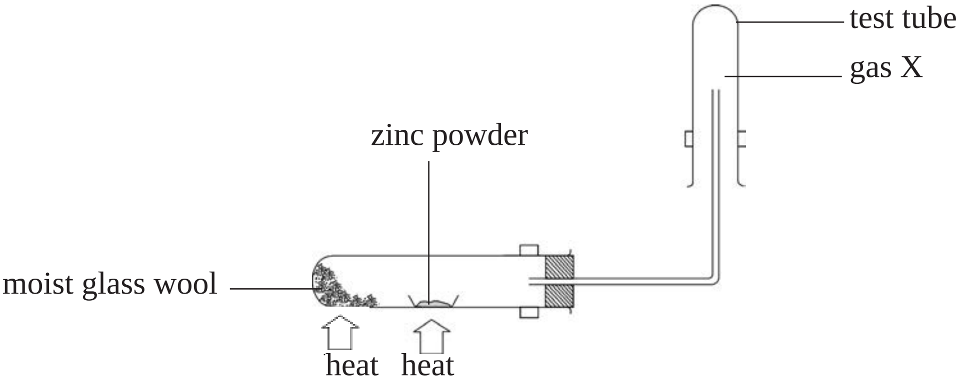
   (c) What is burning at A and why should it be burnt?

   (d) Write a chemical equation of the reaction taking place at A.

   (e) (i) Which other gas burns with a characteristic blue flame?

   (ii) How can you differentiate the two gases?

4. A student set up the experiment below to collect gas X. The glass wool was heated before heating the zinc powder.



(a) Why was it necessary to heat moist glass wool before heating zinc powder?
(b) What would happen if the zinc powder was heated before heating the glass wool?
(c) What property of gas X makes it possible for it to be collected as shown in the diagram?
(d) Write the chemical equation for the reaction taking place in the boiling tube.
(e) State the colour changes that would take place in the boiling tube on heating and cooling.

5. A few drops of silver nitrate solution were added to sodium chloride solution in a test-tube.
   (a) What observation was made?
   Explain.
   (b) Write a balanced chemical equation for the reaction.
   (c) Write the ionic equation for the reaction.

6. Classify each of the following as element, compound or mixture:
sugar, ink, copper(II) sulphate crystals, milk, copper, neon, sand.

Element	**Compound**	**Mixture**

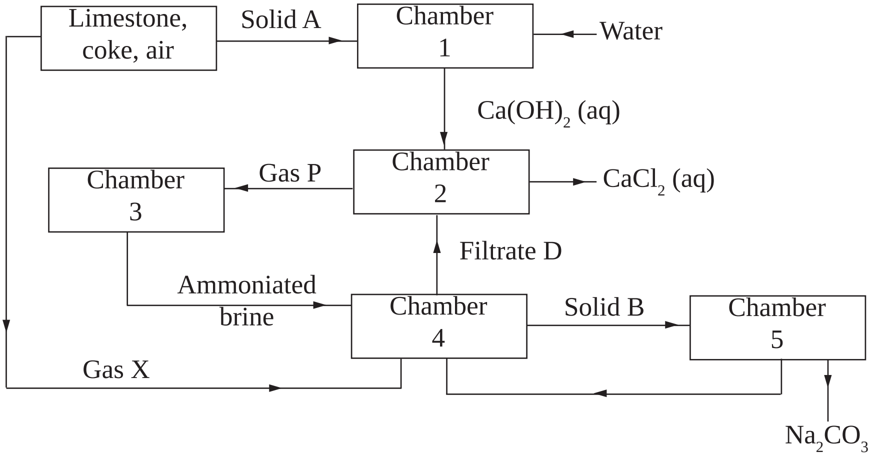
7. An atom of element X has 15 protons and mass number of 31 in its nucleus while an atom of element Y has 17 and a mass number of 35.
   (a) What are the atomic numbers of these elements?
   (b) How many electrons are in the outermost energy levels of these atoms?
   (c) What type of bond exists when X and Y combine?

8. An ion P^{2-} has electron arrangement of 2.8.
   (a) What is the atomic number of the element?
   (b) To which group and period in the periodic table does the element belong?
   (c) What type of bonds exist between the atoms of element P?
   (d) What type of bonds exist between atoms of element X whose ions are represented by X^{2+}.
   (e) Name the solution and the catalyst used in the preparation of element P in the laboratory.
   (f) What is the general name of elements which appear in the group after group of P?

SELF CHECK TEST PAPER 2

1. (a) Explain why solutions of chlorides of silicon and phosphorus react with sodium carbonate to form carbon (IV) oxide. Use chemical equations where necessary.
   (b) Explain why chloride of silicon and phosphorus produce fumes in moist air.

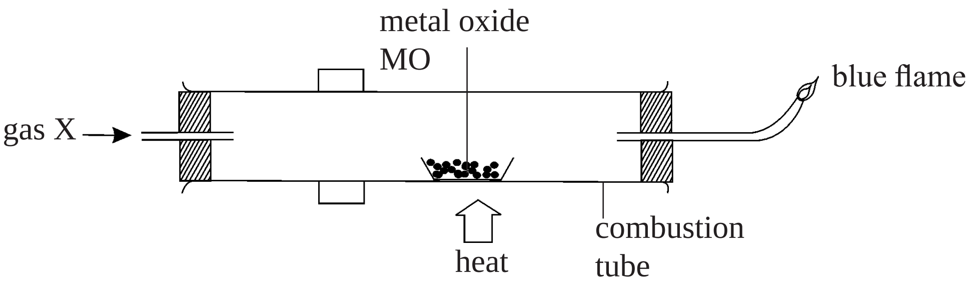
2. Study the flow chart below and answer the questions that follow.



   (a) Name the following substances
      (i) Gas P
      (ii) Gas X
      (iii) Solid A
      (iv) Solid B
      (v) Filtrate D
   (b) Write chemical equations for the reactions that occur in
      (i) Chamber 2
      (ii) Chamber 4

3. Name the catalysts used in the following chemical reactions.
   (a) Decomposition of hydrogen peroxide to oxygen and water.
   (b) Hardening of oils to form solid fats.
   (c) Preparation of hydrogen gas from dilute sulphuric acid and zinc metal.

4. The apparatus below can be used to obtain a red-brown metal M from its oxide MO which is black. Gas X, the lightest gas known is used.



(a) Name gas X.
(b) Name metal oxide, MO
(c) Write the chemical equation of the reaction taking place in the combustion tube. (symbol of gas = X, symbol of metal = M).
(d) When one stops heating, it is advisable to let a stream of gas X to pass through the combustion tube as the solid residue cools. Give the reason for doing so.
(e) What is the reason for igniting the gas coming out of the combustion tube?

5. Lead metal and molten lead(II) bromide are known to conduct electric current. Outline the difference in their mode of conduction.

6. The products formed by the action of heat on nitrates of metals P, Q and R are as shown in the table below:

Nitrate of metal	**Products formed**
P	Metal oxide, nitrogen(IV) oxide and oxygen gas
Q	Metal, nitrogen(IV) oxide and oxygen
R	Metal nitrite and oxygen gas

(a) Arrange the above metals in order of reactivity starting with the least reactive.
(b) Name <u>two</u> metals that could possibly be R.
(c) What is the general name of the metals in (b)?
(d) What is the group number of metals in (b) ?
(e) How many electrons are in their:

(i) Second energy level
(ii) Outermost energy level

(f) How are metals in (b) extracted from their ores?

7. Study the table below and answer the questions that follow. The letters X, Y and Z are not actual symbols of the elements.

Element	**Protons**	**Neutrons**	**Mass number**
X	12	-	24
Y	17	-	35
Z	-	12	23

(a) Complete the table.
(b) Write down the electronic arrangement of Z.
(c) Write down the formula of the compound formed when Z combines with Y.

8. A student from Muniu High School started with a metal oxide of iron and carried out a series of chemical reactions in order to prepare iron(II) sulphate. Study it and answer the questions that follow:

$$\begin{tikzcd}
FeO(s) \arrow[r, "\mathbf{A}"] & Fe(NO_3)_2(aq) \arrow[rr, "\mathbf{B}"] & & FeCO_3(s) \arrow[d, "\mathbf{C}"] \\
FeSO_4(s) & FeSO_4(aq) \arrow[l] & X(s) \arrow[l, "\mathbf{D}"'] & FeCl_2(aq) \arrow[l, "NaOH"']
\end{tikzcd}$$

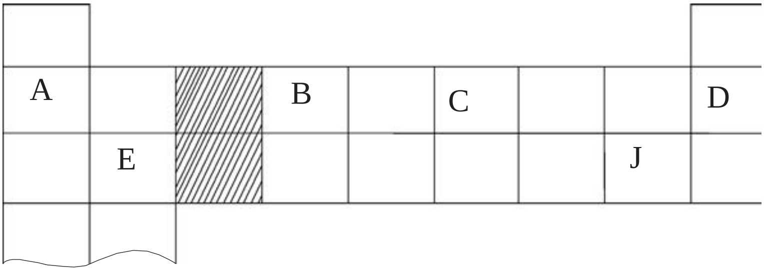
(a) Name the chemical reagents A, B, C, D.
(b) Name the solid X and write an equation of the reaction leading to the formation of iron(II) sulphate.
(c) It is possible to obtain the required salt by a single reaction. Write the word and chemical equations for this single reaction.

SELF CHECK TEST PAPER 3

1. (a) A student from Muchai Secondary School mixed copper(II) oxide with a black substance which is formed when burning magnesium ribbon is dipped into a gas jar of dry carbon(IV) oxide. She decided to use part of the mixture to prepare copper(II) sulphate crystals. How did she go about it?

   (b) Another student from the same class decided to use the other part of the mixture to prepare pure copper metal. How did he go about it?

2. Study the part of the periodic table below and answer the questions that follow (The letters do not represent actual chemical symbols of the elements).



   (a) Choose the element which is
      (i) Alkali metal
      (ii) Halogen
      (iii) Most reactive metal
      (iv) Has an octet in the outermost energy level

   (b) Write the electronic arrangement of
      (i) Ion of E
      (ii) Ion of J
      (iii) Ion of A
      (iv) Atom of B

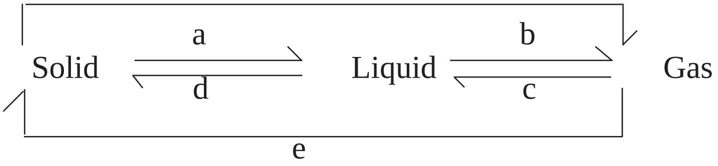
   (c) An element W is in Group V and period 3. Indicate the position of W on the grid and write its electron arrangement in the same grid.

   (d) Explain why the atomic radius of C is smaller than that of B.

   (e) (i) Explain why the ionic radius of A is smaller that its atomic radius.
      (ii) Explain why the ionic radius of J is greater than its atomic radius.

(f) Which element forms an ion of oxidation number of:
   (i) +2
   (ii) -1
   (iii) +1

3. (a) What is the use of the thermometer in the study of the Kinetic Theory of matter?
   (b) Why is heat used in the study of Kinetic Theory of matter?
   (c) Study the flow-chart below



Name the processes
a
b
c
d.

4. Fill the table below. An example is given.

Salt	**Ions present**	**Valency of ion**
$FeCl_3$	Fe^{3+}	3
	Cl^-	1
$AgNO_3$		
$(NH_4)_2SO_4$		
H_3PO_4		
$NaHSO_4$		

5. Give <u>four</u> reasons why the reactions which take place when hydrated

copper(II) sulphate is heated and when water is added to the residue is described as a reversible chemical reaction. Use a chemical equation to explain.

6. When sodium metal is left exposed to the atmosphere it may undergo the following processes.

$$\begin{tikzcd}
Na \arrow[r, "\mathbf{a}"] & Na_2O \arrow[r, "\mathbf{b}"] & NaOH \arrow[r, "\mathbf{c}"] & Na_2CO_3 \arrow[dl, "\mathbf{d}"] \\
 & Na_2CO_3.H_2O & Na_2CO_3.10H_2O \arrow[l, "\mathbf{e}"'] &
\end{tikzcd}$$

(a) What substances are absorbed at the stages a, b and c?
(b) Name process
(c) How can you stop these processes?

7. Explain the following:
(a) Carbon(IV) oxide is one of the causes of greenhouse effect.
(b) Carbon(IV) oxide is suitable for extinguishing fire.

8. Explain the following:
(a) Sodium chloride becomes damp when left in air.
(b) Graphite is a non-metal yet it conducts heat and electricity.
(c) Many cooking pans are made of aluminium although it is a reactive metal.
(d) Diamond and methane have the same type of bonding, yet methane is a gas at room temperature and pressure, while diamond is a very hard solid.
(e) The reaction between dilute sulphuric acid and barium carbonate is not suitable for preparing carbon(IV) oxide.

SELF CHECK TEST PAPER 4

1. Explain why aluminium chloride exist as a dimer and not a single molecule.

2. The table below gives information of four elements represented by letters P, Q, R and S. Study it and answer the questions that follow. The letters do not represent the actual symbols of the elements.

Element	**Electron arrangement**	**Atomic radius (nm)**	**Ionic radius (nm)**
P	2.8.2	0.136	0.065
Q	2.8.7	0.099	0.181
R	2.8.8.1	0.203	0.133
S	2.8.8.2	0.174	0.099

(a) Which <u>two</u> elements have similar chemical properties?
(b) What is the likely formula of the sulphide of R?
(c) Which element is a non-metal? Explain.
(d) Which elements belong to the same group? Explain.
(e) Explain the trend in the atomic radii of the elements named in (d).
(f) Between the elements named in (d) above which one is more reactive? Explain.
(g) Explain why the atomic radius of S is less than that of R.
(h) Explain why ionic radius of Q is larger than its atomic radius.

3. The electronic arrangement of ions of X^{3+} and Y^{2-} are 2.8 and 2.8.8 respectively.
(a) Write the electronic arrangement of their atoms.
(b) Write the formula of the compound formed by X and Y.
(c) Why is hydrogen atom sometimes placed under Group I and other times in Group VII in the periodic table?

4. In the separation of a mixture of ethanol and water what is the use of
(a) Fractionating column.
(b) Glass beads.

5. (a) Explain the role of propanone in the extraction of oil from groundnuts/cashew nuts/coconuts.

(b) What do we call this method of extraction?

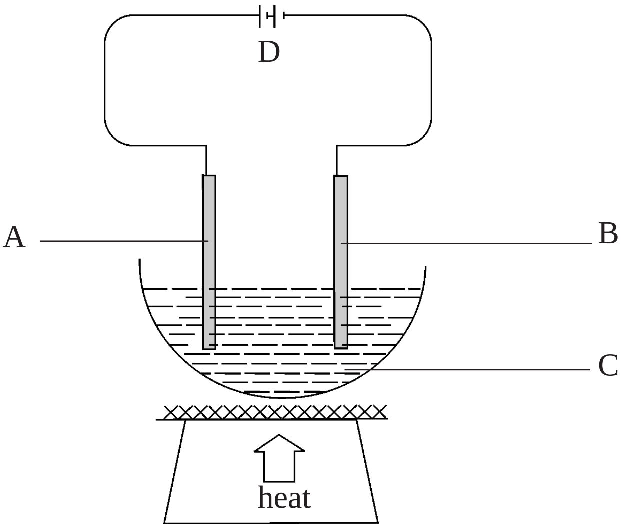
(c) What is the simple test of oils?

6. (a) Label the following diagram by name instead of letters.

(b) Assume the above diagram represents the set-up used during the electrolysis of molten lead(II) iodide, using graphite electrodes. What is observed at:

(i) A

(ii) B



(c) Why is it necessary to melt lead(II) iodide during electrolysis process?

7. (a) Name the type of bonds between the atoms/ions in the following substances.
Use the table below.

Substance	**Type of bond**
Aluminium chloride vapours	
Magnesium chloride	
Calcium	
Liquid water	
Ammonia gas	

(b) Write ionic equations for the following reactions:
    (i) Silver nitrate and sodium chloride solutions.
    (ii) Sodium carbonate solution and dilute hydrochloric acid.
    (iii) Lead nitrate solution and potassium sulphate solution.

8. (a) What is carbon cycle?
   (b) What role does respiration, decay and combustion of organic compounds play in the carbon cycle?

SELF CHECK TEST PAPER 5

1. An electric current was passed through two molten substances M and N in different containers. The observations made are in the table below:

Substance	**Observation**
Molten M	Conducts electric current and is not decomposed
Molten N	Conducts electric current and a pale green yellow gas formed around the anode

(a) Suggest the type of bond present in substances M and N.
(b) Suggest the name of the anion in the substance N.
(c) Suggest the name of the gas produced around the anode from substance N.

2 (a) List five physical properties of carbon (II) oxide.
(b) What is the test for this gas?
(c) Give one use of this gas.

3. Explain the following observations.
(a) Graphite is a good conductor of electricity and yet it is a non-metal.
(b) A solution of aluminium chloride in water reacts with sodium carbonate to produce carbon(IV) oxide.
(c) Aluminium is a better thermal and electrical conductor than magnesium and magnesium is better than sodium.

4. (a) State two observations made when:
   (i) Zinc carbonate is strongly heated, the gas produced passed through calcium hydroxide solution and the solid residue cooled.
   (ii) Copper nitrate is heated strongly and the residue cooled.
(b) Which salt in (a) above produces a gas which relights a glowing splint. Write a chemical reaction to show how this gas is formed.

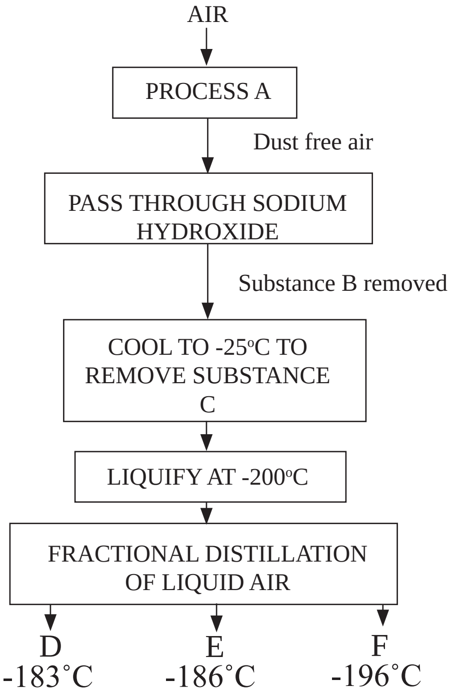
5. An ion Q^{2+} has an electronic arrangement of 2.8.
(a) What is the atomic number of the element.
(b) To which group and period in the periodic table does the element belong?

(c) What type of bond exists between the atoms of this element?
(d) How does the element conduct electricity when it is in molten form?

6. (a) Explain the role of propanone and filter paper in the separation of coloured matter in leaves/ink?
(b) What do we call this method of separation?
(c) Write chemical equations to show the action of heat on lead carbonate and zinc carbonate salts. Their residues undergo a physical change. Write the equations for changes physical changes and indicate colour changes.

7. (a) Name <u>two</u> substances recycled during the manufacture of sodium carbonate by the Solvay process.
(b) Which is the by-product of this process?
(c) Write the chemical equation to show how this by-product is produced?

8. Study the flowchart below and use it to answer the questions that follow.



(a) Name
(i) Process A
(ii) Substance B
(iii) Substance C
(iv) Substance D
(v) Substance E
(vi) Substance F

SELF CHECK TEST PAPER 6

1. (a) An element Y has 11 electrons in each of its atoms. With reasons, state the group and period to which the element belongs.
   (b) Electronic arrangement of an atom of element X is 2.8.6.
      (i) Write the electronic arrangement of an ion of X.
      (ii) Ionic radius of X is larger than its atomic radius. Explain.

2. Explain the following
   (a) Melting point increases from sodium to aluminium in the third period of the periodic table.
   (b) Atomic radius decreases across the period of the periodic table.

3. The following is an outline of one of the methods for the preparation of crystals of copper(II) sulphate ($CuSO_4.5H_2O$). An excess of copper carbonate is added to 50cm^3 of dilute sulphuric acid in a beaker and the mixture warmed until no further reaction takes place. The mixture is filtered and the filtrate evaporated until the volume is about 25cm^3. This solution is left to cool and the crystals are separated and dried.
   (a) Why was the mixture warmed?
   (b) Why is excess copper carbonate used?
   (c) How can you tell when the reaction has stopped?
   (d) Why is it necessary to filter the mixture after the reaction has stopped?
   (e) Why is the filtrate not evaporated to dryness.
   (f) What is the colour of the crystals obtained?
   (g) What observation is made when these crystals are heated in a beaker? What do we call the residue?

4. The structure of nitrogen and hydrogen are as shown below.

   $N \equiv N$ and H–H

   Hydrogen readily burns in oxygen but nitrogen does not. Explain.

5. (a) Water (H_2O) has molecular mass of 18 but is a liquid at room temperature while hydrogen sulphide (H_2S) has molecular mass of 34 but gaseous at room temperature. Why is this so?
   (b) A water molecule is said to be polar. Explain.

(c) State the chemical test of water.
(d) How can you confirm whether water is pure?

6. The table below gives some properties of three substances A, B and C.

Substance	**A**	**B**	**C**
Appearance	Brown solid	Yellow solid	Yellow solid
Melting point (°C)	1017	115	402
Solubility in water	Insoluble	Insoluble	Sparingly soluble
Electrical conductivity when solid	Conducts	Does not conduct	Does not conduct
Electrical conductivity when molted	Conducts	Does not decompose	Conduct

(a) Giving reasons for your answers which of the substances A, B and C has a:
  (i) Giant ionic structure/lattice.
  (ii) Giant molecular structure/lattice.
  (iii) Giant metallic structure/lattice.
(b) Why does C conduct electricity only in molten state but not in solid state?
(c) Why is A not decomposed when it conducts electricity?
(d) Explain in terms of bonding and structure why the melting point of B is lower than that of C.
(e) Which of these substances, if any, would dissolve in organic solvent like methyl-benzene? Explain.

8. The table below shows part of the periodic table. The letters used are not actual symbols of elements. Use the letters to answer the questions that follow:

	S	▨			T	U	V	W
	X	▨				Y		
							Z	

(a) Which <u>two</u> elements will form divalent cations?
(b) Which metallic element will react more vigorously with hot water?
(c) Write down the formula of the compound formed between S and
  (i) Y
  (ii) Z
(d) Name the element that does not form compounds.
(e) Which <u>two</u> elements have the greatest number of valency electrons?
(f) Explain briefly how the atoms of X are bonded.
(g) Write the chemical equations of the reaction which takes place when U is prepared in the laboratory and no heat is required. Use actual symbols.
(h) An isotope of Q has mass number of 35 and 18 electrons.
  (i) State its electronic arrangement.
  (ii) To which period and group does Q belong?

Glossary

A

Absorption:	A soaking up of one substance through the entire mass of another.
Acid salt:	A compound or salt forman acid in which not all the hydrogens have been replaced by a metal. eg. $NaHCO_3$, $NaHSO_4$ etc.
Active substance:	A substance which reacts vigorously with other substances, e.g., potassium which reacts vigorously with water.
Adsorption:	The acquisition of one substance, e.g. a gas by the surface of another, e.g. solid surface.
Allotropes:	Different forms of the same element in the same state, e.g. diamond and graphite is one allotrope of carbon.
Amphoteric:	Capable of acting either as an acid or as a base, e.g. aluminium oxide is an amphoteric oxide.
Amphoteric oxide:	A metal oxide which reacts with an acid or a base producing salts in both cases, e.g. aluminium oxide.
Anhydrous:	Containing no water of crystallisation.
Anion:	A negatively charged ion, e.g. chloride ion (Cl^-) ion.
Anode:	The positively charged electrode. It is joined to the positive terminal of the battery/cell. The anions move towards it during electrolysis process.
Atomic number:	The number of protons in an atom, e.g. the atomic number of carbon is 6 because it has six protons.

B

Baking soda:	Sodium hydrogencarbonate ($NaHCO_3$).
Basicity:	The number of hydrogen atoms in each molecule of an acid replaceable directly or indirectly by a metal or an ammonium radical, e.g. the basicity of sulphuric acid (H_2SO_4) is 2.
Binary compound:	A compound consisting of only two elements.
Bleaching:	The process by which colour is partially or wholly removed from a coloured substance.
Bleaching powder:	A chlorine containing substance used in bleaching textiles or sterilising water. It is made of slaked lime ($Ca(OH)_2$) and chlorine.
Boiling point:	The temperature at which equilibrium vapour of a liquid is equal to prevailing atmospheric pressure. The lower the atmospheric pressure, the lower the boiling temperature of a liquid at which the equilibrium vapour pressure are equal.

C

Carbon:	A non-metallic element of atomic number 6 and found in group 4 and period 2 of the periodic table.
Carbon cycle:	The process of removing and replacing of carbon(IV) oxide in the air thus keeping its percentage in air constant.
Carbon(II) oxide:	A colourless, odourless, tasteless, neutral and poisonous gas formed when carbon, hydrocarbons or organic matter burns in insufficient supply of air/oxygen. It burns with a blue flame.

Carbon(IV) oxide:	A heavy colourless and odourless gas which forms a white precipitate with calcium hydroxide solution.
Cathode:	A negatively charged electrode. It is joined to the negative terminal of the battery/cell. The cations move towards it during electrolysis process.
Cation:	A positively charged ion, e.g. sodium ion (Na^+).
Charged particle:	A particle with either negative or positive charge, e.g. electron (–), proton(+), chloride ion (Cl^-), sodium ion (Na^+).
Chemical bond:	A force which holds particles (atoms, ions or molecules) together.
Chemical equation:	A statement for the overall results of a chemical change and represented by symbols or formulae of the reactants and products, e.g. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$
Chemical formula:	A shorthand notation using chemical symbols and numerical subscripts to represent the composition of substances, e.g. Na_2CO_3 (sodium carbonate).
Chemical properties:	Those properties which pertain to the behaviour of a material in changes in which its identity is altered completely.
Combustion:	Any chemical reaction that occurs so rapidly that noticeable heat and light are produced, e.g. reaction between hydrogen and oxygen.
Conductor:	A substance which allows flow of an electric current, e.g.all metals and graphite. It can be in solid or molten form.
Covalent bond:	Bonding in which two atoms share a pair of electrons, e.g., the bond between hydrogen and chlorine atoms.

D

Dative bond:	Bonding in which two atoms share a pair of electrons. The pair of electrons are donated by one atom. The other atom only accepts the sharing. It is also called *co-ordinate bond*.
Dehydrating agent:	A substance which removes water vapour from another substance e.g., concentrated sulphuric acid.

E

Effervescence:	The process of giving off bubbes of a gas from a liquid in which it is dissolved or by reacting two reacting substances, e.g; reaction of calcium carbonate and dilute hydrochloric acid.
Electric current:	A flow of electrons from one point to another.
Electrode:	A metal or carbon rod (graphite) used to conduct electricity into (or out of) an electrolyte.
Electrolysis:	The decomposition of a compound in solution or molten form by electricity. D.C is used.
Electron:	A negatively charged particle revolving around the nucleus of an atom. It has negligible mass.
Electrolyte:	A substance which conducts electricity when in solution or molten form, e.g. aqueous or molten sodium chloride.
Electronegative element:	An element having a relatively strong attraction of electron of an atom within a molecule.
Electronegativity:	The tendency of an atom to attract the electron(s) from another atom forming a chemical bond between it and that atom.

Electroplating:	The process of coating the surface of metallic surface with another metal using electrolysis process, e.g. elecroplating iron with tin.
Electropositive element:	An element having a relatively weak attraction for its outermost energy level electron(s) or outermost energy level electron(s) of other elements.
Energy level:	A region outside the nucleus of an atom in which electrons may be considered to circulate.
Exothermic:	Pertaining to a chemical reaction which occurs with the evolution of heat energy.

G

Group:	A vertical column of elements in the periodic table, e.g. Group I having, Li, Na, K, Rb, Cs and Fr.

H

Halogen:	Name given to the family of elements having seven electrons in the outermost energy level of their atoms. i.e group VII
Homogeneous substance:	A substance which has similar properties throughout.
Hydrate:	A crystalline substance that contains water of crystallisation, e.g. $CuSO_4 \cdot 5H_2O$
Hydration:	The chemical attachment of water molecules to molecules of the solute, i.e., the process of adding water of crystallisation to a molecule.
Hydrogen bond:	A weak chemical bond between a hydrogen atom in one polar molecule and a negative atom in a second polar molecule e.g.

$$H^{\partial +} \text{———} Cl^{\partial -} \text{ --------- } H^{\partial +} \text{———} Cl^{\partial -}$$

Hydrolysis:	The reaction of a salt with water to form a solution which is acidic or basic, e.g. hydrolysis of aluminium chloride to form aluminium hydroxide and hydrogen chloride gas.
Hydroxonium ion:	A hydrated proton, i.e. H_3O^+ ion.

I

Inactive substance:	A substance which reacts passively with other substances e.g. nitrogen reacts with air only at very high temperature.
Inert substance:	A substance which does not react with other substances under the ordinary conditions of chemical reactions e.g. argon.
Insulator:	A non-conductor of electricity used to protect something from electricity e.g., plastic.
Insoluble:	Incapable of being dissolved.
Ion:	An atom or a group of atoms which possesses one or more positive or negative charges, e.g., Na^+, Cl^-, NO_3^-, SO_4^{2-}, Ca^{2+}, S^{2-} etc.
Ionic bonding:	Bonding in which one or more electrons are transferred from one atom to another.
Isotope:	One of two or more forms of atoms with the same atomic number but with different mass number which comes as a result of different number of neutrons in the nucleus e.g. $^{35}_{17}Cl$ and $^{37}_{17}Cl$
Isotopy:	The existence of elements in form of isotopes.

L

Lime water: Calcium hydroxide solution, it used for testing carbon(IV) dioxide.

Litmus: A dye extracted from lichen, which is used as an indicator. In the laboratory it is either in solution, crystals or soaked and dried into filter paper, usually known as litmus paper.

M

Mass: The quantity of matter which a body possesses.

Mass number: The number of protons plus the number of neutrons in an atom of an element.

N

Neutral particle: A particle with no charge, e.g. neutron, atom.

Neutron: A neutral particle found in the nucleus of an atom. It has a unit mass of one and no charge.

Non-conductor: A solid which does not allow an electric current to pass through e.g. wood, plastic, etc.

Non-electrolyte: A substance which does not conduct electricity in solution or molten form, e.g. wax, benzene.

Normal salt: A compound consisting of metal or ammonium radical and non-metallic radical derived from an acid. The non-metallic radical is incapable of further yielding more hydrogen. e.g. sodium carbonate ($Na_2 CO_3$).

Nucleus: The positively charged, dense and central part of an atom. It consists of protons($^+$), neutrons (no charge) and protons($^+$).

P

Period:	Horizontal row of elements in the periodic table, e.g. period 3 - containing, Na, Mg, AL, Si, Pi, S, Cl and Ar.
Periodic table:	A tabular arrangement of the chemical elements based on their atomic structure and number. Currently, the periodic table has 119 elements.
Physical properties:	These are properties which can be determined without causing a change in the identity of a substance e.g. density.
Polar molecule:	A molecule which has regions of positive and negative charge caused by unequal attraction of the shared electrons, e.g. $H^{\partial+}$ — $Cl^{\partial-}$
Precipitate:	A substance usually solid which separates from a solution as a result of some chemical or physical change e.g. lime water forms a white precipitate of calcium carbonate when reacted with carbon(IV) oxide. The white precipitate is calcium carbonate because it is insoluble in water.
Product:	An element or compound resulting from a chemical reaction e.g. the product of the reaction between hydrogen and oxygen is water.
Proton:	A positively charged particle found in the nucleus of an atom. It has one unit mass and one positive charge.

R

Radical:	A group of atoms which usually behaves like it were a single atom, e.g. SO^{2-}_4, NH^+_4.
Reactant:	An element or compound entering in a chemical reaction, e.g., hydrogen and oxygen are the reactants which form water.

S

Salt:	A compound consisting of metal or ammonium radical and non-metallic radical derived from an acid, e.g. sodium carbonate (Na_2CO_3), sodium hydrogencarbonate ($NaHCO_3$).
Soluble:	Capable of being dissolved.
Sublimation:	The process in which a solid changes to vapour or vapour to solid, without first forming liquid, e.g. iodine, ammonium chloride, solid carbon(IV) oxide.
Synthesis:	The process of combining of simple substances to make a more complex substance, e.g. synthesis of iron(III) chloride from iron and chlorine.

T

Triatomic molecule:	A molecule containing three atoms e.g. ozone (O_3), water (H_2O).
Trona:	A double salt of sodium carbonate and sodium hydrogencarbonate. It contains sodium chloride dissolved in its water of crystallisation.

V

Valence/ valency:	The number of electrons gained, lost or shared by an atom in chemical bonding with one or more atoms, e.g. The valency of iron(III) ion(Fe^{3+}) = 3 and chlorine ion (Cl^{-}) = 1. It is also the number of hydrogen atoms that an atom of an element or a radical can combine with, e.g., the valence of sulphate ion (SO_4^{2-}) = 2, nitrate ion (NO^{-}_{3}) = 1.
Valence or valency electrons:	Electrons in the unfilled outermost energy level.

Van der Waal's bond:	A comparatively weak bond (force) which holds molecules together in molecular crystals or liquids. The crystals have a low melting point and the liquids a low boiling point.

W

Water gas:	A fuel gas which is a mixture of carbon(II) oxide and hydrogen. It is made by allowing a blast of steam through a bed of red-hot coke.
Water of crystallisation:	The water present in the structure of the crystals in fixed amount, e.g. hydrated copper(II) sulphate ($CuSO_4.5H_2O$) has five molecules of water of crystallisation.
Word equation:	A brief statement which identifies the reactants in a chemical reaction and the products formed, e.g., Hydrogen + oxygen → water.

APPENDIX I
Periodic table

x Z y Name

x is the Relative Atomic Mass (RAM)
Z is the symbol of the element
y is the Atomic Numbers of the element

Periods / Groups	I	II											III	IV	V	VI	VII	O
1	1 H 1 Hydrogen																	4 He 2 Helium
2	7 Li 3 Lithium	9 Be 4 Beryllium											11 B 5 Boron	12 C 6 Carbon	14 N 7 Nitrogen	16 O 8 Oxygen	19 F 9 Fluorine	20 Ne 10 Neon
3	23 Na 11 Sodium	24 Mg 12 Magnesium	← Transition metals →										27 Al 13 Aluminium	28 Si 14 Silicon	31 P 15 Phosphorus	32 S 16 Sulphur	35.5 Cl 17 Chlorine	40 Ar 18 Argon
4	39 K 19 Potassium	40 Ca 20 Calcium	54 Sc 21 Scandium	48 Ti 22 Titanium	51 V 23 Vanadium	52 Cr 24 Chromium	55 Mn 25 Manganese	56 Fe 26 Iron	Co Cobalt	59 Ni 28 Nickel	64 Cu 29 Copper	65 Zn 30 Zinc	70 Ga 31 Gallium	73 Ge 32 Germenium	75 As 33 Arsenic	79 Se 34 Selenium	80 Br 35 Bromine	84 Kr 36 Krypton
5	86 Rb 37 Rubidium	88 Sr 38 Strontium	89 Y 39 Yttrium	91 Zr 40 Zirconium	93 Nb 41 Niobium	96 Mo 42 Molybdenum	99 Tc 43 Technetium	101 Ru 44 Ruthenium	103 Rh 45 Rhodium	106 Pd 46 palladium	108 Ag 47 Silver	112 Cd 48 Cadmium	115 In 49 Indium	119 Sn 50 Tin	122 Sb 51 Antimony	128 Te 52 Tellurium	127 I 53 Iodine	131 Xe 54 Xenon
6	133 Cs 55 Caesium	137 Ba 56 Barium	175 Lu 71 lutetium	178 Hf 72 Hafnium	181 Ta 73 Tantalum	184 W 74 Tungsten	186 Re 75 Rhenium	190 Os 76 Osmium	192 Ir 77 Iridium	195 Pt 78 Platinum	197 Au 79 Gold	201 Hg 80 Mercury	204 Ti 81 Thallium	207 Pb 82 Lead	209 Bi 83 Bismuth	210 Po 84 Polonium	210 At 85 Astatine	222 Rn 86 Radon
7	223 Fr 87 Francium	226 Ra 88 Radium	257 Lr 103 Lawrencium	261 Rf 104 Rutherfordium	262 Db 105 Dubnium	266 Sg 106 Seaborgium	264 Bh 107 Bohrium	269 Hs 108 Hassium	268 Mt 109 Meitnerium	271 Uun 110 Ununniliam	272 Uuu 111 Unununium	285 Uub 112 Ununbium	Uut 113	289 Uuq 114 unnuqnabium	Uup 115	292 Uuh 116 ununhexium	Uns 117	Uuo 118 ununoctium

Lanthanoid elements														
Lanthanoid elements	257 La 57 Lanthanum	140 Ce 58 Cerium	141 Pr 59 Praseodymium	144 Nd 60 Neodymium	147 Pm 61 Promethium	150 Sm 62 Samarium	152 Eu 63 Europium	157 Gd 64 Gadolinium	159 Tb 65 Terbium	163 Dy 66 Dysprosium	165 Ho 67 Holmium	167 Er 68 Erbium	169 Tm 69 Thulium	173 Yb 70 Ytterbium
Actinoid elements	227 Ac 89 Actium	232 Th 91 Thorium	231 Pa 91 Protactinum	238 U 92 Uranium	237 Np 93 Neptunium	242 Pu 94 Plutonium	243 Am 95 Americium	247 Cm 96 Curium	249 Bk 97 Berkelium	251 Cf 98 Carlifonium	254 Es 99 Einsteinium	253 Fm 100 Fermium	256 Md 101 Mendelevium	254 No 102 Nobelium

APPENDIX II

Safety symbols

Warning signs

Electrical shock

Oxidising

Flammable

Explosive

Corrosive

Irritant

Dangerous

Radioactive

Danger

Laser radiation