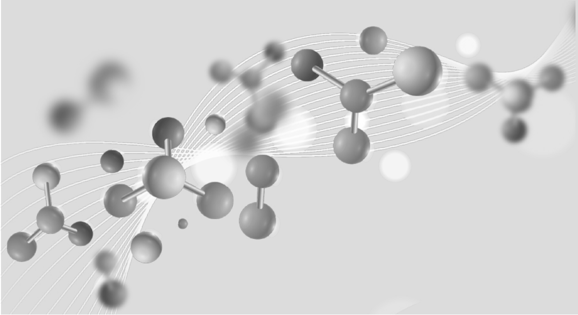
CHEMISTRY

Grade 10

Student Textbook



Content

CHEMISTRY10

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1
TOPIC

Introduction to Chemistry

Learning Objectives

Upon completion of this topic, learners will:

- Demonstrate knowledge about the origins and various stages in the development of chemistry
- Express appreciation for the scientific method
- Explain the word Chemistry and other related terminologies
- Distinguish the systems of units of measurement
- Solve simple conversion problems
- Discuss the origin of symbols of element
- Apply the symbols to write the formula and the names of compounds
- Apply the laboratory safety rules and
- Identify apparatus in the Lab.

Introduction

Science can be viewed as a continuing human effort to systematise knowledge for describing and understanding nature. In our daily life, we come across different substances present in nature and changes in them. Formation of curd from milk and rusting of iron are some of the examples of changes which we come across many times. For the sake of convenience,science is sub-divided into various disciplines such as chemistry,physics, biology, geology, etc.

Definition of Chemistry

The word "chemistry" comes from the Arabic word "*al-kimia*" meaning "the art of transformation". *Chemistry is the branch of science that studies the preparation, properties, structure and reactions of material substances.*

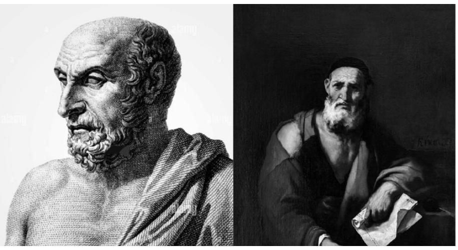
Terminologies Related to Chemistry

Some terminologies related to Chemistry are:

- **Acid:** Any substance that will release hydrogen ions when mixed with water.
- **Atom:** The smallest piece of an element that maintains the identity of that element.
- **Base:** Any substance that accepts proton and releases hydroxide ions (OH^-).
- **Compound:** A combination of more than one element.
- **Element:** A substance that cannot be broken down into simpler chemical substances by ordinary chemical means.
- **Matter:** Anything that has mass and occupies space.
- **Mixture:** A physical combination of more than one substance.
- **Molecule:** The smallest part of a substance that has the physical and chemical properties of that substance.
- **Neutron:** A subatomic particle with no charge.
- **Nucleus:** The centre of an atom that contains protons and neutrons.
- **Proton:** A subatomic particle with a positive charge.
- **Salt:** Any ionic compound that is formed from a reaction between an acid and a base.
- **Solution:** A homogeneous mixture of two or more components in which the particle size is smaller than 1 nm.
- **Solvent:** A chemical substance that dissolves another chemical substance to form a solution.
- **Valency:** The number of electrons that can be donated, accepted or shared by an atom of an element during a chemical reaction.

1.1. DEVELOPMENT OF CHEMISTRY

It was not until the era of the ancient Greeks that we have any record of how people tried to explain the chemical changes they observed and used. At that time, natural objects were thought to consist of only four basic elements: earth, air, fire, and water. In the fourth century BC, two Greek philosophers, Democritus and Leucippus, suggested that matter was not infinitely divisible into smaller



Democritus & Leucippus

particles but instead consisted of fundamental, indivisible particles called **atoms**. Unfortunately, they did not have the technology to test their hypothesis.

Over the next two millennia, alchemists achieved many advances in chemistry. Their major goal was to convert certain elements into others by a process they called **transmutation**.

In particular, alchemists wanted to find a way to transform cheaper metals into gold. Although most alchemists did not approach chemistry systematically and many appear to have been outright frauds, alchemists in China, the Arab kingdoms, and medieval Europe made major contributions, including the discovery of elements such as quicksilver (mercury) and the preparation of several strong acids.

The 16th and 17th centuries saw the beginnings of the modern chemistry. During this period, great advances were made in metallurgy and the first systematic quantitative experiments were carried out. In 1661, Robert Boyle (1627–91) published "*The Sceptical Chymist*," which described the relationship between the pressure and the volume of air. In the 18th century, Joseph Priestley (1733–1804) discovered oxygen gas and found that many carbon-containing materials burn vigorously in an oxygen atmosphere, a process called **combustion**.

Robert Boyle (1627-1691)

In the late 18th century, the French scientist Antoine Lavoisier (1743–94) showed that combustion is the reaction of a carbon-containing substance with oxygen to form carbon dioxide and water. His most important contribution was the '*law of conservation of mass.*' J. L. Proust formally stated the '*law of definite proportions*' in 1797.

In 1803, John Dalton, an English school teacher, expanded the findings if Lavoisier and Proust, and wrote his '*atomic theory of matter.*'

Joseph Gay Lussac (1778–1850) states that gases combine in specific ratio to form a compound.

Amedeo Avogadro (1776–1856) introduced the hypothesis that equal volumes of gases at the same pressure and temperature contain the same number of molecules. It is popularly known as **Avogadro's hypothesis**. It provided the first link between the macroscopic properties of a substance (volume) and the number of atoms or molecules present.

1.1.1. Scientific Method and its Steps

The scientific method is a logical and rational order of steps by which scientists come to conclusions about the world around them. **The scientific method helps to organize thoughts and procedures** so that scientists can be confident in the answers they find. The **steps of the scientific method** are:

1. **Observation:** This step could also be called "research". It is the first stage in understanding the problem you have chosen. For this stage of the scientific method, it is important to use as many sources as you can find. The more information you have; the better the design of your experiment is going to be.

2. **Hypothesis:** The next stage of the scientific method is known as the "hypothesis". This word basically means "a possible solution to a problem, based on knowledge and research". The hypothesis is a simple statement that defines what you think the outcome of your experiment will be.

3. **Prediction:** The hypothesis is your general statement of how you think the scientific phenomenon in question works. Your prediction lets you get specific—how will you demonstrate that your hypothesis is true? The experiment that you will design is done to test the prediction. An **important thing to remember** during this stage of the scientific method is that once you develop a hypothesis and a prediction, you shouldn't change it, even if the results of your experiment show that you were wrong. An incorrect prediction doesn't mean that you "failed". It just means that the experiment brought some new facts to light that may be you hadn't thought about before.

4. **Experimentation:** This is the part of the scientific method that tests your hypothesis. An experiment is a tool that you design to find out if your ideas about your topic are right or wrong. The experiment is the most important part of the scientific method.

5. **Conclusion:** The final step in the scientific method is the conclusion. This is a summary of the experiment's results, and how those results match up to your hypothesis.

Group Activity

In the guidance of your teacher, visit few areas. Investigate and apply scientific method to find possible solution to common Liberian problems.

1.1.2. Contributors of Chemistry

There were many persons who contributed to the development of chemistry. The following table lists some major contributors of chemistry.

Table 1.1. Some Major Contributors of Chemistry

S. No.	Contributors of Chemistry	Their Contribution
1.	Democritus and Leucippus	They suggested that matter was not infinitely divisible into smaller particles but instead consisted of fundamental, indivisible particles called atoms.
2.	Robert Boyle	He formulated the Boyles's law which describes the relationship between pressure and volume of air.
3.	Antoine Lavoisier	His most important contribution was the '*law of conservation of mass.*'
4.	J. L. Proust	He stated the '*law of definite proportions.*'
5.	John Dalton	Wrote atomic theory of matter.
6.	Joseph Gay Lussac	He formulated the Gay-Lussac's Law.
7.	Amedeo Avogadro	He formulated Avogadro's hypothesis.

1.1.3. Branches of Chemistry

Chemistry plays a central role in science and is often intertwined with other branches of science. It has been further divided into different branches depending upon specialised fields of study. The various branches of chemistry are:

(i) **Inorganic Chemistry:** This branch of chemistry deals with the study of compounds of all other elements except carbon.

(ii) **Organic Chemistry:** This branch of chemistry deals with the study of carbon compounds especially hydrocarbons and their derivatives.

(iii) **Physical Chemistry:** This branch of chemistry deals with the explanation of fundamental principles governing various chemical phenomena.

(iv) **Industrial Chemistry:** This branch deals with the chemistry of processes involved in various industrial processes.

In addition to above branches there are other branches of chemistry developed in recent years. These include, *analytical chemistry,* *bio-chemistry, nuclear chemistry, pharmaceutical chemistry, geochemistry, agricultural chemistry, solid state chemistry, etc.*

1.2. UNITS OF MEASUREMENT

Measurement of any physical quantity involves comparison with a certain basic, arbitrarily chosen, internationally accepted reference standard called **unit**. The result of a measurement of a physical quantity is expressed by a number (or numerical measure) accompanied by a unit. Although the number of physical quantities appears to be very large, we need only a limited number of units for expressing all the physical quantities, since they are interrelated with one another. The units for the fundamental or base quantities are called **fundamental** or **base units**. The units of all other physical quantities can be expressed as combinations of the base units. Such units obtained for the derived quantities are called **derived units**.

1.2.1. System of Units

A complete set of both the base and derived units, is known as the ***system of units***.

Earlier, two different systems of measurement, i.e., the **English System** and the **Metric System** were being used in different parts of the world. The metric system,which originated in France in late 18th century, was more convenient as it was based on the decimal system. Late, need of a common standard system was felt by the scientific community. Such a system was established in 1960 and is known as International System of Unit (SI).

1.2.1.1. International System of Units (SI)

The International System of Units (in French *Le Systeme International d'Unités*—abbreviated as SI) was established by the 11th General Conference on Weights and Measures (CGPM from *Conference Generale des Poids at Measures*). The CGPM is an inter governmental treaty organization created by a diplomatic treaty known as Meter Convention which was signed in Paris in 1875. The metric system was found to be more convenient as it was based on the decimal system. The fundamental units of **metric system** are, *gram* **for mass**, *the metre* **for length** and *the litre* for the **volume** of fluids.

The SI system has seven base units which are listed in Table 1.2.

Table 1.2. The Seven Basic SI Units

Physical Quantity	Abbreviation	Name of Unit	Symbol
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Luminous intensity	I_v	candela	cd
Amount of substance	n	mole	mol

Decimal fractions or multiples of units are expressed by putting certain prefixes before the names of the units. The various prefixes used for this purpose are listed in the Table 1.3.

Table 1.3. Prefixes for Expressing the Decimal Fractions in the SI System

Multiple	Prefix	Symbol
10^{12}	tera	T
10^{9}	giga	G
10^{6}	mega	M
10^{3}	kilo	k
10^{2}	hecto	h
10	deka	da
10^{-1}	deci	d
10^{-2}	centi	c
10^{-3}	milli	m
10^{-6}	micro	µ
10^{-9}	nano	n
10^{-12}	pico	p

The idea of using prefixes is illustrated by the following example:

A hundredth of a metre, corresponding to 10^{-2}, is a *centimetre* for which the symbol *cm* is used. Similarly, for 10^{3} metre we can use the term *kilometre*, symbolized as *km*.

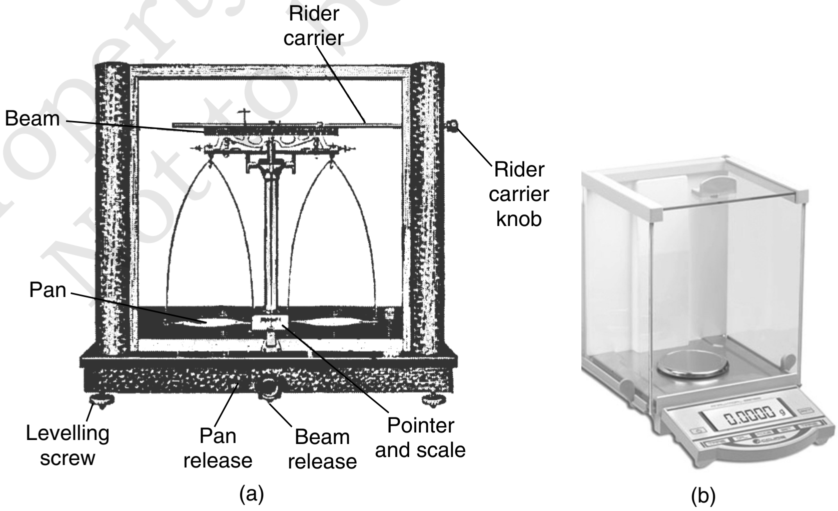
In addition to the seven basic units, there are many derived units. These units are obtained by combination of basic units. For example, *square metre* (m^2) is the unit for area and the *cubic metre* (m^3) is the unit for volume. Some of the commonly used SI derived units are given in Table 1.4. These derived units are obtained by defining the physical quantities for which they are used.

Table 1.4. Some Common SI Derived Units

Physical Quantity	**Name of Unit**	**Symbol for Unit**	**Definition of Quantity Unit**	**Expression in Terms of SI Basic Units**
Area	—	—	Length squared	m^2
Volume	—	—	Length cubed	m^3
Density	—	—	Mass per unit volume	kg/m^3 or $kg\ m^{-3}$
Frequency	hertz	Hz	Cycles per second	s^{-1}
Electric charge	coulomb	C	Ampere times second	A s
Electric potential difference	volt	V	Energy per unit charge	$JA^{-1}\ s^{-1}$ or $kg\ m^2\ s^{-3}\ A$

1.2.2. Measurement of Mass, Length, Time, Temperature and Volume

Mass of a substance is the amount of matter present in it. The mass of a substance is constant.



(a) (b)

Fig. 1.1. Laboratory Balances

Standard laboratory balances can determine the mass of an object to a minimum of three significant figures.

The mass of a substance can be determined very accurately in the laboratory by using analytical balances (Fig. 1.1).

You are already familiar with some direct methods for the measurement of length.

1.2.2.1. Measurement of Length

You are already familiar with some direct methods for the measurement of length. For example, a metre scale is used for lengths from 10^{-3} m to 10^{2} m. A Vernier callipers is used for lengths to an accuracy of 10^{-4} m. A screw gauge and a spherometer can be used to measure lengths as less as to 10^{-5} m.

1.2.2.2. Measurement of Time

To measure any time interval we need a clock. We now use an **atomic standard of time**, which is based on the periodic vibrations produced in a caesium atom. This is the basis of the **caesium clock**, sometimes called **atomic clock**. Such standards are available in many laboratories. In the caesium atomic clock, the second is taken as the time needed for 9,192,631,770 vibrations of the radiation corresponding to the transition between the two hyperfine levels of the ground state of caesium-133 atom. The vibrations of the caesium atom regulate the rate of this caesium atomic clock just as the vibrations of a balance wheel regulate an ordinary wristwatch or the vibrations of a small quartz crystal regulate a quartz wristwatch.

The caesium atomic clocks are very accurate. In principle they provide portable standard. The national standard of time interval 'second' as well as the frequency is maintained through four caesium atomic clocks.

1.2.2.3. Measurement of Temperature

There are three common scales to measure temperature—°C (degree celsius), °F (degree fahrenheit) and K (kelvin). Here, K is the SI unit. The thermometers based on these scales are shown in Fig. 1.2. Generally, the thermometer with celsius scale are calibrated from 0° to 100° where these two temperatures are the freezing point and the boiling point of water respectively. The fahrenheit scale is represented between 32° to 212°.

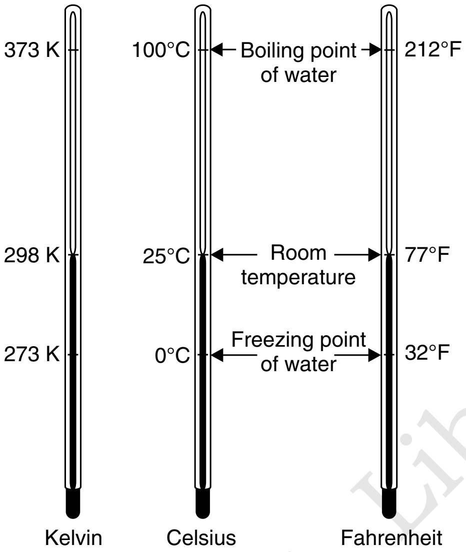


Fig. 1.2. Thermometers based on different temperature scales

The temperatures on two scales are related to each other by the following relationship:

$$°F = \frac{9}{5}(°C) + 32$$

The kelvin scale is related to celsius scale as follows :

$$K = °C + 273.15$$

It is worthwhile to note that the temperature below 0°C (*i.e.*, negative values) are possible in Celsius scale but in Kelvin scale, negative temperature is not possible.

1.2.2.4. Measurement of Volume

Volume has the units of $(\text{length})^3$. So in SI system, volume has units of m^3. But since smaller volume are used in laboratories, hence the units of cm^3 or dm^3 are used. A common unit, litre (L) which is not an SI unit, is used for the measurement of volume of liquids.

$$1\ L = 1000\ mL \text{ or } 1000\ cm^3$$
$$= 1\ dm^3 = 10^{-3}\ m^3$$
$$1\ cm^3 = 10^{-3}\ dm^3 = 10^{-6}\ m^3.$$

In the laboratory, volume of liquids or solutions can be measured by graduated cylinder, burette, pipette, etc. A volumetric flask is used to prepare a known volume of a solution. These measuring devices are shown in Fig. 1.3.

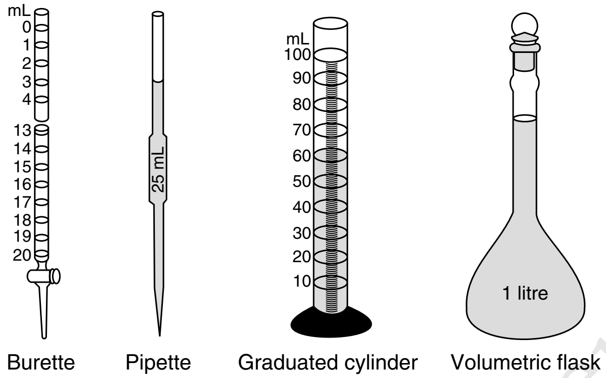


Fig. 1.3. Some volume measuring devices

1.2.3. Conversion of Units

The rule of conversion is "**multiplication is done for converting higher units to smaller units**, whereas **division is carried out for converting smaller units to higher units.**"

The units are expressed using scientific notation and converted into numerical values as per the quantities.

We can convert meter to centimeter by using the conversion factor 1 m = 100 cm. For example,

$$50 \text{ m} = 50 \times 1 \text{ m} = 50 \times 100 \text{ cm} = 5000 \text{ cm}$$

Similarly, we can convert centimeter to meter by using the conversion factor $1 \text{ cm} = \frac{1}{100}$ m. For example,

$$50 \text{ cm} = 50 \times \frac{1}{100} \text{ m} = \frac{50}{100} \text{ m} = 0.5 \text{ m}$$

1.2.3.1. Conversion of Unit of Mass

Below is the conversion table for unit of mass.

Conversion Table for Unit of Mass

1 kg = 10 hg = 100 dag = 1000 g
1 g = 10 dg = 100 cg = 1000 mg
1 dg = 10 cg = 100 mg
1 cg = 10 mg

Example 1.1: *Convert the following:*

(i) *5 g to cg* (ii) *500 g to kg*

Solution:

(i)
$$1\ \text{g} = 100\ \text{cg}$$
$$5\ \text{g} = 5 \times 100\ \text{cg} = 500\ \text{cg}$$

(ii)
$$1000\ \text{g} = 1\ \text{kg}$$
$$1\ \text{g} = \frac{1}{1000}\ \text{kg}$$
$$500\ \text{g} = \left(\frac{1}{1000}\right) \times 500\ \text{kg} = 0.500\ \text{kg}$$

1.2.3.2. Conversion of Unit of Length

Below is the conversion table for unit of length.

Conversion Table for Unit of Length

1 km = 10 hm = 100 dam = 1000 m
1 m = 10 dm = 100 cm = 1000 mm
1 dm = 10 cm = 100 mm
1 cm = 10 mm

Example 1.2: *Convert the following:*

(i) *5 m to cm* (ii) *500 m to km*

Solution:

(i)
$$1\ \text{m} = 100\ \text{cm}$$
$$5\ \text{m} = 5 \times 100\ \text{cm} = 500\ \text{cm}$$

(ii)
$$1000\ \text{m} = 1\ \text{km}$$
$$1\ \text{m} = \frac{1}{1000}\ \text{km}$$
$$500\ \text{m} = \left(\frac{1}{1000}\right) \times 500\ \text{km} = 0.500\ \text{km}$$

1.2.3.3. Conversion of Unit of Time

Below is the conversion table for unit of time.

Conversion Table for Unit of Time

1 minute = 60 seconds
1 hour = 60 minutes= 3600 seconds
1 day = 24 hours
1 week = 7 days
1 year = 365 days

Example 1.3: *Convert the following:*

(i) *30 minutes to seconds*
(ii) *1200 seconds to hour*

Solution:

(i)
$$1 \text{ minute} = 60 \text{ seconds}$$
$$30 \text{ minutes} = 60 \times 30 = 1800 \text{ seconds}$$

(ii)
$$3600 \text{ seconds} = 1 \text{ hour}$$
$$1 \text{ second} = \frac{1}{3600} \text{ hour}$$
$$1800 \text{ seconds} = \left(\frac{1}{3600}\right) \times 1800 = \frac{1}{2} \text{ hour}$$

1.2.3.4. Conversion of Unit of Temperature

The three main conversions of temperature take place between:

- Celsius and Kelvin
- Fahrenheit and Kelvin
- Celsius and Fahrenheit

Here is the list of the conversion formulas for the different units of temperature.

Conversion Table for Unit of Temperature

Conversion	Formulas
Celsius to Kelvin	$K = °C + 273.15$
Kelvin to Celsius	$°C = K - 273.15$
Fahrenheit to Celsius	$°C = (°F - 32) \times \left(\frac{5}{9}\right)$
Celsius to Fahrenheit	$°F = °C \times \left(\frac{9}{5}\right) + 32$

Example 1.4: *Convert the following:*

(i) *20° Celsius to Fahrenheit* (ii) *212 Fahrenheit to Celsius*
(iii) *30° Celsius in Kelvin* (iv) *400 Kelvin to Celsius*

Solution:

(i) Placing the Celsius value in the formula:

$$°F = \left(°C \times \frac{9}{5}\right) + 32 = \left(20 \times \frac{9}{5}\right) + 32$$
$$= 36 + 32 = 68\ °F$$

Therefore, 20° Celsius is equal to 68° Fahrenheit.

(ii) Placing the Fahrenheit value in the formula:

$$
\begin{aligned}
°C &= (°F - 32) \times \frac{5}{9} \\
&= (212 - 32) \times \frac{5}{9} = 180 \times \frac{5}{9} \\
&= 100\ °C
\end{aligned}
$$

Therefore, 212° Fahrenheit is equal to 68° Celsius.

(iii) Placing the Celsius value in the formula:

$$K = 30 + 273.15 = 303.15$$

Therefore, 30° Celsius is equal to 303.15 Kelvin.

(iv) Placing the Kelvin value in the formula:

$$°C = K - 273.15 = 400 - 273.15 = 126.85\ °C$$

1.2.3.5. Conversion of Unit of Volume

We normally use liter to represent as the standard unit, and the other standard units are in the units in capacity and volume conversion chart.

Conversion Table for Unit of Volume

1 milliliter	=	0.001 liter
1 centiliter	=	0.01 liter
1 deciliter	=	0.1 liter
1 decaliter	=	10 liters
1 hectoliter	=	100 liters
1 kiloliter	=	1000 liters

Example 1.5: *Convert the following:*

(i) *65 milliliter into liter* (ii) *8 hectoliter into liter.*

Solution:

(i)

$$
\begin{aligned}
65 \text{ milliliter} &= 65 \times 0.001 \text{ liter} \\
&= 0.065 \text{ liter}
\end{aligned}
$$

(ii)

$$8 \text{ hectoliter} = 8 \times 100 = 800 \text{ liters}$$

1.2.4. Scientific Notations and Significant Figures

1.2.4.1. Scientific Notations

Chemistry is the study of atoms and molecules which have extremely low masses and are present in extremely large numbers. For example, 2 g of hydrogen gas has 602,200,000,000,000,000,000,000 molecules of hydrogen gas. The mass of each H atom is also found to be as small as 0.00000000000000000000000166 g. Similarly, many other constants such as Planck's constant, speed of light, charges on particles, etc. involve numbers of very small magnitude.

Looking at the above two figures and many others like those, it may look funny for a moment to write or count numbers involving so many zeroes but it offers a real challenge to do simple mathematical operations of *addition*, *subtraction*, *multiplication* or *division* with such numbers. Thus, it is always difficult to handle such small numbers.

This problem is solved by using scientific notation for such numbers. The scientific notation is also called **standard form** or **exponential notation**. In this notation, any number can be represented in the form $N \times 10^n$ where n is an exponent having positive or negative values and N is a number that can vary between 1 to 10.

While writing the scientific notation, the decimal in the given number is placed after the first non-zero digit. For doing this, the decimal has to be shifted towards right or left of the original decimal place. The number of places by which the decimal is shifted gives the value of exponent 'n'. Further, n is positive if the decimal is shifted towards left and it is negative if the decimal is shifted towards right. For example,

(i) **Number** : 136.325
Scientific notation : 1.36325×10^2

Here, the decimal has been shifted by two places towards left. Hence, $n = 2$.

(ii) **Number** : 0.00025
Scientific notation : 2.5×10^{-4}

Here, the decimal has been shifted by four places towards right. Hence, $n = -4$.

Example 1.6: *Express the following in the scientific notation.*

(i) *1,86,000* (ii) *0.00683* (iii) *7.0042.*

Solution:

(i) 1.86×10^5 (ii) 6.83×10^{-3} (iii) 7.0042.

1.2.4.2. Significant Figures

Experimental measurements have some uncertainty associated with them. However, one would always like the results to be ***precise*** and ***accurate***. **Precision** means how closely the individual measurements agree with one another. **Accuracy**, on the other hand, means the closeness to the experimental measurements and the true or accepted value. These aspects further depend on the accuracy of **measuring device** and the **skill of the operator**.

For example, the magnitude of the uncertainty associated with measurement of length of an object will depend upon accuracy of scale used. If a centimetre scale is used to measure the length of a page, we might find that it is 14.6 centimetre. In such a measurement we would be sure that the number of *tens* in the numerical expression is 1 and the number of *ones* is 4. However, we cannot be sure of the number corresponding to tenths of a centimetre. This last digit is, therefore, called doubtful digit. One of the method of expressing uncertainty is to use the notation ± 1 alongwith the doubtful digit. For example, in the above measurement of length, the doubtful digit is 6, which corresponds to the tenths of a centimetre. Hence, the uncertainty in length can be expressed by recording the above measurement as **14.6 ± 0.1 cm**.

A convenient method of expressing the uncertainty in measurement is to express it in terms of *significant figures* instead of using the notation ± 1. In this method, it is assumed that all the digits are known with certainty except the last digit which is uncertain to the extent of ± 1 in that decimal place. Thus, a measured quantity is expressed in terms of such a number which includes all digits which are certain and a last digit which is uncertain. The total number of digits in the number is called the number of **significant figures**.

The number of significant figures in a measurement is the number of figures that are known with certainty plus one that is uncertain, beginning with the first non-zero digit.

In order to determine the significant figures in a measured quantity the following rules should be applied.

1. ***All non-zero digits are significant.***

   For example, 165 cm has **three** significant figures; 0.165 has also **three** significant figures. Similarly, 2006 has four significant figures, 9.05 has three significant figures, etc.

2. ***Zeros to the left of the first non-zero digit in the number are not significant.***

   For example, 0.005 g has only **one** significant figure, 0.026 g has **two** significant figures.

3. ***Zeros between non-zero digits are significant.***

   For example, 2.05 g has **three** significant figures.

4. ***Zeros to the right of the decimal point are significant.***

   For example, 5.00 g, 0.050 g, 0.5000 g have **three**, **two** and **four** significant figures respectively.

5. ***If a number ends in zeros that are not to the right of a decimal, the zeros may or may not be significant.***

   For example, 1500 g may have two, three or four significant figures.

   The ambiguity in the last point can be removed by expressing the number in **scientific notation**.

In scientific notation the number is written in the standard exponential form as $N \times 10^n$.

N = a number with a single non-zero digit to the left of the decimal point.

n = some integer.

For example, a mass of 1500 g can be expressed in scientific notation in the following forms depending upon whether it has two, three or four significant figures.

1.5×10^3 g (**Two** significant figures)

1.50×10^3 g (**Three** significant figures)

1.500×10^3 g (**Four** significant figures)

In these expressions all the zeros to the right of the decimal point are significant. The exponential notation is an excellent way of expressing significant figures in very large or very small measurements. For example, Avogadro's constant is expressed as 6.022×10^{23} mol^{-1} and Planck's constant as 6.62×10^{-34} Js.

Calculations with Significant Figures

During quantitative studies, the scientists have to do calculations with numbers used for various measured physical quantities. These numbers generally, have different number of significant digits depending upon the accuracy with which a particular measurement is made. While carrying out calculations with these numbers, the rule used is that *the accuracy of the final result is limited to the least accurate measurement.* In other words*, final result cannot be more accurate than the least accurate number involved in the calculation.*

Rounding Off

While carrying out calculations with experimentally measured quantities, the final result often contains figures that are not significant. When this happens, the final result is rounded off. In rounding off, the extra digits are dropped with or without minor changes in the figures retained. The rules employed for rounding off a number to the required number of significant digits are as follows:

1. *If the digit following the last digit to be retained is less than five, the last digit is left unchanged.* For example, suppose in the final result obtained as 46.32 only two figures are to be retained as significant figures. The last figure to be retained is 6 and the figure following it is 3. Since 3 is less than five, the 6 will be retained as such without change and the final result would be expressed as 46.
2. *If the digit following the last digit to be retained is more than five, the last digit retained is increased by one.* For example, suppose the result 52.87 is to be rounded off to three significant digits. The last digit to be retained is 8 and the digit following it is 7. Since 7 is more than five, therefore 8 would be increased by one to 9. The result in terms of significant figures would be expressed as 52.9.
3. *If the digit following the last digit to be retained is equal to five, the last digit is left unchanged if it is even and is increased by one if it is odd.*

Example 1.7: *Express the following numbers to four significant figures:*

(i) *5.607892* (ii) *32.392800*

Solution:

(i) As the fifth digit 8 is greater than 5, therefore the result will be expressed as 5.608.

(ii) It will be expressed as 32.39. The digit 2 is dropped and since it is less than 5, the figure is not rounded off to next number.

1.2.5. Accuracy and Precision

1.2.5.1. Accuracy

Accuracy is how close a measured value is to the **actual (true) value**. Accuracy describes how close an approximation is to a correct answer. Thus, accuracy "*is the measure of the difference between the experimental value or mean value of a set of measurements and the true value.*"

Accuracy = *Mean value – True value*

Smaller is the difference between the mean value and *the true value more is the accuracy.*

For *example*, suppose your math textbook tells you that the value of pi (π) is 3.14. You do a careful measurement by drawing a circle and measuring the circumference and diameter, and then you *divide the circumference by the diameter* to get a value for pi (π) of 3.16.

The accuracy of your answer is how much it differs from the accepted value.

In this case, the accuracy is 3.16 – 3.14 = **0.02.**

1.2.5.2. Precision

Precision is how close the measured values are **to each other**. Precision describes how many digits we use to approximate a particular value. In simple words, it is the difference between the measured value and the arithmetic mean value for a series of measurements, i.e.,

Precision = *Individual value – Arithmetic mean value*

Look at the following example:

The value of π (pi) is 22/7. Suppose the value is written as 3.1417 and 3.1392838. Looking at the value, the second number has higher precision, but it would appear that the first is more accurate. (Actual value is 3.142857143.)

1.3. CHEMICAL SYMBOLS, FORMULAE AND NAMING COMPOUNDS

1.3.1. Origin of Symbols

In order to represent the elements, instead of using full lengthy names, scientists use abbreviated names. These abbreviated names of the elements are known as symbols. Thus, ***symbol** of an element may be defined as the abbreviation used for the name of the element.*

The abbreviation used for the name of an element is more precisely known as chemical symbol of that element.

Dalton suggested symbols for the atoms of different elements as shown in Fig. 1.4. He was the first scientist to use the symbols for elements in a quantitative sense. When he used symbol for an element he meant a definite quantity of that element, that is, one atom of the element.

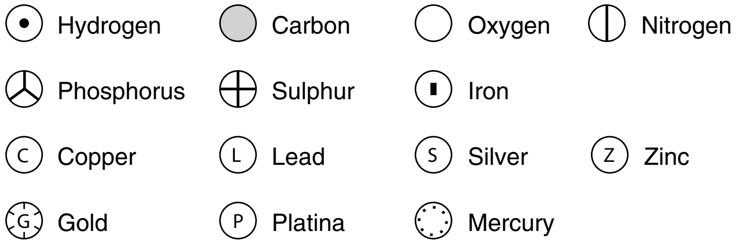


Fig. 1.4. Symbols for some elements as proposed by Dalton

Nowadays, IUPAC (International Union of Pure and Applied Chemistry) approves names of elements. The symbols of elements are generally either the first letter or the first two letters or the first and the third letters of the name of the element. Some symbols derived from the first letter of the names of the elements. For example, Hydrogen (H), Nitrogen (N), Carbon (C), etc.

Some symbols derived from the first two letters of the names of the elements. For example, Aluminium (Al), Barium (Ba), Lithium (Li), etc.

Some symbols derived from the first and the third letters of the names of the elements. For example, Arsenic (As), Magnesium (Mg), Chlorine (Cl), etc.

However, there are certain symbols which seem to have no relationship to their names. The symbols of these elements are, in fact, derived from their latin names. For example, Iron (*Ferrum*) –Fe, Gold (*Aurum*)– Au, etc.

It is important to note that the first letter of every chemical symbol is capital letter but, if the symbol consists of two letters, the second letter is not capital letter. Thus:

Symbol for aluminium is Al and not AL

Symbol for lead is Pb and not PB

1.3.2. Writing Chemical Formula

As already discussed, the elements are represented in the abbreviated form by their symbols. Similarly, a compound is represented in the abbreviated form by its chemical formula.

The expression of the composition of a substance by chemical symbols and numerical subscripts is called the ***chemical formula*** *of the substance.*

The formula of a simple substance is obtained by writing the symbol of the element and indicating the number of atoms in a molecule of the substance by a subscript. For example, a molecule of hydrogen contains two atoms. Hence, hydrogen is represented by the formula H_2. A molecule or ozone contains three atoms of oxygen. Hence, ozone is represented by the formula O_3.

In order to represent the chemical formula of a compound, one must know what are the elements present in the compound and what is the number of atoms of one element that combine with a definite number of atoms of the other element. *For example*, a molecule of water contains two hydrogen atoms and one oxygen atom. It is represented by the formula H_2O.

1.3.2.1. Writing the Formula of a Binary Molecular Compound

In binary molecular compounds, the atoms of the two elements are held by covalent bonds. A covalent bond, as you know, is formed by sharing of electrons between the two atoms. The number of electrons that an atom of the elements contribute for sharing is known as valency of the element. Knowing the valencies of the two elements involved in the formation of a binary molecular compound, the formula of the compound can be derived.

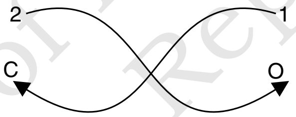
Step 1. *Write the constituent elements and their valencies as shown below:*

Valency	**4**	**2**
Element	**C**	**O**

Step 2. *Reduce the valency numerals to simplest whole numbers by dividing by some common factor, if any*

2	**1**
C	**O**

Step 3. *Criss-cross the reduced valency numerals and write them as subscripts at bottom right hand side of the symbols.*



The subscript 1 is not written. Thus, the formula of the compound is CO_2.

1.3.2.2. Writing the Formula of an Ionic Compound

The formula of an ionic compound represents the simplest whole number ratio of ions in it. The total positive charge on cations is equal to the total negative charge on anions. Knowing the formulae of the ions present in the compound, the formula of the compound can be written by the following steps:

Step 1. *Write the formulae of the ions or radicals of the compound side by side with cation on the left hand side and anion on the right hand side.*

Step 2. *Enclose the compound ion (if any) in a bracket.*

Step 3. *Reduce the valency numerals to a simple ratio by dividing with a common factor, if any.*

Step 4. *Criss-cross the valencies, i.e., shift the valency numerals crosswise to the lower right-hand corner of the ions. This is done to achieve electrical neutrality.*

Let us apply the above steps to write formula of **calcium phosphate**.

Step 1. Writing the formulae of the ions.

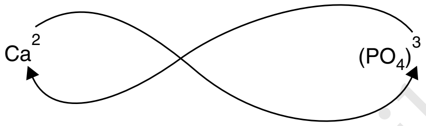
Ca^{2+} PO_4^{3-}

Step 2. Enclose the compound ion phosphate in a bracket.

Ca^{2+} $(PO_4)^{3-}$

Step 3. Not applicable, because ratio is already simple.

Step 4. *Criss-cross the valencies.*



Thus, the formula of calcium phosphate is: $\mathbf{Ca_3(PO_4)_2}$

On the basis of above steps the formulae of some ionic compounds are:

(i) Potassium chlorate ($KClO_3$) (ii) Lithium nitride (Li_3N)

(iii) Sodium oxalate [$Na_2C_2O_4$] (iv) Calcium cyanide [$Ca\ (CN)_2$]

(v) Magnesium nitride (Mg_3N_2)

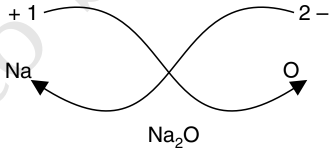
(vi) Potassium manganate (K_2MnO_4)

Example 1.8: *Write down the formulae of*

(i) *Sodium oxide* (ii) *Magnesium hydroxide*

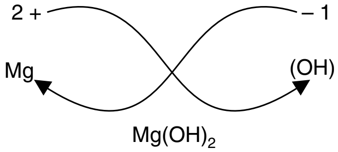
Solution:

(i) Sodium oxide contains Na^+ and O^{2-} ions.



The formula of sodium oxide is $\mathbf{Na_2O}$.

(ii) Magnesium hydroxide



The formula of magnesium hydroxide is $Mg(OH)_2$.

Example 1.9: *Write names of the compounds represented by the following formulae.*

(i) $Al_2(SO_4)_3$ (ii) $CaCl_2$

Solution:

(i) Aluminium sulphate (ii) Calcium chloride

Group Activity: Playing Game for Writing Formula

Make placards with symbols and valencies of the elements separately. Each student should hold two placards, one with the symbol in the right hand and the other with the valency in the left hand. Keeping the symbols in place, students should criss-cross their valencies to form the formula of a compound.

1.3.3. Types of Formula

There are different types of formulas. These include *empirical*, *molecular structural* formulas. In this section, we will study empirical and molecular formulas.

1.3.3.1. Empirical Formula

The empirical chemical formula represents the relative number of atoms of each element in the compound. Some compounds, like water, have the same empirical and molecular formula, because they are small and have the same ratio of atoms in molecules and number of atoms in a molecule. The empirical and molecular formula for water looks like this:

$$H_2O$$

The empirical formula is determined by the weight of each atom within the molecule. Therefore, for a slightly bigger molecule like hydrogen peroxide, the empirical formula shows only the ratio of atoms. In this case:

$$HO:$$

However, this empirical chemical formula only shows the basic foundation of the molecule. In reality, two HO: molecules come together to form a hydrogen peroxide molecule.

1.3.3.2. Molecular Formula

The molecular formula comes in to show the actual number of atoms within each molecule. Thus, for hydrogen peroxide the molecular formula is thus:

$$H_2O_2$$

As you can see, this somewhat confuses the actual structure of hydrogen peroxide. While the empirical chemical formula gives clues that the molecule has two oxygen atoms bonded together in the middle, the molecular formula does not make that clear at all. However, the molecular formula is often used to describe molecules, simply because it is convenient and most molecules can be looked up after their formula is identified.

1.3.4. Naming Compounds

The simplest compounds are binary compounds. *A binary compound is a compound that contains only two elements.* For example, NaCl is a binary compound of sodium and chlorine. CaF_2 is binary compound of calcium and fluorine. It may be noted that a binary compound may contain more than two atoms.

While writing the formula of a binary compound the symbol of the more electronegative element is written on right hand side while that of less electronegative element is written on left hand side. The number of atoms of each element are indicated by subscripts on the right hand side bottom of the symbol. For example, for a compound formed by combination of a metal with a non-metal, the symbol of the metal element is written first (left hand side) and the symbol of the non-metal element is written on right hand side.

While naming the binary compounds, the first element (less electronegative element) is named as such while the name of the second element (more electronegative element) is written with an-ide ending. For example,

KI is named as potassium iodide.

$CaCl_2$ is named as calcium chloride.

The names of some non-metallic elements with-*ide* endings are given below:

Hydrogen — Hydride
Fluorine — Fluoride
Chlorine — Chloride
Oxygen — Oxide
Sulphur — Sulphide
Carbon — Carbide

While naming binary compounds of metals and non-metals, the subscript numerals are ignored. For example, $BaCl_2$ *is named barium chloride and not barium dichloride.*

On the other hand, *while naming the binary compounds of two non-metals, the subscript numerals have to be taken into consideration* and are indicated as a part of the name. The reason for using prefixes is that the same two non-metallic elements may combine to form many compounds. For example, phosphorus and chlorine combine to form two compounds PCl_3 and PCl_5. A subscript 2 is indicated by the prefix *di*; subscript 3 by *tri*; subscript 4 by tetra, and so on. There should be no gap between the prefix and the name of the element. Some examples are given in Table 1.5.

Table 1.5. Names of some Binary Compounds of two Non-metals

Formula	Name	Formula	Name
CO	Carbon monoxide	CCl_4	Carbon tetrachloride
CO_2	Carbon dioxide	PCl_5	Phosphorus pentachloride
SO_2	Sulphur dioxide	N_2O_4	Dinitrogen tetroxide
PCl_3	Phosphorus trichloride	N_2O	Dinitrogen oxide

1.4. APPARATUS AND SAFETY RULES

1.4.1. Laboratory Apparatus

Experiments cannot be carried out in the absence of scientific apparatus and equipment. Let us discuss about some common laboratory apparatuses.

The common laboratory apparatuses are:

(a) **Beaker:** A beaker is a simple container commonly used in laboratories for mixing and heating liquids. They are generally cylindrical in shape, with a flat bottom.

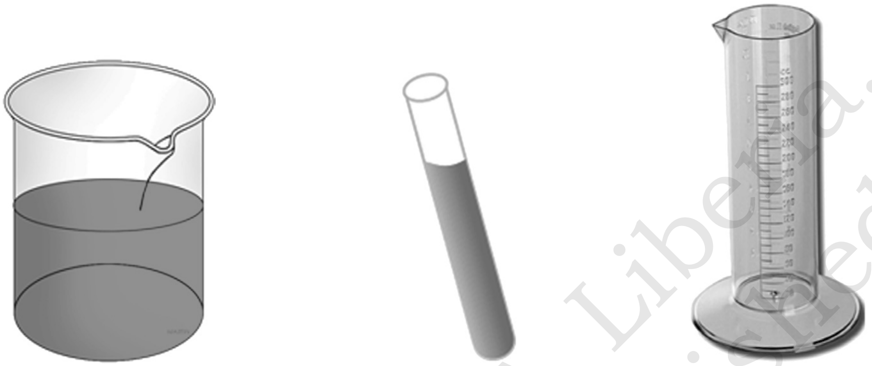
(b) **Test tube:** A thin cylindrical glass tube closed at one end, used to hold small amounts of material for laboratory testing or experiments.

(c) **Measuring cylinder:** A measuring cylinder is used to measure the volume of a liquid in laboratory. It has a narrow cylindrical shape. Each marked line on the measuring cylinder represents the amount of liquid that has been measured.

(d) **Separating funnel:** A separating funnel is a laboratory glassware used in liquid-liquid extractions to separate (partition) the components of a mixture into two immiscible solvent phases of different densities.

(e) **Pipette:** A pipette, pipet, or chemical dropper is a laboratory tool commonly used in chemistry, biology and medicine to transport a measured volume of liquid, often as a media dispenser.

(f) **Balance:** Weighing scales are used to measure weight or mass of the chemicals.



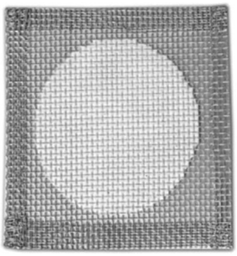
(a) Beaker (b) Test tube (c) Measuring cylinder

(d) Separating funnel (e) Pipette (f) Balance

Fig. 1.5. Some common Laboratory Apparatuses

Besides, the other apparatuses used in a laboratory are test tube stand, wash bottle, China dish, Petri dish, tripod, wire gauze, water condenser and watch glass. We will study different experiments using these laboratory apparatuses in some of the subsequent units.

Test tube stand



Wire gauze

Wash bottle

China dish Petri dish Tripod

Fig. 1.6. Some other Laboratory Apparatuses

ACTIVITY 1.1

Making a Poster of Laboratory Apparatuses

Take a drawing sheet and draw various laboratory apparatuses on it. Label these apparatuses properly. Display this poster to the whole class.

1.4.2. Basic Safety Rules in the Laboratory

In a laboratory, improper handling of chemicals and equipment can cause injury or accident. Thus, we must follow safety rules inside the laboratory to avoid accidents and injuries. These rules are:

- Always wear lab-coats and closed shoes.
- Never taste any chemical or transfer chemicals by mouth pipette.
- Read labels carefully before using chemicals.
- Do not use any chemicals in the absence of teacher.
- Do not perform unauthorized experiments.
- Keep your lab space clean and organized.
- Do not leave an on-going experiment unattended.
- Check your glassware for cracks and chips each time you use it.
- Safety glasses must be worn whenever chemicals are being heated, mixed or poured.
- Wash hands before leaving the lab and before eating.
- If any accident (burns) or injury (cut) occurs, immediately call your teacher for help.
- Before leaving a lab unattended, turn off all ignition sources and lock the doors.

Group Activity: Discussing Laboratory Safety Rules

In group, discuss some laboratory safety rules. Students should note these on their notebooks and apply them in the laboratory.

1.4.2.1. Some Hazard Signs

Following are some hazard signs:

Electrical Hazard Combustible Materials No Open Flames

Poison or Toxic Fire Extinguisher Non-potable Water

Fig. 1.7. Some Safety Hazard Signs

GLOSSARY

- **Accuracy:** The quality or state of being correct.
- **Chemistry:** The branch of science that studies the preparation, properties, structure and reactions of material substances.
- **Combustion:** The process where a substance burns in the presence of Oxygen, giving off heat and light in the process.
- **Hypothesis:** A possible solution to a problem, based on knowledge and research.
- **Metric System:** The decimal measuring system based on the metre, litre, and gram as units of length, capacity, and weight or mass.

- **Precision:** The quality, condition, or fact of being exact and accurate.
- **Transmutation:** Conversion of one chemical element into another.

SUMMARY

- Chemistry is the branch of science that studies the preparation, properties, structure and reactions of material substances.
- In the fourth century BC, two Greek philosophers, Democritus and Leucippus, suggested that matter was not infinitely divisible into smaller particles but instead consisted of fundamental, indivisible particles called atoms.
- The 16th and 17th centuries saw the beginnings of the modern chemistry.
- In 1661, Robert Boyle (1627–91) published "The Sceptical Chymist," which described the relationship between the pressure and the volume of air.
- Amedeo Avogadro (1776-1856) introduced the hypothesis that equal volumes of gases at the same pressure and temperature contain the same number of molecules.
- Measurement of any physical quantity involves comparison with a certain basic, arbitrarily chosen, internationally accepted reference standard called unit.

EVALUATION

I. Multiple Choice Questions

1. This branch deals with the study of compounds of other elements except carbon.
   (a) Physical Chemistry (b) Industrial Chemistry
   (c) Inorganic Chemistry (d) Bio-Chemistry
2. How many base units are there?
   (a) Three (b) Five
   (c) Seven (d) None of these
3. Which of the following is NOT an SI unit?
   (a) kelvin (b) metre
   (c) ampere (d) yard.

4. The expression of the composition of a substance by chemical symbols and numerical subscripts is called chemical _________.

(a) symbol (b) formula

(c) Both (a) & (b) (d) None of these.

5. In a laboratory, improper handling of chemicals and glassware can cause

(a) injury (b) accident

(c) both (a) and (b) (d) neither (a) nor (b).

II. State True or False

1. Antoine Lavoisier published "*The Sceptical Chymist*," which described the relationship between the pressure and the volume of air.

2. The scientific method helps to organize thoughts and procedures.

3. Joseph Gay Lussac stated the law of definite proportions.

4. Stapler and calculator are examples of laboratory apparatus.

5. We should not return chemicals or reagents to bottles.

III. Answer the Following Questions

1. Define chemistry.

2. Convert each of the following:

(i) 5 gram to centigram

(ii) 25000 millimeter to meter

(iii) 1200 seconds to minutes

(iv) 40° Celsius to Fahrenheit

3. Express the following in the scientific notation.

(i) 3,86,000 (ii) 9007

(iii) 0.02683 (iv) 9000.0

4. Express the result of the following calculation to the appropriate number of significant digits.

$816 \times 0.02456 + 215.67$.

5. Write at least five symbols of atoms which were proposed by Dalton.

6. Write down the formulae of:

(i) aluminium chloride (ii) magnesium hydroxide.

7. Write names of the compounds represented by the following formulae.

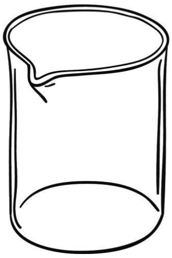
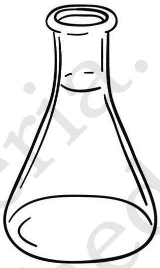
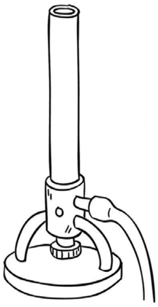
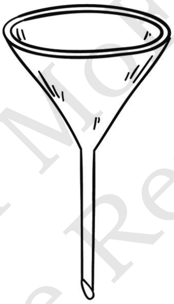
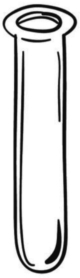
(i) K_2SO_4 (ii) KNO_3

8. Give the names of the elements present in the following compounds:

(i) Baking powder (ii) Aluminium chloride

9. Write five safety rules of laboratory.

10. Identify the following objects and write their names in your notebook.

		
(a)	(b)	(c) ....
		
(d)	(e)	(f) ....

SEMESTER–I (Period–I)

C10CH2

2
TOPIC

Matter and its Properties

Learning Objectives

By the end of this topic, learners will be able to:

- Discuss the states of matter.
- Identify physical and chemical changes.
- Make use of separation techniques to purify impure substances.
- Utilize various techniques to separate mixtures.

Introduction

Anything that occupies space and has mass is called **matter**. All the materials present in this world such as air and water, hydrogen and oxygen, sugar and sand, silver and steel; copper and coal, iron and wood, ice and alcohol, milk and oil, kerosene and petrol, carbon dioxide and steam, carbon and sulphur, rocks and minerals are different kinds of matter as they all occupy space and have mass. There are also things such as television signals, radio signals, electricity, light, sound, friendship, love, affection, hatred, thoughts, ideas, taste and smell that do not occupy space and do not have mass. Such things are not matter. This is one way of classifying things. There are other ways of classifying things such as on the basis of their physical properties and chemical properties. On the basis of physical properties, the matter is classified into solids, liquids, and gases. On the basis of chemical properties, matter is classified into elements, compounds, and mixtures.

2.1. STATES OF MATTER AND THEIR CHANGES

Observe different types of matter around you. What are its different states? We can see that matter around us exists in three different states—solids, liquids and gases.

(a) Ice is a solid (b) Water is a liquid (c) Steam is a gas

Fig. 2.1. Three States of Matter

Let us classify the substances around us into solids, liquids and gases.

ACTIVITY 2.1

Grouping Substances into Solid, Liquid and Gas

Group the following substances into solid, liquid, and gas in your exercise book.

Solid	
Liquid	
Gas	

2.1.1. Changes of States of Matter

We now know that matter exists in three physical states—solid, liquid, and gaseous state. The physical state of matter can be changed by changing the temperature, *i.e.* by heating it or cooling it. Let's perform the following activity to understand this.

ACTIVITY 2.2

To Demonstrate the Change of States of Matter

Materials Required:

A kettle, a frying pan, ice cubes, water, a beaker and a burner.

Procedure:

Take some water in the kettle and few ice cubes in the frying pan and arrange the set-up as shown in Fig 2.2.

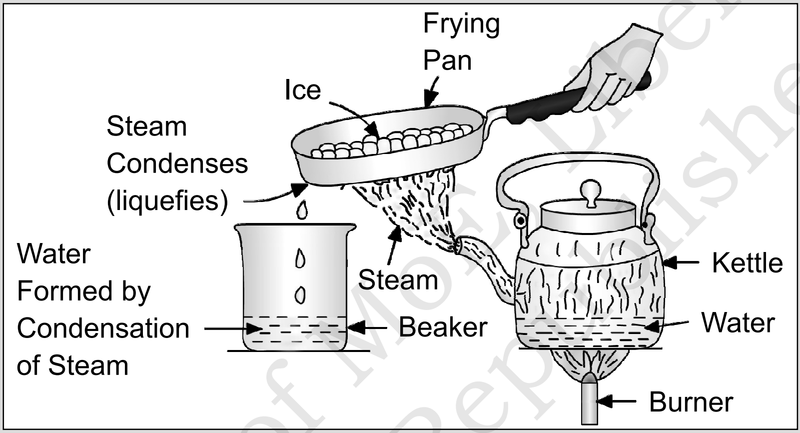


Fig. 2.2. Set-up to Demonstrate Change of State of Matter

Observation:

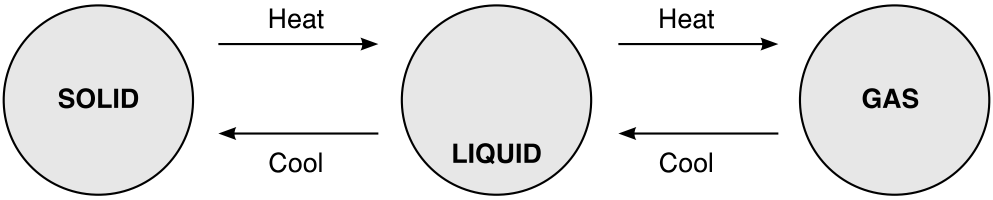
On heating the kettle, you will see steam coming out of the spout of the kettle. When the hot steam coming out of the spout comes in contact with the ice-cold bottom of the frying pan, it gets cooled and converts to liquid water.

Conclusion:

The formation of steam by boiling of water is a change of state of matter from liquid to gas. The formation of water drops by cooling the steam is a change of state of matter from gas to liquid.

As observed in the Activity 2.2, above, when water is heated, it converts into steam (liquid to gas change). The steam when cooled, converts into liquid (gas to liquid change). Thus, depending on the temperature, a substance can exist in all the three physical states—solid, liquid, and gas.

When a solid melts to form a liquid or a liquid evaporates to form a gas, we say that there is a **change of state**. There is also a change of state when a gas is cooled, and this change from gas to a liquid is called **condensation**. The change of state from solid to liquid is called **melting**.



2.2. PROPERTIES AND CHANGES OF MATTER

2.2.1. Properties of Matter

The states of matter have different properties. The properties of matter depend upon the arrangement of atoms or molecules in them. The atoms or molecules may be tightly packed or loosely packed. The arrangement of particles decides whether a given substance will exist as a solid, a liquid or a gas.

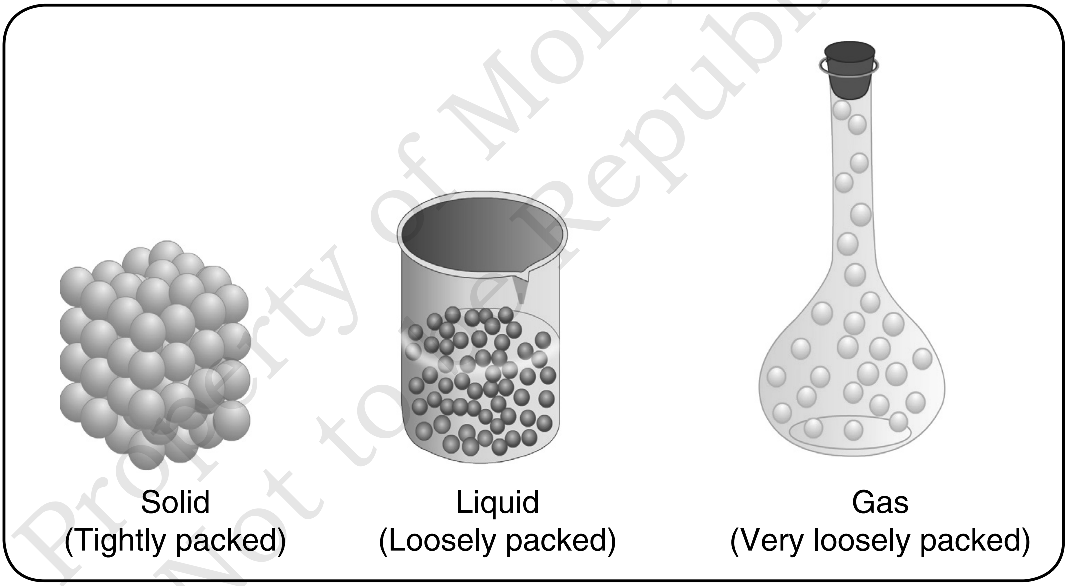


Fig. 2.3. Arrangement of Particles in Solids, Liquids and Gases

2.2.1.1. Properties of Solids

Some of the characteristic properties of solid are:

(i) Particles in a solid are closely packed.

(ii) The matter in solid state possesses *a definite volume*, *a definite shape*, distinct boundaries and *a definite mass*.

(iii) Solids are *rigid* and *almost incompressible*.

(iv) Solids generally possess *high densities.*

(v) Solids *do not exhibit diffusion.*

Some common **examples** of solids are: table, chair, common salt, silver, ice, diamond, etc.

Solids generally do not exhibit diffusion due to smaller interparticle spaces and absence of translatory motion. However, on rare occasions we can observe diffusion of solid materials, but to a very limited extent. It is observed that if we write something with chalk on a blackboard and leave it as such for a few days, it becomes difficult to clean. This is probably due to diffusion of chalk particles into the surface of the blackboard.

2.2.1.2. Properties of Liquids

Some of the characteristic properties of liquids are:

(i) Particles in a liquid are loosely packed.

(ii) The matter in liquid state possesses *a definite volume, a definite mass,* but *no definite shape.*

(iii) Liquids are also *almost incompressible* but are *not rigid.* In fact, they can flow and acquire the shape of the container in which they are kept.

(iv) Liquids can *undergo diffusion.*

(v) Liquids also have *high densities* but less than that of solids.

Some **examples** of liquids are: milk, water, alcohol, petrol, kerosene, fruit juices, etc.

Solids, liquids as well as gases can diffuse into liquids. This is due to the fact that the interparticle spaces in liquids are larger and the particles in liquid state move freely. For example, when we add a few crystals of sugar to water they intermix (dissolve) with water spontaneously. Similarly, when we add few drops of ink to water, the colour of the ink gets dispersed evenly in the entire liquid. The gases also diffuse into liquids. Aqueous solution of ammonia contains ammonia diffused in water.

2.2.1.3. Properties of Gases

Some of the characteristic properties of gases are:

(i) Particles in a gas are very loosely packed.

(ii) The matter in gaseous state *has neither definite volume nor definite shape but it has definite mass.* It acquires the shape and volume of the container.

(iii) *Gases are highly compressible.* For example, natural gas in compressed form is used as fuel (Compressed Natural Gas—**CNG**) in internal combustion engines. Oxygen supplied to hospitals in cylinders is also in compressed form.

(iv) *The gases exhibit the property of diffusing very fast into other gases. This is because the particles in gases move very fast in all directions and there are large interparticle spaces.*

(v) *Gases exert pressure on the walls of the container in which they are stored.*

Some common **examples** of gases are: air, hydrogen, carbon dioxide, hydrogen sulphide, ammonia, oxygen, nitrogen, etc.

A comparison chart of the characteristic properties of solids, liquids and gases are given in Table 2.1.

Table 2.1. Comparison of Characteristic Properties of Solids, Liquids and Gases.

Property	Solids	Liquids	Gases
Arrangement of particles	Tightly packed	Loosely packed	Very loosely packed.
Shape	Definite	Take the shape of the container, but do not necessarily occupy all of it.	Take the shape of the container by occupying whole of the space available to them.
Volume	Definite	Definite	Not definite, take the volume of the container.
Compressibility	Almost nil	Almost nil	Very large. Gases are highly compressible
Fluidity or Rigidity	Rigid	Fluid	Fluid
Density	Large	Large	Very small.
Diffusion	Generally do not diffuse	Diffuse slowly	Diffuse rapidly.
Free surfaces	Any number of free surfaces	Only one free surface	No free surface.

2.2.2. Changes of Matter

Every matter exists in particular state under ordinary conditions. However, it can be changed from one state to another by changing the conditions. The changes are of two types—(i) Physical change and (ii) Chemical change.

2.2.2.1. Physical Changes

ACTIVITY 2.3

Demonstrating Physical Changes of Matter

Cut a piece of paper in four square pieces. Cut each square piece further into four square pieces. Lay these pieces on the floor or a table so that the pieces acquire the shape of the original piece of paper.

Obviously, you cannot join the pieces back to make the original piece, but is there a change in the property of the paper?

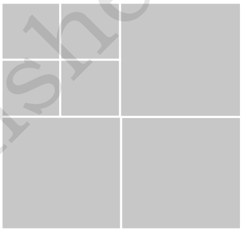


Fig. 2.4.

In Activity 2.3 above, you saw that paper underwent changes in size.

Properties such as shape, size, colour and state of a substance are called its **physical properties**. *A change in which a substance undergoes a change in its physical properties is called a **physical change***. A physical change is generally reversible. In such **a change no new substance is formed**.

Some more examples of **physical changes** are:

1. Evaporation of alcohol.
2. Formation of crystals on cooling a saturated solution of sugar.
3. Melting of ice.
4. Evaporation of sea water.

2.2.2.2. Chemical Changes

If you leave some iron nails in the open for sometime, it acquires a film of brownish substance (Fig. 2.5). This substance is called **rust** and the process is called **rusting**.

At home you must have seen hoe and spades getting rusted when exposed to the atmosphere for some time (Fig. 2.6).Rust is not iron. It is different from iron on which it gets deposited.

Fig. 2.5. Rusted nails

Fig. 2.6. Rusted hoe

Let us consider a few more changes where new substances are formed.

ACTIVITY 2.4

Demonstrating Chemical Changes of Matter

This activity can be demonstratered by the teacher.

CAUTION: Be careful while handling a flame.

Get a small piece of a thin strip or ribbon of magnesium. Clean its tip with sandpaper. Bring the tip near a candle flame. It burns with a brilliant white light (Fig. 2.7).

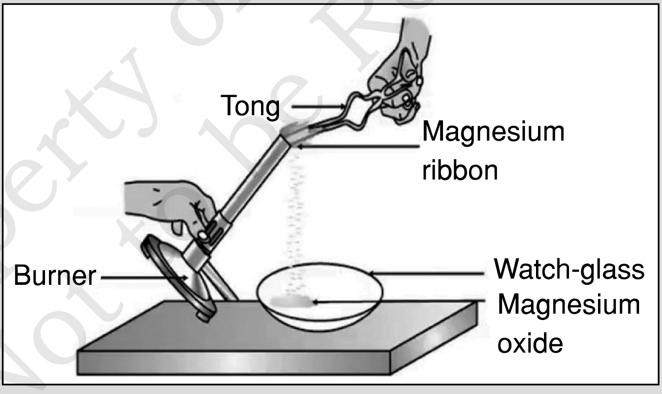


Fig. 2.7. Burning of Magnesium Ribbon

When it is completely burnt it leaves behind a powdery ash.

Does the ash look like the magnesium ribbon?

Collect the ash and mix it with a small amount of water. Stir the mixture (aqueous solution) well. Test the mixture with blue and red litmus papers.

Does the mixture turn red litmus blue?

Does the mixture turn blue litmus red?

When we burn the ribbon of magnesium with flame it forms a new substance (ash). This change can be written in the form of following equation:

Magnesium (Mg) + Oxygen (O_2) → *Magnesium oxide (MgO)*

On dissolving the ash in water it forms a new substance. This change can be written in the form of the following equation:

Magnesium oxide (MgO) + Water (H_2O) → *Magnesium hydroxide* [$Mg(OH)_2$]

Here, magnesium oxide and magnesium hydroxide are two new substances.

A change in which one or more new substances are formed is called a ***chemical change.*** A chemical change is also called a **chemical reaction**.

Some examples of chemical changes are:

1. Digestion of food
2. Burning of coal
3. Setting cement
4. Cooking of food

Some differences between physical and chemical changes are shown in Table 2.2.

Table 2.2. Differences between Physical and Chemical Change.

	Physical Change	**Chemical Change**
1.	It is temporary change.	It is a permanent change.
2.	No new substances are formed.	New substances with quite different properties are formed.
3.	Only the physical properties of the substance undergo change. Composition remains same.	Both the chemical composition and properties of the substance change.
4.	It can be reversed by physical means.	It cannot be reversed by physical means.

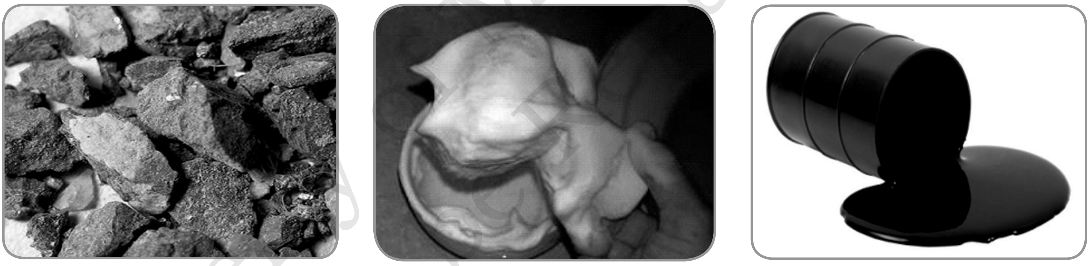
2.3. MIXTURE AND ITS CLASSIFICATION

2.3.1. What is Mixture?

Most substances around us are not pure. They are mixtures of two or more pure substances—elements or compounds. The pure substances can be mixed in different ways. The air we breathe is not a single, pure substance rather it is a mixture of different gases. River water consists of solid particles. Thus, *a* ***mixture*** *is a substance which consists of two or more elements or compounds not chemically combined together*. All the mixtures are impure substances because they contain more than one kind of particles. Some examples of mixtures are:

- Sea water—it has several salts dissolved in water
- Sugar solution—a mixture containing mainly sugar and water
- Air—a mixture of several gases such as oxygen, carbon dioxide, nitrogen, rare gases, water vapours and dust particles.

Sea water Brass Soft drinks



Minerals Shaving cream Petroleum

Fig. 2.8. Some Examples of Mixtures

2.3.2. Classification of Mixtures

We now know that different substances are present in a mixture. These substances are called **components** or **constituents of the mixture**. These constituents may be completely mixed together, or may remain separate, or a substance is diffused throughout another substance.

The mixtures containing its constituents completely mixed together in a way not distinguishable from one another are called ***homogeneous mixtures***, for example, sugar solution in water.

In some mixtures, the constituents are not completely mixed and are distinguishable from each other. Such mixtures are known as ***heterogeneous mixtures***, for example, the mixture of sugar and sand. This is one way of classifying mixtures.

The mixtures are also classified on the basis of the nature of their constituents. The constituents may be solid-solid, solid-liquid, liquid-liquid, solid-gas, gas-liquid and gas-gas.

Solid-solid Mixtures

Mixtures of this class consist of two or more solid substances that have been physically combined.

Examples

- Brass (copper and zinc)
- Steel (copper and iron)
- Sulphur and iron filings
- Sand and salt

Solid-liquid Mixtures

Mixtures of this class consist of a solid and a liquid in which the solid is soluble.

Examples

- Sugar and water
- Salt and water
- Sugar and milk
- Chalk and water

Liquid-liquid Mixtures

Mixtures of this class consist of two liquids which are miscible.

Examples

- Oil and water
- Petrol and kerosene

Solid-gas Mixtures

Mixtures of this class consist of a solid and a gas.

Examples

- Smoke
- Hydrogen gas absorbed by palladium

Gas-liquid Mixtures

The mixtures of this class consist of a gas and a liquid.

Examples

- Foam, hair spray, shaving cream
- Soda water (Water and carbon dioxide gas)
- Hydrochloric acid (Water and HCl gas)

Gas-gas Mixtures

The mixtures of this class consist of two or more gases which are diffusible into each other.

Examples

- Air (Mixture of oxygen, nitrogen, carbon dioxide, argon and water vapour)

2.4. STANDARD SEPARATION TECHNIQUES FOR MIXTURE

Most of the substances around us are not pure. They are mixture of two or more pure substances.

Let us discuss some techniques of separating constituent of different types of mixtures.

2.4.1. Separating Constituents of Solid-Solid Mixtures

This type of mixtures can be separated by the following methods:

2.4.1.1. Sieving

This method is used to separate mixtures that contain substances mostly of different sizes. The mixture is passed through the pores of the sieve. All the smaller substances pass through easily while the bigger components of the mixture are retained.

Fig. 2.9.Sieving

2.4.1.2. Magnetic Separation

This method is applied when one of the components of the mixture is magnetic. For example, a mixture of iron filings and sand can be separated by using a magnet.

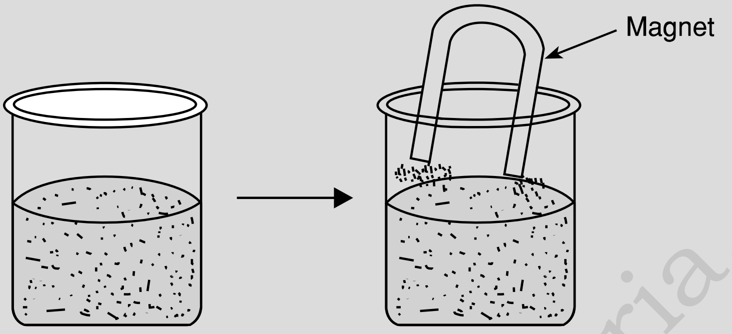
Let us demonstrate this activity to understand the technique of magnetic separation.

ACTIVITY 2.5

Demonstrating the Process of Magnetic Separation

1. Mix up the iron filings and the sand in glass jar (Fig. 2.10).
2. Bring the magnet closer to the mixture.

   What do you observe?



Mixture of iron filing and sand

Separating the mixture by removing iron filings

Fig. 2.10. Magnetic Separation of Mixtures

We observe that iron filings get stuck to the magnet.

2.4.2. Separating Constituents of Solid-Liquid Mixtures

This type of mixtures can be separated by the following methods:

2.4.2.1. Filtration

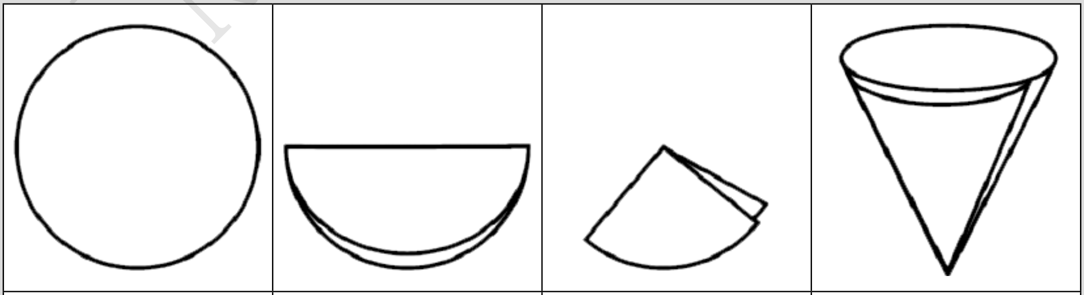
This method is applied for separating a heterogeneous mixture containing a solid component suspended in liquid component. A mixture containing sand and water can be separated by this method.

Let us demonstrate this activity to understand the technique of filtration.

ACTIVITY 2.6

Demonstrating the Separation of a Mixture Using Filtration

1. Take a filter funnel and fold it in the form of a cone as shown in Fig. 2.11.



Filter paper	Folded once	Folded twice	Opened

Fig. 2.11. Making Filter Paper cone

2. Fit the cone in a funnel and moisten so that it sticks to the surface of the funnel.

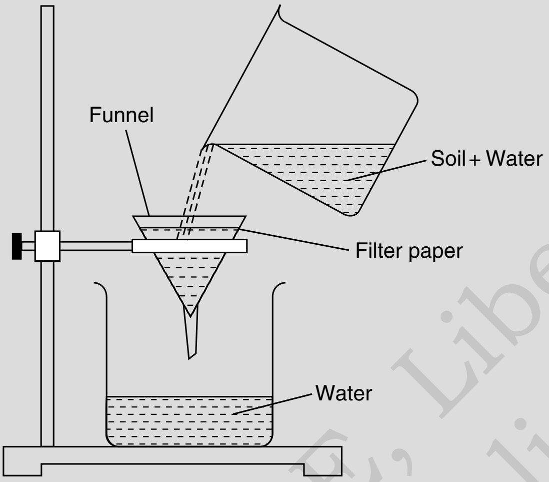


Fig. 2.12. Method of Filtration

3. Place the funnel in a funnel stand(Fig. 2.12).
4. Carefully pour some of the mixture down a glass rod (stirrer) and into the filter funnel.
5. The purpose of the glass rod is to prevent any splashing and to make sure that the mixture does not get between the paper and the funnel.
6. Open the filter paper and allow the soil (the residue) to dry.

   What is the basis for separation of this mixture?

The soil remains behind on the filter paper. The water passes through the filter paper.

2.4.2.2. Evaporation

This method is applied when the solvent is volatile liquid and the solute is non-volatile solid or liquid. For example, sodium chloride can be separated from its aqueous solution by evaporation. Similarly, the coloured dye present in the blue or black ink can be separated by evaporation (Fig. 2.13). The volatile component evaporates leaving behind non-volatile component.

Let us demonstrate this activity to understand the techniques of evaporation.

ACTIVITY 2.7

Demonstrating Separation of Mixture Using Evaporation

1. Place an evaporating dish on a beaker containing some water.
2. Put a few drops of the ink on the evaporating dish.
3. Place the beaker along with the evaporating dish on a tripod stand. (Fig. 2.13)

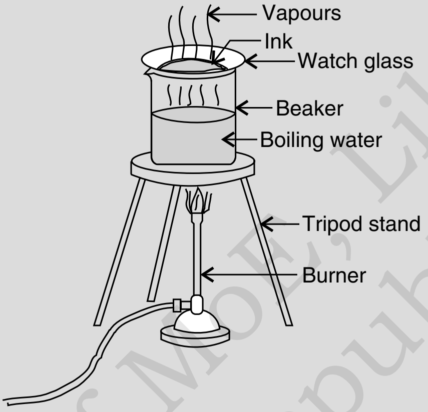


Fig. 2.13. Separation by Evaporation

4. Start heating.
5. The ink on the evaporating dish starts evaporating.
6. Continue heating. Finally, a residue is left on the dish. This is the dye present in the ink.

2.4.3. Separating Constituents of Liquid-Liquid Mixtures

This type of mixtures can be separated by the following methods:

2.4.3.1. Distillation

This method is applied for the separation of a mixture containing two miscible liquids having large difference in their boiling points or a mixture containing solid dissolved in liquid. Distillation involves vaporization followed by condensation. In this method both the components are recovered.

Let us study the separation of a mixture containing petrol and kerosene.

ACTIVITY 2.8

Demonstrating Separation of Mixture Using Distillation

1. Take the mixture in a distillation flask and fit it with the thermometer (Fig. 2.14).

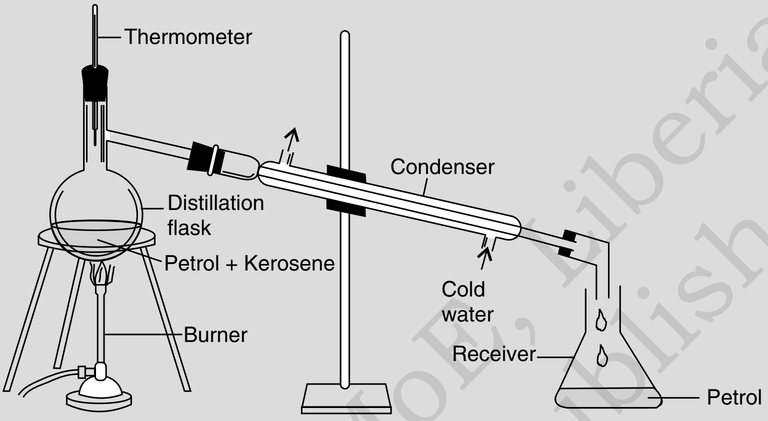


Fig. 2.14. Separation by Distillation

2. Heat the mixture slowly.
   What do you observe?

As the boiling point of petrol is lower than that of kerosene, so, petrol vaporises first. It condenses in the condenser and is collected from the outlet.

And thus kerosene is left in the flask.

2.4.3.2. Using a Separating Funnel

A mixture of two immiscible liquids can be separated into its components by using a separating funnel. The separating funnel has a tap in its stem and a stopper in its mouth (Fig. 2.15). It can be used for separating the mixtures such as kerosene and water, water and benzene, water and carbon tetrachloride, etc.

The principle of this method is that immiscible liquids separate out in layers depending on their densities.

ACTIVITY 2.9

Demonstrating Separation of Mixture Using a Separating Funnel

In this activity, you will study the separation of components of a mixture containing kerosene oil and water using a separating funnel.

1. Pour the mixture of kerosene oil and water in a separating funnel (Fig. 2.15).

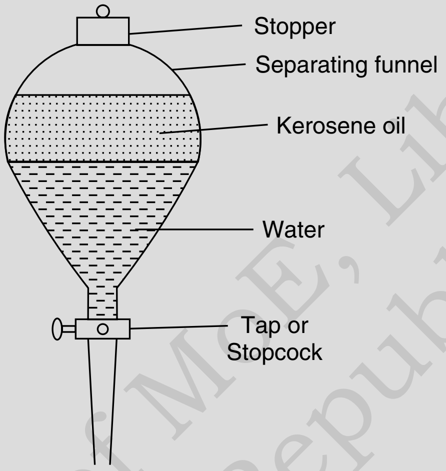


Fig. 2.15. Separation of a mixture of two immiscible liquids.

2. Allow the mixture to stand for some time so that separate clear layers of oil and water are formed. Water being heavier forms the lower layer.
3. Open the stopcock of the separating funnel and pour out the lower layer in a beaker.
4. Close the stopcock when the oil reaches the stopcock. The kerosene oil remains in the separating funnel.

GLOSSARY

- **Condensation:** The process in which water vapour changes to liquid water on cooling.
- **Diffusion:** The movement of molecules from a region of higher concentration to a region of lower concentration.
- **Distillation:** A method of separating mixtures based on differences in their volatilities in a boiling liquid mixture.
- **Evaporation:** The process in which liquid water changes to vapour on heating.

- **Filtration:** Method of separating heterogeneous mixture containing a solid component suspended in liquid components.
- **Matter:** Anything that occupies space and has mass.
- **Translatory motion:** Motion in which all points of a moving body move uniformly in the same direction.

SUMMARY

- Anything that occupies space and has mass is called matter.
- On the basis of physical properties, the matter is classified into solids, liquids, and gases.
- On the basis of chemical properties, matter is classified into elements, compounds, and mixtures.
- Matter around us exists in three different states—solids, liquids and gases.
- When a solid melts to form a liquid or a liquid evaporates to form a gas, we say that there is a change of state.
- Matter can be changed from one state to another by changing the conditions. Changes can be of two types—
    (i) physical change and
    (ii) chemical change.
- Properties such as shape, size, colour and state of a substance are called its physical properties.
- A change in which a substance undergoes a change in its physical properties is called a physical change.
- A change in which one or more new substances are formed is called a chemical change.
- A mixture is a substance which consists of two or more elements or compounds not chemically combined together.
- The mixtures containing its constituents completely mixed together in a way not distinguishable from one another are called homogeneous mixtures, for example, sugar solution in water.
- The mixture in which, the constituents are not completely mixed and are distinguishable from each other are known as heterogeneous mixtures, for example, the mixture of sugar and sand. This is one way of classifying mixtures.

- Based on the nature of constituents, mixtures are of the following types:
  - (i) Solid-solid
  - (ii) Solid-liquid
  - (iii) Liquid-liquid
  - (iv) Solid-gas
  - (v) Liquid-gas
  - (vi) Gas-gas

EVALUATION

I. Multiple Choice Questions

1. Which of the following is a matter?

(a) Love
(b) Thought
(c) Cold
(d) Cold drink.

2. The state of the matter that can be rapidly diffused is

(a) solid
(b) liquid
(c) gas
(d) none of these

3. Which of the following are physical properties of a matter?

(a) shape
(b) colour
(c) size
(d) All of these

4. In which change, one or more new substances are formed?

(a) Physical change
(b) Chemical Change
(c) Biological change
(d) None of these

5. Which of the following is an example of gas-gas mixture?

(a) Air
(b) Sugar solution
(c) Sea water
(d) None of these

II. State True or False

1. Gaseous state of matter is highly compressible.
2. Particles in a liquid are closely packed.
3. Physical change is reversible.
4. The change of state from solid to liquid is called condensation.
5. Sieving is used to separate mixture containing substances of different sizes.

III. Answer the Following Questions

1. Define matter. Give some examples of matter.
2. What are the three states of matter?

3. State at least two characteristic properties each of:
   (a) a solid (b) a liquid
   (c) a gas.

4. What do you mean by a physical change? State any two examples of physical change.

5. What do you mean by a chemical change? State any two examples of chemical change.

6. Distinguish between physical and chemical changes.

7. Explain the term 'Mixture'. State four examples of mixture.

8. State the main points of differences between homogeneous and heterogeneous mixtures.

9. Define the following:
   (i) Solid-solid mixture (ii) Solid-liquid mixture
   (iii) Liquid-liquid mixture (iv) Gas-gas mixture

10. What is magnetic separation? Explain with the help of an activity.

11. What is evaporation? Explain with the help of an activity.

12. What is distillation? Explain with the help of an activity.

PROJECT

Cut out a mineral water bottle and make a filtration apparatus for yourself.

SEMESTER–I (Period–II)

C10CH3

3
TOPIC

Atomic Structure

Learning Objectives

Upon completion of this lesson, learners will be able to:

- Discuss contributors to atomic structures.
- Explain the arrangement of fundamental particles.
- Discuss the concept of atomic theories.
- Discuss atomic number and mass number and their relationship to isotopes.
- Discuss the four quantum numbers.
- Construct electronic configurations for atoms.
- Explain the rules and principles for filling in electrons.

3.1. HISTORY OF ATOMIC STRUCTURE

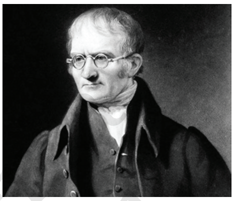
3.1.1. History of Atomic Chemistry

In the fifth century B.C. the Greek Philosopher Democritus proposed that all matter consists of very small indivisible particles called *atoms* (meaning uncuttable or indivisible). These earlier ideas were mere speculations and did not have any experimental basis. It was in 1808 that an English scientist, John Dalton formulated a precise definition of the indivisible building blocks of matter that are known as atoms. By the end of 19th century it was discovered that atoms consist of subatomic particles such as electrons, protons and neutrons. The **protons** are positively charged particles and are present in the nucleus of the atom. The **electrons** are negatively charged particles and are present in the extra-nuclear part of the atom. The **neutrons** are neutral particles and are present in the nucleus along with the protons. Most of the mass of the atom is concentrated in the **nucleus**.

During the past few years it has been found experimentally that some of the subatomic particles can be further split into two elementary particles, **quarks** and **leptons**.

3.1.2. Dalton's Atomic Theory

On the basis of laws of chemical combination John Dalton, an English school teacher in Manchester, proposed that behaviour of matter could be explained using an atomic theory. He published his work about atomic theory in 1808. The main points of **Dalton's atomic theory** are:



John Dalton (1766–1844)

1. Matter is composed of extremely small particles called atoms.
2. An element consists of only one type of atoms. They have identical properties such as mass, shape, colour, density, chemical properties, etc.
3. Atoms of one element differ from atoms of all other elements in mass, size and chemical properties.
4. Atom is the smallest particle that takes part in chemical reactions.
5. Atoms of different elements can combine in a fixed ratio to form compound.
6. Atoms can neither be created nor destroyed during chemical reactions.

3.1.2.1. Limitations of Dalton's Atomic Theory

Dalton's atomic theory was the first successful attempt which gave us some idea about the inner structure of matter. The main failures of Dalton's atomic theory are:

1. It failed to explain how atoms of different elements differ from each other, i.e., it did not tell anything about internal structure of the atom.
2. It could not explain how and why atoms of different elements combine with each other to form compound-atoms or molecules.
3. It failed to explain the nature of forces that hold together different atoms in a molecule.
4. It did not make any distinction between ultimate particle of an element that takes part in reactions (atom) and ultimate particle that has independent existence (molecule).

Towards the end of 19th century and in the beginning of 20th century new discoveries were made by Sir J.J. Thomson, Neils Bohr, Chadwick and others which revealed the inner structure of atom. In the light of these findings Dalton's atomic theory was suitably modified.

3.1.3. Discovery of Electron

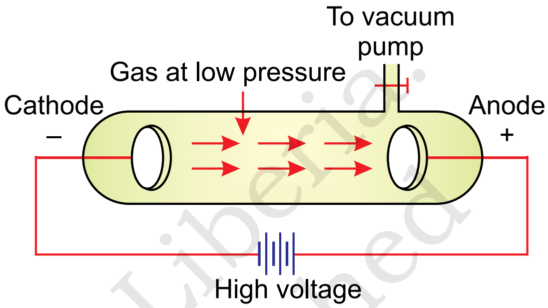


Fig. 3.1. Cathode rays

Electron was the first fundamental particle that was discovered. The credit for the discovery of electron goes to J.J. Thomson, a British physicist. Most of the information about electrons is obtained from the study of cathode rays, which were discovered during the experiments with gas discharge tubes. A discharge tube is long glass tube which is fitted with metal electrodes on either end, across which high voltage can be applied. The electrode which is connected to the negative terminal of the power source is called **cathode** while the electrode which is connected to the positive terminal is called **anode**. The tube is also connected to a vacuum pump for controlling the pressure of gas inside the discharge tube.

When the gas pressure inside the discharge tube is one atmosphere, no electric current flows through the tube. If the gas pressure is reduced to about 10^{-2} atmospheres and a potential difference of about 10000 volts is applied to the electrodes, an electric current flows and at the same time light is emitted by the gas. As the gas pressure in the discharge tube is reduced further, a dark space appears in the vicinity of cathode and alternate light and dark bands can be seen between the two electrodes. If the gas pressure is reduced to 10^{-6} atmospheres, the emission of light ceases, instead the end of the glass tube opposite to cathode glows (fluoresces) with a *faint greenish light*. Further investigations revealed that the fluorescence was caused due to the bombardment of the walls of the tube by rays emanating from cathode. These rays were called **cathode rays**. These rays were found to consist of negatively charged material particles, called **electrons**.

3.1.3.1. Properties of Cathode Rays

The cathode rays possess the following properties:

1. ***Cathode rays travel in straight lines.*** An object placed in the path of cathod rays casts a sharp shadow. It shows that cathode rays travel in straight lines.

2. ***Cathode rays consist of material particles.*** This was indicated by the fact that a light paddle wheel placed in the path of cathode rays starts rotating.
3. ***Effect of electric field.*** When electric field is applied to a stream of cathode rays, they get deflected towards positive plate (Fig. 3.2). It showed that cathode rays themselves are negatively charged.

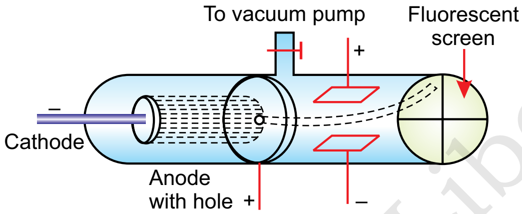


Fig. 3.2. Effect of Electric Field on Cathode Rays

4. ***Effect of magnetic field.*** When magnetic field is applied, the cathode rays get deflected. The direction of deflection again indicates that cathode rays are negatively charged.
5. On striking against walls of the discharge tube *cathode rays produce **faint greenish fluorescence***.

The above mentioned properties of cathode rays indicate that the cathode rays consist of a fast-moving stream of negatively charged material particles. These particles were named **electrons**.

3.1.3.2. Characteristics of Electron

Following are the characteristics of electron:

1. The charge to mass ration, e/m was found to be -1.76×10^8 coulombs/g or -1.76×10^{11} coulombs/kg.
2. Charge of electron is 1.60×10^{-19} coulomb.
3. Mass of electron is 9.11×10^{-31} kg, which is nearly equal to $\frac{1}{1840}$th of mass of an atom of hydrogen.

3.1.4. Discovery of the Proton

E. Goldstein, a German scientist, in 1886, discovered the existence of a new type of rays in the discharge tube. He used a perforated cathode (Fig. 3.3) in the discharge tube. The cathode divided the discharge tube in two chambers. On passing the electric discharge at low pressure he observed a new type of rays streaming behind the cathode. The path of these rays became visible due to the glow of the residual gas.

These rays also produced fluorescent glow on striking the walls of the tube behind cathode. These rays were named **anode rays** or **canal rays**. These rays were named canal rays because they passed through '*holes*' or '*canals*' in the cathode.

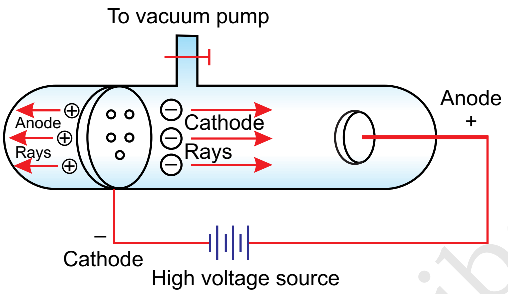


Fig. 3.3. Canal rays

Further investigations of these rays showed that they consist of positively charged material particles.

Some of the **characteristic properties of anode rays** are:

1. Anode rays consist of material particles.
2. Anode rays are deflected by electric field towards negatively charged plate.

   This indicates that *they are positively charged*.
3. The charge to mass ratio of the particles in the anode rays was determined by using Thomson's technique. Charge to mass ratio of the particles in the anode rays depends upon the nature of the gas taken in the discharged tube.

It was observed that e/m ratio was maximum when hydrogen gas was taken in the discharge tube. This indicated that positive ions formed from hydrogen are the lightest. These lightest positively charged particles were named **protons**. The charge and mass of the proton were determined, in the same manner as the one discussed in case of electron. Charge to mass ratio for protons was found to be **9.58 × 10^7 C/kg**. Charge on proton is opposite but equal in magnitude to the charge on the electron i.e , **1.60 × 10^{-19} C**. From these two observations mass of a proton works out to be **1.67 × 10^{-27} kg**. It is practically the same as the mass of a hydrogen atom and is about 1837 times the mass of an electron.

In general, a proton is represented as p^+.

A **proton** is a fundamental particle of atom carrying one unit positive charge and having mass 1.672 × 10^{-27} kg, which is nearly equal to the mass of an atom of hydrogen.

THOMSON'S MODEL OF ATOM

Thomson was the first to propose a detailed model of the atom. He proposed that an atom consists of a uniform sphere of positive electricity in which the electrons are distributed more or less uniformly. The negative and the positive charge are equal in magnitude. Thus, the atom as a whole is electrically neutral. This model of atom is known as the ***"Plum pudding model "*** because of its similarity to a christmas pudding with currants in it (Fig. 3.4).

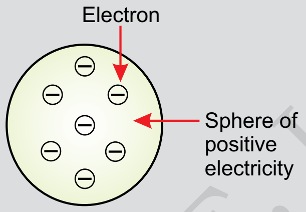
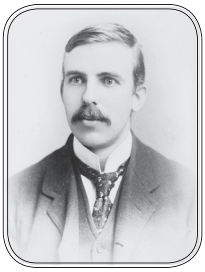


Fig. 3.4. Thomson's model of atom.

3.1.5. Discovery of Nucleus (Rutherford's Gold Foil Experiment)

Ernest Rutherford in 1911 performed an experiment which led to the downfall of Thomson's model. Most of the experimental work was carried out by Geiger and Marsden, two of Rutherford's students at Manchester University. Geiger and Marsden built the apparatus as shown in Fig. 3.5. The experiment involved the bombardment of a thin sheet of gold (thickness 4×10^{-5} cm) by α-particles. The source of α-particles was a piece of radium placed in a lead block. A narrow hole in the lead block allowed the α-particles to travel only in one direction through the evacuated vessel towards a fluorescent zinc sulphide screen.



Ernest Rutherford (1871-1937)

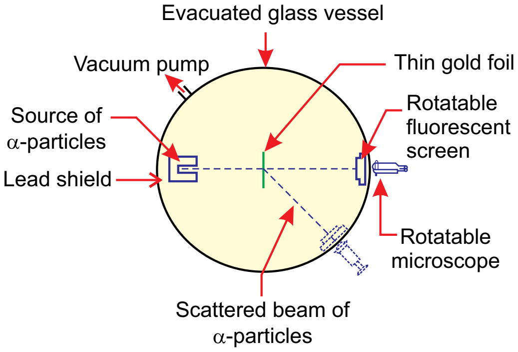


Fig. 3.5. The apparatus used by Geiger and Marsden to investigate the scattering of a-particles by a thin metal foil.

When an α-particle hits the fluorescent screen a tiny flash of light is produced which is observed with the help of a microscope. A thin gold foil was placed between the slit and the screen. The screen and the microscope were rotated to detect scattering of α-particles.

Rutherford observed that:

1. Most of the α-particles (nearly 99%) passed through the gold foil undeflected.
2. Some of the α-particles were deflected by small angles.
3. Very few particles (1 in about 10^4) were either deflected by very large angles or were actually reflected back along their path.

In order to explain the observations of his experiment, Rutherford assumed that the solid gold foil consists of layers of individual atoms which are touching each other so that there is hardly any empty space between them. As such the α-particles striking the gold foil must pass through the atoms. Rutherford explained his observations as follows:

1. Since most of the α-particles pass through the foil undeflected, it indicates that *the most of the space in an atom is empty* (Fig. 3.6).

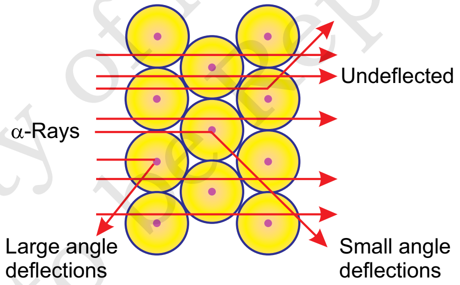


Fig. 3.6. Scattering of a-particles by gold atoms

2. α-particles being positively charged and having considerable mass could be deflected only by some heavy, positively charged centre. The small angle of deflection of α-particles indicates the presence of a *heavy positive centre in the atom.* Rutherford named this positive centre as **nucleus** (Fig. 3.7).

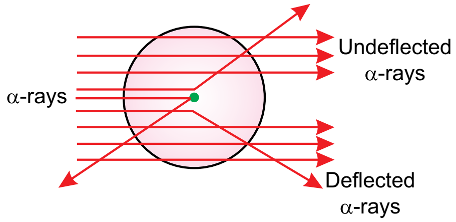


Fig. 3.7.

3. α-particles which make head-on collision with heavy positive centre are deflected through large angles. Since the number of such α-particles is very small, *the space occupied by the heavy positive centre must be very small.*

From the data of scattering experiment, Rutherford was able to calculate the radius of the nucleus. Rutherford calculated that the nucleus of an atom would have radius of about 10^{-14} m. He showed that the radius of the nucleus is about 10^{-4} times the radius of the atom which is about 10^{-10} m.

3.1.5.1. Rutherford's Nuclear Model of Atom

On the basis of scattering experiment Rutherford put forward *nuclear model of atom*. Main points of this model are:

1. *Most of the mass and all the positive charge of an atom is concentrated in a very small region called nucleus. Size of the nucleus is extremely small as compared with the size of the atom.* Radius of the nucleus is of the order of 10^{-15} m, whereas radius of atoms is of the order of 10^{-10} m.
2. *The nucleus is surrounded by electrons which are revolving around it at very high speeds. The electrostatic force of attraction between electrons and the nucleus is balanced by the centrifugal force acting on the revolving electrons.*
3. *Total negative charge on the electrons is equal to the total positive charge on the nucleus so that atom on the whole is electrically neutral.*

Nuclear model of atom can be compared with the solar system. In an atom electrons revolve around the nucleus in just the same way as the planets revolve around the sun. Due to this comparison, revolving electrons are sometimes called ***planetary electrons***.

Drawbacks of Rutherford's Model of Atom

Rutherford model failed in view of electromagnetic theory given by Maxwell. According to this theory a charged particle when accelerated emits energy in the form of electromagnetic radiation. According to Rutherford's model, electrons are revolving around the nucleus. This means, electrons would be in a state of acceleration all the time. Since electrons are charged particles, therefore, electron revolving in an orbit should continuously emit radiations. As a result of this, it would slow down and would no longer be able to withstand the attractive force of the nucleus. Hence, it would move closer and closer to the nucleus and would finally fall in the nucleus by following a spiral path (Fig. 3.8). This means atom should collapse. But actually we know atom is stable. Thus, *Rutherford's model failed to explain stability of atoms.*

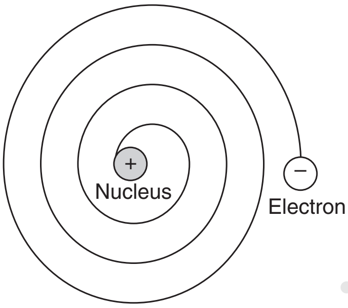


Fig. 3.8. Gradual decrease in the radius of orbit

3.1.6. Discovery of the Neutron

The existence of neutrons in the nucleus was first predicted by **Rutherford** in 1920 to account for the difference in the mass of the atom and the total mass of protons. Neutrons were discovered experimentally by **James Chadwick** in 1932. He bombarded a thin foil of beryllium with fast moving α-particles and observed that highly penetrating rays consisting of neutral particles were produced. These neutral particles were found to have mass 1.675×10^{-27} kg and were named *neutrons*.

In general, a neutron is represented as '***n***'.

> A ***neutron*** is a subatomic particle carrying no charge and having mass 1.675×10^{-27} kg which is almost equal to that of a hydrogen atom.

The mass of the atom is largely due to protons and neutrons in the nucleus of the atom. *Atoms of all elements except ordinary hydrogen contain one or more neutrons in their nuclei.*

The relative masses and relative charges of these three subatomic particles are summarized in Table 3.1.

Table 3.1. The Relative Masses and Relative Charges of Three Fundamental Particles

Particle	**Relative Mass**	**Mass in kg**	**Relative Charge**	**Charge**
Proton (p^+)	1	1.673×10^{-27} kg	+ 1 unit	$+ 1.602 \times 10^{-19}$ C
Neutron (n)	1	1.675×10^{-27} kg	0	0
Electron (e^-)	$\frac{1}{1840}$	9.110×10^{-31} kg	– 1 unit	$- 1.602 \times 10^{-19}$ C

3.1.7. Bohr's Model of Hydrogen Atom

In order to overcome the shortcomings of Rutherford's model, Neils Bohr (1913) proposed a new model of hydrogen atom based on radically new concepts. Bohr made a bold suggestion that particles at atomic level behave differently from the macroscopic objects. He proposed that at the atomic level, electron could revolve around the nucleus in stable orbits without continuously radiating energy in the form of electromagnetic radiations. Main postulates of this model are:

1. *The electrons in the hydrogen atom revolve around the nucleus only in certain selected circular orbits. These orbits are associated with definite energies and are called* ***energy shells*** *or* ***energy levels****. These are numbered as 1, 2, 3, 4, ... ,etc., or designated as K, L, M, N, ..., etc. shells (Fig. 3.9).*

   The energy of the electron is minimum in the orbit nearest to the nucleus i.e., K shell. The energy of the electron increases as it moves away from the nucleus (Fig. 3.9).

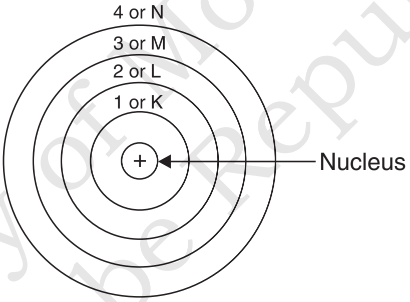


Fig. 3.9. Bohr's orbits

2. *As long as the electron remains in a particular orbit, it does not lose or gain energy. This means that energy of an electron in a particular orbit remains constant. That is why, these orbits are also called* ***stationary states****.*
3. *When energy from some external source is supplied to the electron, it may jump to some higher energy level by absorbing a definite amount of energy (equal to the difference in energy between the two energy levels). When the electron jumps back to the lower energy level it radiates the difference in energy in the form of electromagnetic radiation (Fig. 3.10).*

   The wavelength (λ) of the radiation emitted depends upon the energies of the two levels between which the transition is taking place.

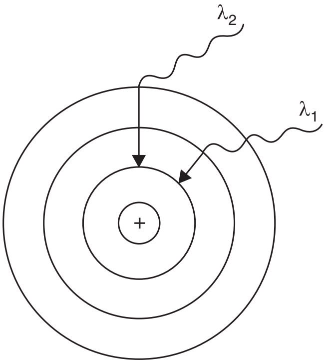


Fig. 3.10.

3.1.7.1. Success of Bohr's Model

Niels Bohr
(1885–1962)

Bohr's model could explain the stability of an atom. According to Bohr's model, an electron revolving in a particular orbit cannot lose energy. The electron can lose energy only if it jumps to some lower energy level. If no lower energy level is vacant then electron will keep on revolving in the same orbit without losing energy and hence it explains the stability of atom.

3.2. FUNDAMENTAL PARTICLES OF AN ATOM

Atoms are made from three fundamental (or sub-atomic)) particles. These are include electrons, protons and neutrons. The atom contain nucleus at its center, which has positively charged protons and neutrons. Electrons are revolving around the nucleus and they carry negative charge.

3.2.1. Arrangement of Particles in an Atom

The fundamental particles are arranged in an atom as follows:

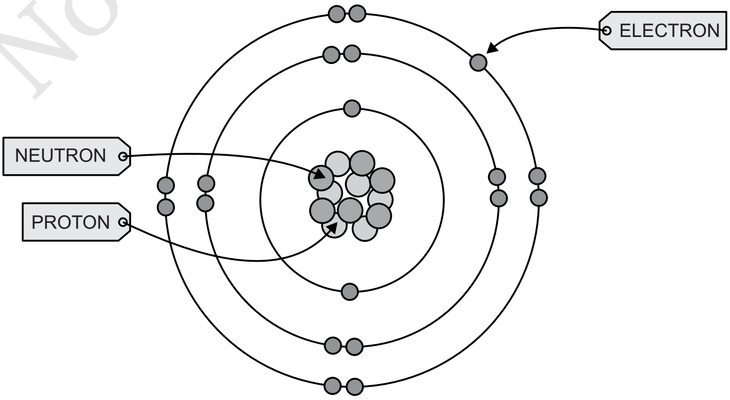


Fig. 3.11. Arrangement of particles in atom

Neutrons and protons are present in the nucleus while electrons revolves around the nucleus.

ACTIVITY 3.1

Aim: ***To Make Model of an Atom***

Materials Required: Some ping-pong balls of different colours (one colour for protons, one for the neutron and other for electrons), glue, a piece of paper and a cardboard.

Procedure:

1. Take a piece of paper and paste it on the cardboard.
2. Draw 4-5 concentric circle on the paper.
3. Now, glue the balls together to represent protons and neutrons. This is the nucleus.
4. Glue the nucleus to the center of the cardboard.
5. Starting from the innermost ring, you can have up to two electrons in the first ring, up to eight in the second ring, up to 18 in the third ring and up to 32 in the fourth ring.
6. Now, glue the electrons on the rings, spacing them evenly. Make sure you don't exceed each ring's maximum number of electrons.

3.3. ISOTOPES

All the atoms of a particular element have same number of protons in their nuclei, however, the number of neutrons may be different. Such atoms have same atomic number but different mass numbers and are known as **isotopes** of the element. Thus:

Isotopes *of an element are the atoms of the element with the same atomic number but different mass numbers.*

For example, hydrogen has three isotopes, protium (H), deuterium (D) and tritium (T). All the three isotopes have atomic number 1, however, their mass numbers are 1, 2 and 3 respectively. The isotopes of other elements do not have special names; they are indicated by giving mass number value on the symbol.

Mass number and atomic number of an atom are generally indicated as follows:

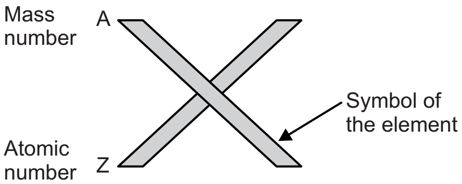


Fig. 3.12. Representation of an isotope

Thus, three isotopes of Hydrogen can be represented as:

$$ {}^{1}_{1}\mathrm{H} \qquad {}^{2}_{1}\mathrm{H} \qquad {}^{3}_{1}\mathrm{H} $$

Protium **Deuterium** **Tritium**

As isotopes are the atoms with the same atomic number but different mass numbers, then we need to know about atomic number and mass number of the atom.

The two isotopes of chlorine are represented as:

$$ {}^{35}_{17}\mathrm{Cl} \qquad {}^{37}_{17}\mathrm{Cl} $$

Chlorine-35 **Chlorine-37**

Another way to describe an isotope is to cite its elemental name and mass number. The two isotopes of chlorine may be represented as chlorine-35 and chlorine-37. Similarly, carbon exists in the form of three isotopes C-12, C-13 and C-14.

3.3.1. Atomic Number

Nucleus carries positive charge due to the presence of protons in it. In 1913, H.G.J. Moseley, a young British physicist, discovered a way to determine nuclear charge accurately. From the charge on the nucleus, the number of protons in it could be determined.

The atomic number of an element is equal to the number of protons in the nucleus of its atom. The atomic number is also known as **proton number** of the element. Further, in an atom, number of protons is equal to the **number of electrons**. Hence, atomic number is also equal to the number of electrons in an atom of the element. Thus,

Atomic number of an element is equal to the number of protons in the nucleus of its atom or the number of extra-nuclear electrons.

Atomic number is denoted by the letter Z.

Atomic Number (Z) = Number of protons = Number of electrons.

For example, an atom of magnesium contains 12 protons, therefore, its atomic numberis 12.

All the atoms of a particular element contain same number of protons in their nuclei.

Therefore, all the atoms of an element have same atomic number. Atoms of different elements contain different number of protons and hence no two elements can have same atomic number. Thus, *each* *element has its characteristic atomic number which can be used to identify the element.*

For example, when we say an element with atomic number 13, we are referring to aluminium. No other element has atomic number 13.

3.3.2. Mass Number

It has already been stated that mass of an atom is mainly concentrated in the nucleus. In the nucleus, there are protons and neutrons. From this it follows that mass of an atom is mainly due to protons and neutrons. Protons and neutrons are collectively called **nucleons**.

The total number of protons and neutrons in the nucleus is called mass number of the atom.

It is generally represented by the letter A.

Mass Number (A) = Number of protons + Number of neutrons
= Number of nucleons

The number of protons in the nucleus is equal to atomic number, Z while *the number of neutrons in the nucleus is sometimes called* ***neutron number***; N.

3.3.3. Calculation of Number of Electrons, Protons and Neutrons

From the knowledge of atomic number and mass number of an element it is possible to calculate number of electrons, protons and neutrons in an atom of the element. For example, atomic number and mass number of aluminium are 13 and 27 respectively.

Number of electrons, protons and neutrons in an atom of it can be calculated as under:

Number of protons = Atomic number = 13
Number of electrons = Atomic number = 13
Number of neutrons = Mass number – Atomic number
= 27 – 13 = 14.

3.4. RELATIVE ATOMIC MASS

Mass of an isotope, relative to the mass of an atom of carbon (C-12 isotope) taken as 12 amu or 12 u, is called the **relative isotopic mass**. The relative masses of isotopes and their relative abundance can be determined using a mass spectrometer.

The relative atomic mass of an element can be calculated from the relative masses and relative abundances of its various isotopes. ***Relative***

***atomic mass** of an element is the weighted average of relative masses of its isotopes*. For example, hydrogen exists in the form of three isotopes; protium, deuterium and tritium with relative masses 1.0078, 2.0141 and 3.0161 u respectively. The relative abundance of protium is 99.985%, deuterium is 0.015% and tritium is almost zero. The average atomic mass of hydrogen may be calculated as:

$$\text{Average atomic mass of hydrogen} = \frac{1.0078 \times 99.985 + 2.0141 \times 0.015}{100}$$

$$= \mathbf{1.00794\ u}$$

It may be mentioned here that ***mass number is different from relative atomic mass***. Mass number is always a whole number (as it corresponds to number of protons and neutrons in the nucleus of the isotope) whereas relative atomic mass is generally fractional because it is the weighted average of relative masses of its various isotopes.

Relative atomic mass (RAM) of an element may be defined as the average relative mass of all the isotopes of the element as compared with mass of an atom of carbon (C-12 isotope) taken as 12 u.

Example 3.1: *The nucleus of an atom of a certain isotope contains 6 protons and 7 neutrons. What is the atomic number and the mass number of this isotope? To which element does it belong?*

Solution:

We know that:

Atomic number = Number of protons

∴ Atomic number = 6

Mass number = Number of protons + Number of neutrons

= 6 + 7 = 13.

Since atomic number of the given isotope is 6, it indicates that this isotope is of **carbon** element.

Example 3.2: *The number of protons in the nucleus of an atom of mass number 97 is 41. Find out the number of neutrons in its isotope of mass number 99?*

Solution:

The atomic number of isotopes is same. Therefore, the number of protons in both the isotopes is same and is equal to 41.

Mass number = Number of protons + Number of neutrons

∴ Number of neutrons = Mass number – Number of protons

= 99 – 41 = 58.

Example 3.3: *Boron exists in nature in the form of two isotopes B-10 and B-11 with relative abundance 19.91% and 80.09% respectively. The masses of B-10 and B-11 isotopes as determined by mass spectrometer are 10.0129 and 11.0093 respectively. What is the average atomic mass of boron?*

Solution:

$$\text{\% of B-10} = 19.91\%; \quad \text{\% of B-11} = 80.09\%$$

$$\text{Average atomic mass of boron} = \frac{19.91 \times 10.0129 + 80.09 \times 11.0093}{100}$$

$$= \mathbf{10.811\ u}$$

3.5. QUANTUM NUMBERS

In an atom a large number of electron orbitals are permissible. These orbitals are designated by a set of numbers known as **quantum numbers**. In order to specify energy, size, shape and orientation of the electron orbital three quantum numbers are required. These are, principal quantum number, azimuthal quantum number and magnetic quantum number. These quantum numbers arise as a natural consequence during the solution of the Schrodinger wave equation. In order to designate the electron, an additional quantum number called spin quantum number is needed to specify spin of the electron. Thus, each orbital in an atom is designated by a set of three quantum numbers and each electron is designated by a set of four quantum numbers. These quantum numbers are discussed below.

3.5.1. The Principal Quantum Number (n)

This is the most important quantum number as it determines to a large extent the energy of an electron. It also determines the average distance of an electron from the nucleus. It is denoted by the letter *n*. This quantum number tells us in which principal energy level or shell the electron is present. It can have any whole number value such as 1, 2, 3, 4, ..., etc. The energy levels or energy shells corresponding to these numbers are designated as K, L, M, N, ..., etc. As the value of *n* increases, the electron gets farther away from the nucleus and its energy increases. The higher the value of *n*, the higher is the electronic energy. For hydrogen and hydrogen-like species, the energy and size of the orbital are determined by principal quantum number alone.

Energy of the electron in a hydrogen atom is related to principal quantum number by the following relation:

$$E_n = -\,k^2\,\frac{\text{üπü}^2\text{ü}me^4}{n^2h^2} = -\,\frac{\times\quad{}^{-18}}{n^2}\,\text{J}$$

where,
m = mass of electron
e = charge on electron
h = Planck's constant
E_n = Energy of the electron in nth principal shell
n = Principal quantum number used to designate the principal shell
k = Coulomb's law constant.

This relation is similar to the expression given by Bohr.

3.5.2. The Orbital Angular Momentum Quantum Number or Azimuthal Quantum Number (l)

This quantum number determines angular momentum of the electron. This is denoted by l. The value of l gives the **sub-level** or **sub-shell** in which the electron is located. *It also determines the shape of the orbital* in which the electron is located. The number of sub-shells within a principal shell is determined by the value of n for that principal energy level. *Thus, l may have all possible whole number values from 0 to n—1* for each principal energy level. For a given value of n, l can have n values. The various sub-levels are designated as *s, p, d, f* depending upon the value of l, as follows:

Value of l →	0	1	2	3	4	5	6
Designation of sub-shell →	*s**	*p*	*d*	*f*	*g*	*h*	*i*

For n = 1, l can have only one value *i.e.*, 0. It means that an electron in first energy level can be present only in s-sub-shell (l = 0).

So **first energy level** has only **one sub-shell**, *i.e.*, 1*s*.

For n = 2, l can have values 0 and 1. It means that the electron in second principal energy level may be located either in s-sub-shell (l = 0) or *p*-sub-shell (l = 1).

So **second energy level** has only two **sub-shells**, *i.e.*, 2*s* and 2*p*.

For n = 3, possible values of l are 0, 1 and 2. This implies that an electron in third principal energy level may be present either in *s*-sub-shell (l = 0) or *p*-sub-shell (l = 1) or *d*-sub-shell (l = 2).

So **third energy level** has **three sub-shells,** *i.e.*, 3*s*, 3*p* and 3*d*. Similarly, **fourth energy level** (n = 4) can have **four sub-shells** 4*s*, 4*p*, 4*d* and 4*f*.

---

*The names **s, p, d** and **f** are derived from spectroscopic terms sharp, principal, diffuse and fundamental respectively.

The relation between the orbital angular momentum and azimuthal quantum number, l is

Orbital Angular Momentum

$$= \sqrt{l(l+1)}\,\frac{h}{2\pi} = \sqrt{l(l+1)}\;\hbar$$

Azimuthal quantum number is also known as **subsidiary quantum number**.

3.5.3 The Magnetic Quantum Number (m_l)

This quantum number which is denoted by m_l refers to the different orientations of electron cloud in a particular sub-shell. These different orientations are called **orbitals**. The number of orbitals in a particular sub-shell within a principal energy level is given by the number of values allowed to m_l which in turn depends on the values of l. The possible values of m_l range from $+ l$ through 0 to $- l$, thus making a total of $(2l + 1)$ values. Thus, in a subshell, the number of orbitals is equal to $(2l + 1)$.

For l = 0 (i.e., *s*-sub-shell), m_l can have only one value, $m_l = 0$. It means that ***s*-sub-shell** has only **one orbital**.

For l = 1 (i.e., *p*-sub-shell), m_l can have three values, + 1, 0 and – 1. This implies that ***p*-sub-shell** has **three orbitals**.

For l = 2 (i.e., *d*-sub-shell), m_l can have five values, + 2, + 1, 0, – 1, – 2. It means that ***d*-sub-shell** has **five orbitals**.

For l = 3 (i.e., *f*-sub-shell), m_l can have seven values, + 3, + 2, + 1, 0, – 1, – 2, – 3. It means that ***f*-sub-shell** has **seven orbitals**.

The number of orbitals in various types of sub-shells are given below in tabular form:

Sub-shell	*s*	*p*	*d*	*f*	*g*
Value of l	0	1	2	3	4
No. of orbitals ($2l + 1$)	1	3	5	7	9

The relationship between the principal quantum number (n), angular momentum quantum number (l) and magnetic quantum number (m_l) is summed up in Table 3.2.

Table 3.2. Relationship among Values of n, l and m_l

Energy Level	Principal Quantum Numbers (n)	Possible Values of (l)	Designation of Sub-shell	Possible Values of (m_l)	Number of orbitals	
					In a given Sub-shell	In a given Energy Level
K	1	0	1*s*	0	1	1
L	2	0	2*s*	0	1	4
		1	2*p*	+ 1, 0, – 1	3	
M	3	0	3*s*	0	1	9
		1	3*p*	+ 1, 0, – 1	3	
		2	3*d*	+ 2, + 1, 0, – 1, – 2	5	

3.5.4. The Spin Quantum Number (m_s)

This quantum number which is denoted by m_s does not follow from the wave mechanical treatment but arises from the spectral evidence that electron in its motion about the nucleus also rotates or spins about its own axis.

This quantum number determines the orientation of spin angular momentum. Spin angular momentum is quantised and can have two orientations relative to a chosen axis. The spin quantum number can have only two values which are $+\frac{1}{2}$ and $-\frac{1}{2}$. The $+\frac{1}{2}$ value indicates clockwise spin (generally represented by an arrow pointing upwards, i.e., ↑) and the other indicates anti-clockwise spin (generally represented by an arrow pointing downwards i.e., ↓).

Due to its spin, the electron behaves as a tiny magnet. The spin of the electron is responsible for most of the magnetic properties of atoms, molecules or ions. If all the electrons in an atom or molecule are paired, it behaves as a **diamagnetic substance** i.e., it is weakly repelled by the magnetic field. On the otherhand, if atoms or molecules of a substance have one or more unpaired or odd electrons, it behaves as a **paramagnetic substance**, i.e., it is weakly attracted by magnetic field.

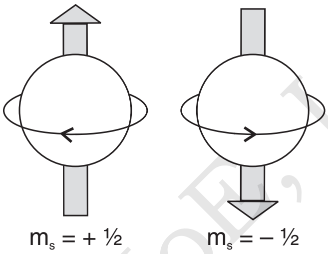


Fig. 3.13. Magnetic field associated with a spinning electron

3.6. ELECTRONIC CONFIGURATION

3.6.1. Dot Notation

A **Lewis electron dot diagram** or **electron dot diagram** is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side.

For example, the Lewis electron dot diagram for calcium is simply

·Ca·

3.6.2. Orbital Notation

Each electron present in an atom is present in a subshell. There are totally four subshells present which are labelled as *s, p, d* and *f*. Each subshell can accommodate only a certain number of electrons, i.e. only 2 electrons can be occupied in *s* subshell, 6 electrons in *p* subshell, 10 electrons in *d* subshell and 14 electrons in *f* subshell.

We can write the electronic configuration of an atom by assigning the number of electrons that are present in the atom as the superscript of the subshell. These electrons are filled in the subshells according to the

Aufbau's principle (each added electron occupies the subshell of lowest energy available). Increasing order of energies of various orbitals are:

$1s > 2s > 2p > 3s > 3p > 4s > 3p > 3d > 4p > 5s > 4d > 5p > 6s > 4f > 5d > 6p > 7s > 5f > 6d > 7p$

Electronic configuration of the first 10 elements are mentioned in Table 3.3 given below.

Table 3.3. Electronic Configuration of the First 10 Elements

No. of Atom	Element	Electronic Configuration
1	Hydrogen (H)	$1s^1$
2	Helium (He)	$1s^2$
3	Lithium (Li)	$1s^2 2s^1$
4	Beryllium (Be)	$1s^2 2s^2$
5	Boron (B)	$1s^2 2s^2 2p^1$
6	Carbon (C)	$1s^2 2s^2 2p^2$
7	Nitrogen (N)	$1s^2 2s^2 2p^3$
8	Oxygen (O)	$1s^2 2s^2 2p^4$
9	Fluorine (F)	$1s^2 2s^2 2p^5$
10	Neon (Ne)	$1s^2 2s^2 2p^6$

3.6.2.1. Electronic Configuration of Transition Elements

The first series of transition elements (At. Nos. = 21 – 30) which follow calcium are scandium (Sc), titanium (Ti), vanadium (V), chromium (Cr), manganese (Mn), iron (Fe), cobalt (Co), nickel (Ni), copper (Cu) and zinc (Zn). In these elements addition of electrons takes place in the 3d-orbitals. All these elements are known as the **transition elements**. The electronic configuration of scandium (At. No. = 21), the first transition element, for example, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$ while that of zinc (At. No. = 30), the last element of the series, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$. Thus, while the filling of $3d$-orbitals begins with scandium, it ends with zinc. The electronic configurations of all these 10 elements are represented in Table 3.4.

Table 3.4. Electronic Configurations of Transition Elements
[Scandium (At. No. 21) to Zinc (At. No. 30)]

Atomic Number	Element	Electronic Configuration			
21	Scandium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^1$	$4s^2$
22	Titanium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^2$	$4s^2$

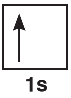
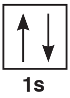
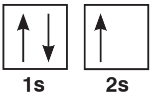
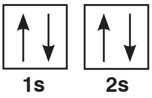
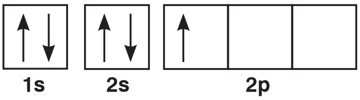
23	Vanadium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^3$	$4s^2$
24	Chromium	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^5$	$4s^1$
25	Manganese	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^5$	$4s^2$
26	Iron	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^6$	$4s^2$
27	Cobalt	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^7$	$4s^2$
28	Nickel	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^8$	$4s^2$
29	Copper	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^{10}$	$4s^1$
30	Zinc	$1s^2$	$2s^2\ 2p^6$	$3s^2\ 3p^6\ 3d^{10}$	$4s^2$

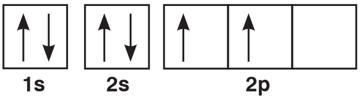
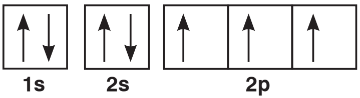
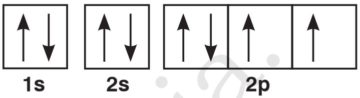
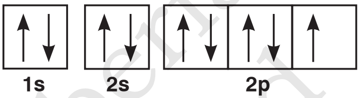
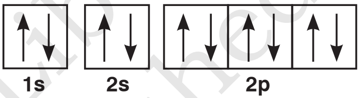
3.6.3. Orbital Diagram

An orbital diagram is a way of representing the electronic configuration of an atom. Electronic configurations can be determined by applying the Aufbau principle (*each added electron occupies the subshell of lowest energy available*), Pauli Exclusion Principle (*no two electrons can have the same set of four quantum numbers*), and Hund's rule of maximum multiplicity (*whenever possible, electrons retain unpaired spins in degenerate orbitals*).

Orbital diagrams (Orbital box diagrams) of the first 10 elements are mentioned in Table 3.5 given below.

Table 3.5. Electronic Configuration of the First 10 Elements using Orbital Diagram

No. of Atom	Element	Electronic Configuration	Orbital Diagram
1	Hydrogen (H)	$1s^1$	
2	Helium (He)	$1s^2$	
3	Lithium (Li)	$1s^2 2s^1$	
4	Beryllium (Be)	$1s^2 2s^2$	
5	Boron (B)	$1s^2 2s^2 2p^1$	

6	Carbon (C)	$1s^2 2s^2 2p^2$	
7	Nitrogen (N)	$1s^2 2s^2 2p^3$	
8	Oxygen (O)	$1s^2 2s^2 2p^4$	
9	Fluorine (F)	$1s^2 2s^2 2p^5$	
10	Neon (Ne)	$1s^2 2s^2 2p^6$	

3.6.4. Noble Gas Notation

It may be noted that configurations of atoms can also be written in condensed form by taking the configurations of noble gases as the core. The configurations of inert gases representing core are written as $[He]^{2}$, $[Ne]^{10}$, $[Ar]^{18}$, $[Kr]^{36}$, $[Xe]^{54}$ and $[Rn]^{86}$. For example, electronic configurations of scandium having atomic number 21 may be written as :

$$_{21}Sc : [Ar]^{18}, 3d^1, 4s^2.$$

Electronic configurations of the first ten elements on this pattern are given in Table 3.6.

Table 3.6. Electronic Configurations of Elements using Noble Gas Notation

Atomic Number	Symbol of Element	Electron Configuration	Atomic Number	Symbol of Element	Electron Configuration
1	H	$1s^1$	11	Na	$[Ne]3s^1$
2	He	$1s^2$	12	Mg	$[Ne]3s^2$
3	Li	$[He]2s^1$	13	Al	$[Ne]3s^2 3p^1$
4	Be	$[He]2s^2$	14	Si	$[Ne]3s^2 3p^2$
5	B	$[He]2s^2 2p^1$	15	P	$[Ne]3s^2 3p^3$
6	C	$[He]2s^2 2p^2$	16	S	$[Ne]3s^2 3p^4$
7	N	$[He]2s^2 2p^3$	17	Cl	$[Ne]3s^2 3p^5$
8	O	$[He]2s^2 2p^4$	18	Ar	$[Ne]3s^2 3p^6$
9	F	$[He]2s^2 2p^5$	19	K	$[Ar]4s^1$
10	Ne	$[He]2s^2 2p^6$	20	Ca	$[Ar]4s^2$

3.6.5. KLMNOPQ Notation

The electrons can be distributed in the KLMN based electron shell. The K shell is the first shell or energy level, L is the second shell, M is third, and so on. The KLMN notations indicate the total number of electrons with each principal quantum number which is ‘n’.

The total number of electrons accommodated by the energy shell is given by $2n^2$, where ‘n’ is the shell number. The values of shell and the principal quantum number is tabulated as:

Shell and ‘n’ value	Max.number of electron
K shell, $n = 1$	$2(1)^2 = 2$
L shell, $n = 2$	$2(2)^2 = 8$
M shell, $n = 3$	$2(3)^2 = 18$
N shell, $n = 4$	$2(4)^2 = 32$

According to **Bohr** and **Bury**:

1. Electrons fill up the lowest energy level first.
2. The maximum number of electrons in any orbit is given by the formula $2n^2$ where n is the number of the orbit, i.e., K = 2, L = 8, M = 18, N = 32, etc.
3. The outermost orbit in a stable atom cannot have more than 8 electrons even if it can accommodate more electrons according to rule 2.
4. The penultimate energy shell, i.e., the energy shell preceding the outermost shell, cannot have more than 18 electrons.

Let us apply the above rules to write electronic configuration of some elements.

I. CASE OF HYDROGEN ATOM

Atomic number of hydrogen is 1. There is only one electron in hydrogen atom which goes into the lowest energy shell i.e., K shell. Thus, the electronic configuration of hydrogen atom is:

$$\begin{array}{ccc} & & \textbf{K} \\ \textbf{H} & : & \textbf{1} \end{array}$$

Atomic diagram of hydrogen atom is shown in Fig. 3.14.

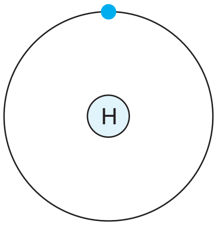


Fig. 3.14. Atomic diagram of hydrogen.

Here, circle in the centre represents the nucleus and the ring drawn around the nucleus represents the first energy shell. The dot in the ring represents the electron.

II. CASE OF OXYGEN ATOM

Now consider an atom of oxygen. Atomic number of oxygen is 8. Thus, there are eight electrons in an atom of oxygen. Two electrons are accommodated in the first shell (K-shell) and the remaining six electrons are accommodated in second shell (L-shell).

	K	**L**
O :	**2**	**6**

The atomic diagram of oxygen is shown in Fig. 3.15.

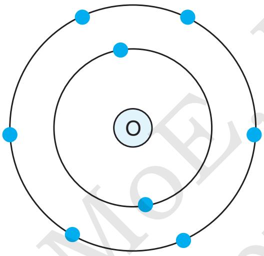


Fig. 3.15. Atomic diagram of oxygen

III. CASE OF CALCIUM ATOM

Atomic number of calcium is 20. An atom of calcium contains 20 electrons. 2 Electrons go in the first shell (K-shell), 8 electrons go to the second shell (L-shell), next 8 electrons go to the third shell (M-shell) and the remaining 2 electrons go to the fourth shell (N-shell). Although the third shell can accommodate a maximum of 18 electrons, all the 10 electrons cannot go into it because that would make 10 electrons in the outermost shell whereas the outermost shell cannot have more than 8 electrons. After the third level acquires 8 electrons, the fourth level begins to fill. The next 2 electrons go to the fourth energy shell (N-shell). Distribution of electrons of various elements has been shown in Table 3.7.

	K	**L**	**M**	**N**
Ca :	**2**	**8**	**8**	**2**

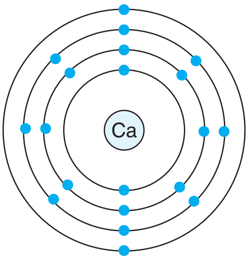
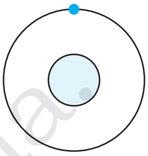
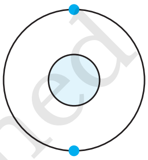
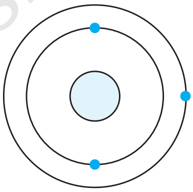
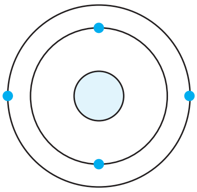
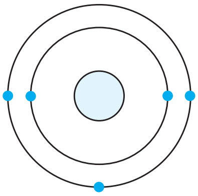
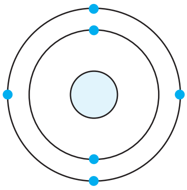
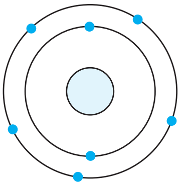
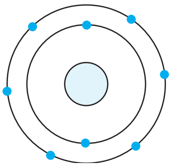
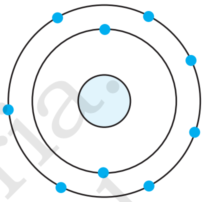
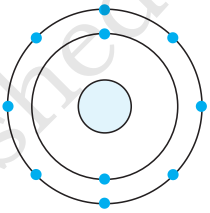
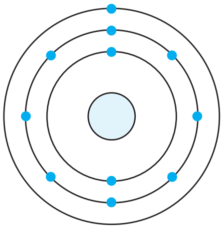
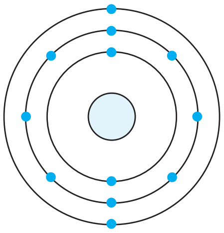
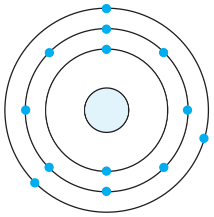
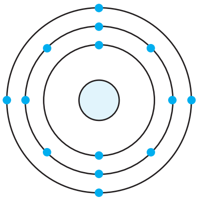
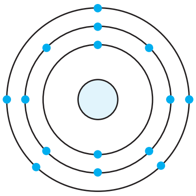
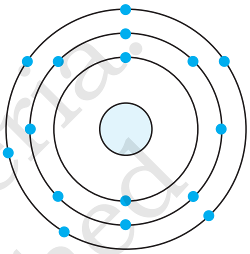
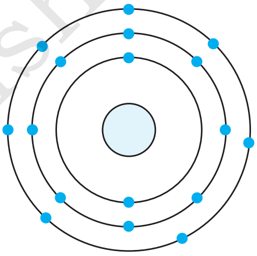
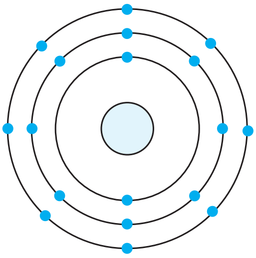
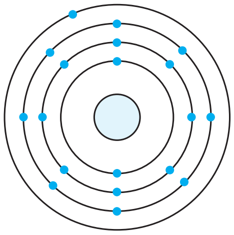
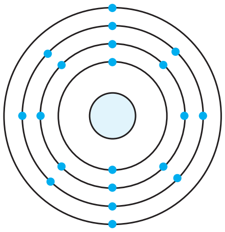


Fig. 3.16. Atomic diagram of calcium

Table 3.7. Distribution of Electrons in Various Energy Shells in the First 20 Elements

Element	Number of Electrons	Electronic Configuration	Atomic Diagram
Hydrogen	1	K 1	
Helium	2	K 2	
Lithium	3	K L 2, 1	
Beryllium	4	K L 2, 2	
Boron	5	K L 2, 3	
Carbon	6	K L 2, 4	
Nitrogen	7	K L 2, 5	

Oxygen	8	K L 2, 6	
Fluorine	9	K L 2, 7	
Neon	10	K L 2, 8	
Sodium	11	K L M 2, 8, 1	
Magnesium	12	K L M 2, 8, 2	
Aluminium	13	K L M 2, 8, 3	
Silicon	14	K L M 2, 8, 4	

Phosphorus	15	K L M 2, 8, 5	
Sulphur	16	K L M 2, 8, 6	
Chlorine	17	K L M 2, 8, 7	
Argon	18	K L M 2, 8, 8	
Potassium	19	K L M N 2, 8, 8, 1	
Calcium	20	K L M N 2, 8, 8, 2	

After the fourth shell acquires 2 electrons, the third shell begins to fill to its capacity of 18. Thus, in the atoms from atomic numbers 21–30, the electrons beyond 20th electron go to the third shell. For example, the electronic configuration of iron (Z = 26) is

	K	L	M	N
Fe :	**2**	**8**	**14**	**2**

3.7. RULES FOR FILLING ORBITALS IN AN ATOM

An atom in its lowest energy state is said to be in the normal state or the *ground state*. The ground state is the most stable state for the atom. The filling of orbitals in the ground state is determined by the following *rules*:

1. AUFBAU RULE

According to Aufbau rule, ***the electrons are added progressively to the various orbitals in their order of increasing energies, starting with the orbital of lowest energy.***

Increasing order of energies of various orbitals is :

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, ...

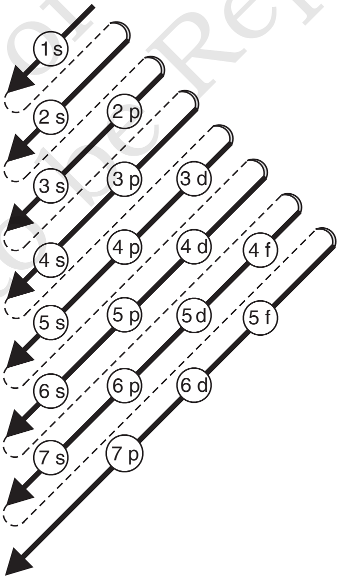


Fig. 3.17. Memory aid for remembering order of energies of various orbitals

Figure 3.17 shows a simple memory aid for remembering the increasing order of energies of various orbitals.

2. PAULI'S EXCLUSION PRINCIPLE

Pauli's exclusion principle states that ***no two electrons in an atom can have same set of all the four quantum numbers***.

From this it follows that **an orbital cannot have more than two electrons**. Moreover, if an orbital has two electrons then **they must have opposite spins**.

3. HUND'S RULE OF MAXIMUM MULTIPLICITY

Hund's rule states that ***the pairing of electrons in the orbitals of a particular sub-shell (p, d or f) does not take place until all the orbitals of the sub-shell are singly occupied. Moreover, the singly occupied orbitals must have the electrons with parallel spins***.

The basis of this rule is that two electrons in a particular orbital feel greater repulsion and hence while filling orbitals of equal energy pairing of electrons is avoided as long as it is possible. Moreover, the singly occupied orbitals should have electrons with parallel spin because this corresponds to state of lower energy. This can be explained in terms of magnetic effects of electron spin. This rule helps us in writing the ground state configurations of those atoms which have partially filled *p*, *d* or *f* sub-shells in them. The application of these rules has been illustrated in the following electronic configurations.

ADDITIONAL ACTIVITIES/EXPERIMENTS

The flame photometer

Flame photometry is one of the branches of atomic absorption spectroscopy. It is also known as flame emission spectroscopy. Currently, it has become a necessary tool in the field of analytical chemistry. Flame photometer can be used to determine the concentration of certain metal ions like sodium, potassium, lithium, calcium and cesium etc. In flame photometer spectra the metal ions are used in the form of atoms. The International Union of Pure and Applied Chemistry (IUPAC) Committee on Spectroscopic Nomenclature has named this technique as flame atomic emission spectrometry (FAES).

In this Experiment, you will calibrate a flame photometer using standard sodium and potassium solutions then measure the Na^+ and K^+concentrations in a redissolved oral rehydration sachet.

EXPERIMENT 3.1

To determine concentration of Na^+ and K^+ in solution by flame photometry. (This activity can be conducted under the supervision of your teacher.)

Reagents: oral rehydration sachet

NaCl standards: 0.25, 0.5, 1.0, 2.0, 4.0 and 5.0 mM

KCl standards: 0.1, 0.2, 0.5, 1.0, 1.5, 2.0 mM

Procedure:

1. Carefully open the oral rehydration sachet and empty the contents into a clean 250 *ml* beaker. Add about 150 *ml* distilled water and gently churn the contents until dissolved.
2. Pour the solution into a 200 *ml* volumetric flask and rinse out the beaker with small amounts of distilled water, adding the washings to the flask. Finally, make up the flask to exactly 200 *ml* and mix thoroughly.
3. Make a 1/50 dilution of the redissolved sachet solution by accurately pipetting 2 *ml* of the solution into a 100 *ml* volumetric flask and making up to 100 *ml* with distilled water.
4. Ensure that the photometer drain is leading into a sink and that the instrument is connected to gas, air and electricity supplies. Ensure the mains supply gas tap is off.
5. Turn the “Sensitivity” and instrument “Gas” controls control fully counterclockwise(towards you).
6. Insert the sodium optical filter.
7. Switch on the instrument and unclamp the galvanometer by turning counterclockwise.
8. Open the mica window, turn on the mains gas supply, light the gas and close the window.

   (**CAUTION:** Do not lean over the instrument or you will set your hair alight.)
9. Turn on the air supply control and adjust the air pressure to 10 lb/in^2. Leave for 1-2 minutes to stabilise.
10. Place a beaker of distilled water into position at the left hand side of the instrument and insert the narrow draw tube into it to allow water to pass through the photometer.

11. Adjust the gas control to give a flame with a large central blue cone then, with water passing through the instrument, slowly close the gas control until ten separate blue cones just form.
12. Set the galvanometer to zero using the “Set zero” control.
13. Replace the distilled water with the 5 mM NaCl standard and adjust the “Sensitivity” control till the galvanometer reads 100.

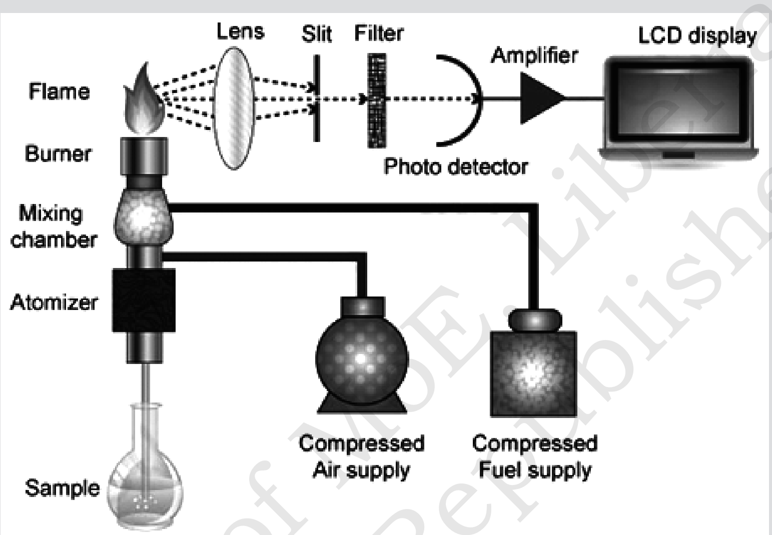


Fig. 3.18. Flame photometer

14. Quickly but carefully, replace the 5 mM NaCl standard with standards of decreasing concentration from 4 mM to 0.25 mM and note the readings in the Table below.
15. Run water through the instrument again for 1-2 min then place the draw tube into a beaker containing the 1 in 50 diluted rehydration sachet solution and note the galvanometer reading.
16. Run water through the instrument again and replace the sodium with the potassium filter.
17. Repeat the above procedure with the KCl standards, setting to 100 with 2.0 mM KCl, then reading the others in reverse order. Then read the 1 in 50 diluted rehydration sachet solution.
18. Finally, run water through the instrument until the flame appears free of colour again.

19. When the instrument is no longer required, switch off in the following sequence:
    (i) Turn off the gas control and the mains gas supply.
    (ii) Wait for the flame to die out.
    (iii) Turn off the air supply.
    (iv) Switch off the electricity.
    (v) Clamp the galvanometer.

$[Na^+]$ (mM)	**5.0**	**4.0**	**2.0**	**1.0**	**0.5**	**0.25**	**0**
Galvo. reading	100						
$[K^+]$ (mM)	**2.0**	**1.5**	**1.0**	**0.5**	**0.2**	**0.1**	**0**
Galvo. reading	100						

20. Plot the galvanometer readings against Na^+ and K^+ concentrations on the graph paper provided (separate graph for each ion) and from these calibration curves determine the Na^+ and K^+ concentrations in the diluted sachet solution. Finally, calculate the Na^+ and K^+ concentrations in the undiluted sachet solution.

	Galvanometer reading	**Diluted concentration (mM)**	**Undiluted concentration (mM)**
Sodium ion			
Potassium ion			

EXPERIMENT 3.2

Aim: ***To compare thermal conductivity of metals and non-metals.***

(*This activity can be conducted under the supervision of your teacher.*)

Materials required: Copper or aluminium wire, carbon or graphite rod, spirit lamp or Bunsen burner, wax, laboratory stand and needles.

Procedure:

1. Take the copper or aluminium wire and fix the wire to a stand. Attach a pin to one of the free ends of the wire using wax, as shown in the figure.

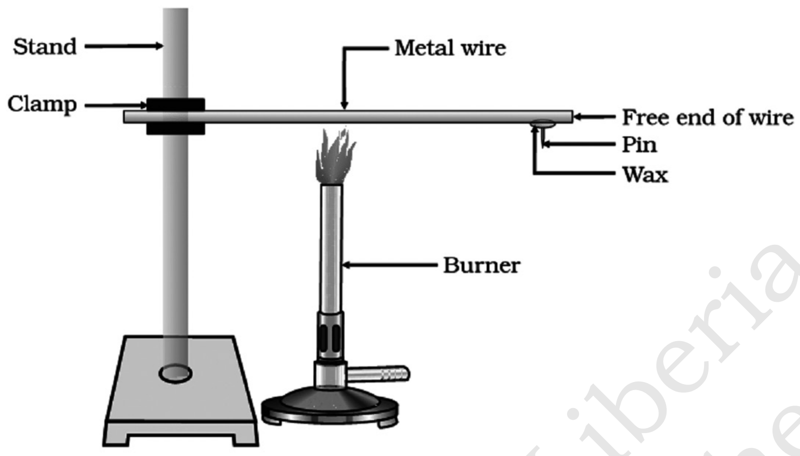


Fig. 3.19. Experimental Set-up for Comparing Thermal Conductivity of Metals and Non-metals

2. Start heating the wire from its other free end using a spirit lamp or a Bunsen burner.
3. Observe whether the wax melts or not and the pin falls or not.
4. Repeat the same procedure with a carbon or graphite rod and record your observations.

Observation:

S.No.	Sample	Observation
1	Copper or aluminium wire	Pin drops
2	Carbon or graphite rod	Pin doesn't drop

Explanation: Metals are a good conductor of heat. Here, once we start heating, heat transfers to the area of wax. It melts the wax. So, the pin drop on the table.

Conclusion: This experiment demonstrates that metals are a good conductor of heat.

GLOSSARY

- **Alpha Particles:** He^{2+} ions or helium nuclei.
- **Atomic Number:** The number of protons present in the nucleus of an atom.
- **Electron:** A sub-atomic particle carrying one unit negative charge and having negligible mass (9.1×10^{-31} kg).

- **Isotopes:** Atom of an element having different mass numbers.
- **Isotopic Mass:** Mass of an isotope relative to the mass of an atom of C-12 taken as 12 u.
- **Mass Number:** The total number of protons and neutrons present in the nucleus an atom.
- **Mass Spectrometer:** An instrument used to find isotopic masses, relative isotopic abundance and relative molecular masses.
- **Neutron:** A neutral sub-atomic particle having mass almost equal to that of a hydrogen atom.
- **Nucleus:** Heavy positively charged centre of an atom. It contains protons and neutrons.
- **Proton:** A sub-atomic particle carrying one unit positive charge and having mass nearly equal to the mass of an atom of hydrogen.
- **Relative Atomic Mass (RAM):** It is the weighted average of isotopic masses of different isotopes of the element.
- **Specific Charge:** e/m ratio for a charged particle.

SUMMARY

- In the fifth century B.C. the Greek Philosopher Democritus proposed that all matter consists of very small indivisible particles called atoms (meaning uncuttable or indivisible).
- The protons are positively charged particles and are present in the nucleus of the atom.
- The electrons are negatively charged particles and are present in the extra-nuclear part of the atom.
- The neutrons are neutral particles and are present in the nucleus along with the protons. Most of the mass of the atom is concentrated in the nucleus.
- Cathode rays consist of negatively charged material particles called electrons.
- Electrons are the fundamental subatomic particles carrying negative charge (-1.602×10^{-19} C) and having mass 9.1×10^{-31} kg. Discovered by J.J. Thomson.
- Charge to mass (e/m) ratio for electrons is 1.76×10^{8} C/g or 1.76×10^{11} C/kg.
- Alpha Particles are He^{2+} ions or helium nuclei.

- Rutherford's Experiment led to the discovery of nucleus. Radius of nucleus (~10^{-14} m) is very small as compared with radius of atom (~10^{-10} m).
- Atomic Number (Z) (Proton Number) the number of protons present in the nucleus of an atom.
- Mass Number (A) (Nucleon Number) the total number of protons and neutrons present in the nucleus of an atom.
- Isotopes are the atoms of an element having different mass numbers.
- Relative Atomic Mass (RAM) of an element is the weighted average of isotopic masses of different isotopes of the element.
- An orbital diagram is a way of representing the electronic configuration of an atom.

EVALUATION

I. Multiple Choice Questions

1. Which of these are negatively charged particles?

(a) Electrons
(b) Protons
(c) Neutrons
(d) None of these

2. Name the subatomic particles that are present in the nucleus of an atom.

(a) Electrons and protons
(b) Electrons and neutrons
(c) Protons and neutrons
(d) None of these

3. Who discovered electrons?

(a) John Dalton
(b) J.J. Thomson
(c) James Chadwick
(d) Ernest Rutherford

4. Who determined the charge to mass ratio of an electron?

(a) Robert Milikan
(b) Robert Boyle
(c) J.J. Thomson
(d) None of these

5. Who discovered nucleus?

(a) John Dalton
(b) J.J. Thomson
(c) James Chadwick
(d) Ernest Rutherford

6. Who discovered proton?

(a) J.J. Thomson
(b) James Chadwick
(c) E. Goldstein
(d) Ernest Rutherford

7. This is an isotope of hydrogen having mass number 2.

(a) Protium (b) Deuterium

(c) Tritium (d) None of these

8. The total number of protons and neutrons in the nucleus is called _______ .

(a) atomic number (b) mass number

(c) Both (a) and (b) (d) None of these

II. State True or False

1. Protons are negatively charged particles.

2. Dalton's atomic theory explains the law of chemical combination by mass.

3. Mass of an electron is nearly equal to $\frac{1}{1840}$th of mass of an atom of hydrogen.

4. An atom as a whole is electrically neutral.

5. Neutrons are experimentally discovered by James Chadwick.

6. Atomic number of an element is equal to the number of neutron in the nucleus of its atom.

7. Mass number is different from relative atomic mass.

III. Answer the Following Questions

1. Write down the brief history of atomic chemistry.

2. Write down the main points of Dalton's atomic theory.

3. Briefly explain the discovery of electrons.

4. Briefly explain the discovery of nucleus.

5. Briefly explain the Bohr's model of hydrogen atom.

6. What do you mean by isotopes? Give two examples of isotopes.

7. What do you mean by quantum number? What are the four kinds of quantum numbers?

8. Explain the rules and principles for filling in electrons.

SEMESTER–I (Period–III)

C10CH4

4
TOPIC

The Periodic Table/Periodic Chemistry

Learning Objectives

Upon completion of this topic, learners will be able to:

- Discuss the history and development of the periodic table
- Identify that elements are placed on the periodic table due to similar properties
- Identify the main blocks, groups and the periods of the periodic table
- Discuss the chemical and physical properties of the groups
- Discuss the periodic trends.

Introduction

Before the beginning of eighteenth century, only about 30 elements were known and it was quite easy to study and remember their individual properties. However, the situation became difficult with the discovery of large number of elements in the later years. At this stage, the scientists felt the need of some simple method to facilitate the study of the properties of various elements and their compounds. After numerous attempts the scientists were ultimately successful in arranging the elements in such a way so that similar elements were grouped together and different elements were separated. This arrangement of elements is known as *classification of elements* and it led to the formulation of *periodic table*. Thus, ***periodic table*** *may be defined as the table giving the arrangement of all the known elements according to their properties so that similar elements fall within the same vertical column and dissimilar elements are separated.*

4.1. HISTORY AND ORIGIN OF THE PERIODIC LAW

Earlier attempts on classification of elements were based on atomic masses. The formulation of a satisfactory periodic law took place only after 1860.

4.1.1. Dobereiner's Triads

The German chemist, Johann Dobereiner (1829) made the first significant attempt to show a relationship between atomic masses and the chemical properties of the elements. He observed that certain similar elements exist in groups of three elements which he named **triads**. An interesting feature of these triads was that the atomic mass of middle member was the arithmetic mean of the atomic masses of the other two members of the triad. For example, lithium, sodium and potassium constituted one triad. Atomic masses of lithium, sodium and potassium are 7, 23 and 39. We can observe that atomic mass of sodium is equal to the average of atomic masses of lithium and potassium.

$$\text{Atomic mass of sodium} = \frac{\text{Atomic mass of lithium} + \text{Atomic mass of potassium}}{2}$$

$$= \frac{7+39}{2} = 23$$

Sulphur, selenium, tellurium and chlorine, bromine, iodine are two more examples of triads. This classification of elements into triads was not satisfactory as it could be applied only to a limited number of elements.

4.1.2. Newland's Law of Octaves

The distinction of correlating the chemical properties of the elements with their atomic masses goes to J.A.R. Newlands. In 1866, he arranged the elements in the order of increasing atomic masses and noted a striking similarity between every eighth element. Newlands named this generalization, the *law of octaves*, due to its similarity to the musical scale. He stated that:

If the elements are arranged in the order of increasing atomic masses, the eighth element, starting from a given one is a kind of repetition of the first—like the eighth note in an octave of music.

Table 4.1 shows the first three rows of Newlands' table.

Table 4.1. Newlands' Arrangement of Elements

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

Newlands' arrangement worked well for the first 17 elements but failed beyond calcium. Moreover, with the discovery of noble gases, the eighth element no longer remained a similar element.

4.1.3. Lothar Meyer's Curves

In 1870, the German chemist Julius Lothar Meyer plotted the atomic volumes (i.e., the atomic mass divided by density) of the elements against their atomic masses. From his graph, Lothar Meyer was able to produce a table showing periodic arrangement of elements. A graph of atomic volume against atomic number has been shown in Fig. 4.1. This graph is similar to the plot of atomic volume versus atomic mass.

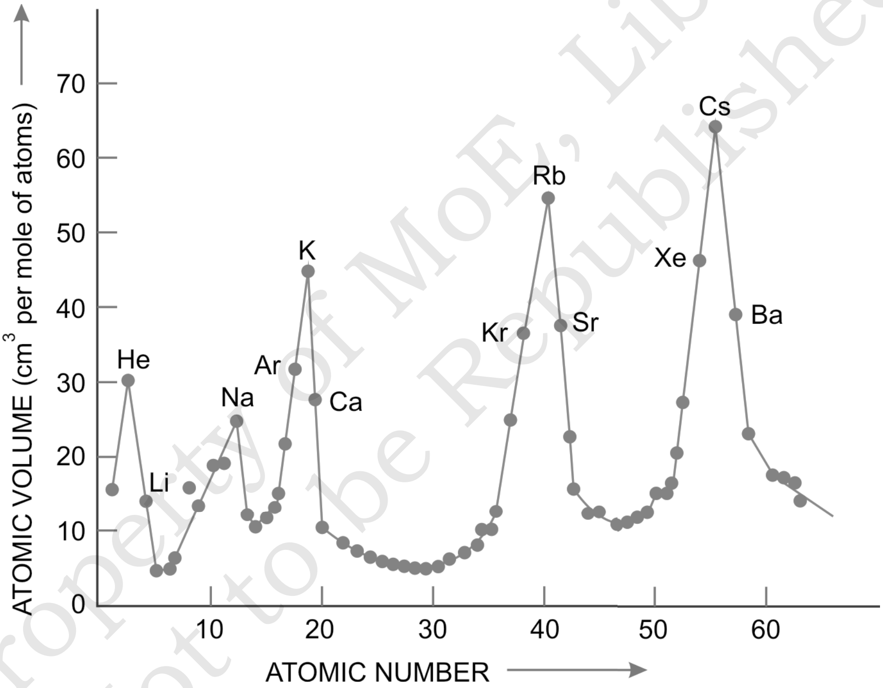


Fig. 4.1. Change of atomic volume with atomic number

It may be noted that the elements belonging to same chemical family occur at similar points on the curves. For example, alkali metals appear on the peaks whereas alkaline earth metals are on the descending portions of the curve while noble gases are on the ascending portions of the curve.

Although Lothar Meyer's curves showed a periodic repetition of properties with atomic masses, most of the credit for arranging the elements in a periodic table is given to Mendeleev.

4.1.4. Mendeleev's Periodic Table

Dmitri Ivanovich Mendeleev (1834–1907)

Dmitri Ivanovich Mendeleev produced a form of periodic table from which the modern periodic table was developed. When Mendeleev started his work 63 elements were known. He arranged all the elements known to him in the order of increasing atomic mass (then known as atomic weight) and showed that elements with similar properties recurred at regular intervals. The elements were arranged in such a way that the elements with similar properties fall in the same vertical column. **These vertical columns of similar elements are called groups and the horizontal rows of elements are called periods**. Mendeleev's periodic table has been shown in Table 4.2.

Table 4.2. Mendeleev's Periodic Table

Group	**I**		**II**		**III**		**IV**		**V**		**VI**		**VII**		**VIII**
Oxide: **Hydride:**	R_2O RH		RO RH_2		R_2O_3 RH_3		RO_2 RH_4		R_2O_5 RH_3		RO_3 RH_2		R_2O_7 RH		RO_4
	A	**B**	**A**	**B**	**A**	**B**	**A**	**B**	**A**	**B**	**A**	**B**	**A**	**B**	
Period 1	H														
Period 2	Li		Be		B		C		N		O		F		
Period 3	Na		Mg		Al		Si		P		S		Cl		
Period 4	K	Cu	Ca	Zn	*	*	*	Ti	As	V	Se	Cr	Br	Mn	Fe Co Ni
Period 5	Rb	Ag	Sr	Cd	In	Y	Sn	Zr	Sb	Nb	Te	Mo	I	Tc	Ru Rh Pd
Period 6	Cs	Au	Ba	Hg	Ti	La	Pb	Hf	Bi	Ta		W			Os Ir Pt

This arrangement of elements was based on the physical and chemical properties of the elements and also the formulae of the compounds they formed with oxygen and hydrogen. Mendeleev selected oxygen and hydrogen as the elements because they are very reactive and formed compounds with most of the elements. The formulae of the hydrides and oxides formed by an element were treated as one of the basic properties for its classification. For example, all the metals of group IA form hydrides with molecular formula MH.

4.1.5. Modern Periodic Law

Mendeleev's periodic table had a number of drawbacks. A large number of scientists made attempts to remove these. In 1914, Henry Moseley, the English physicist, observed that physical and chemical properties of the elements are determined by their atomic numbers instead of their atomic masses. This observation led to the development of **modern periodic law**. The modern periodic law states that:

"The physical and chemical properties of the elements are the periodic function of their atomic numbers."

It means that if the elements are arranged in order of increasing atomic numbers, the elements with similar properties recur after regular intervals. Many new forms of periodic table have been proposed in recent times with modern periodic law as the guiding principle, but the general plan of the table remained the same as proposed by Mendeleev.

Based on this modern periodic law, the **modern periodic table** (Fig. 4.2 on page 94) was prepared. In this periodic table the elements are arranged into different groups and periods based on the electronic configuration.

4.2. STRUCTURE OF THE PERIODIC TABLE

Let us now, study structure of the periodic table.

4.2.1. Groups and Periods

4.2.1.1. Groups

*A vertical column of elements in the periodic table is called a **group**.* A group consists of a series of elements having similar valence shell configuration. There are 18 vertical columns in the long form of periodic table. Thus, there are 18 groups in the long form of periodic table which are numbered from 1 to 18 according to recommendations of IUPAC.

Earlier, the designation of these groups was the same as in the Mendeleev's periodic table. The relationship between the two ways of numbering the groups is given below:

1	2	3	4	5	6	7	8	9	10
IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII		
11	12	13	14	15	16	17	18		
IB	IIB	IIIA	IVA	VA	VIA	VIIA	0		

PERIODS ↓ GROUPS →

This numbering system is used by the international Union of Pure and Applied Chemistry (IUPAC)

12	Atomic number
2 8 2	Electrons in each energy level
Mg	Element symbol
Magnesium 24.305	Average atomic mass

TRANSITION METALS (groups 3–12)

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	1 **H** Hydrogen 1.0079 1																	2 **He** Helium 4.0026 2
2	3 **Li** Lithium 6.941 2 1	4 **Be** Beryllium 9.0122 2 2											5 **B** Boron 10.81 2 3	6 **C** Carbon 12.011 2 4	7 **N** Nitrogen 14.007 2 5	8 **O** Oxygen 15.999 2 6	9 **F** Fluorine 18.998 2 7	10 **Ne** Neon 20.179 2 8
3	11 **Na** Sodium 22.990 2 8 1	12 **Mg** Magnesium 24.305 2 8 2											13 **Al** Aluminium 26.982 2 8 3	14 **Si** Silicon 28.086 2 8 4	15 **P** Phosphorus 30.974 2 8 5	16 **S** Sulphur 32.06 2 8 6	17 **Cl** Chlorine 35.453 2 8 7	18 **Ar** Argon 39.948 2 8 8
4	19 **K** Potassium 39.098 2 8 8 1	20 **Ca** Calcium 40.08 2 8 8 2	21 **Sc** Scandium 44.958 2 8 9 2	22 **Ti** Titanium 47.90 2 8 10 2	23 **V** Vanadium 50.941 2 8 11 2	24 **Cr** Chromium 51.996 2 8 13 1	25 **Mn** Manganese 54.938 2 8 13 2	26 **Fe** Iron 55.847 2 8 14 2	27 **Co** Cobalt 58.933 2 8 15 2	28 **Ni** Nickel 58.71 2 8 16 2	29 **Cu** Copper 63.546 2 8 18 1	30 **Zn** Zinc 65.38 2 8 18 2	31 **Ga** Gallium 69.72 2 8 18 3	32 **Ge** Germanium 72.59 2 8 18 4	33 **As** Arsenic 74.922 2 8 18 5	34 **Se** Selenium 78.96 2 8 18 6	35 **Br** Bromine 79.904 2 8 18 7	36 **Kr** Krypton 83.80 2 8 18 8
5	37 **Rb** Rubidium 85.468 2 8 18 8 1	38 **Sr** Strontium 87.62 2 8 18 8 2	39 **Y** Yttrium 88.906 2 8 18 9 2	40 **Zr** Zirconium 91.22 2 8 18 10 2	41 **Nb** Niobium 92.906 2 8 18 12 1	42 **Mo** Molybdenum 95.94 2 8 18 13 1	43 **Tc** Technetium (97) 2 8 18 14 1	44 **Ru** Ruthenium 101.07 2 8 18 15 1	45 **Rh** Rhodium 102.91 2 8 18 16 1	46 **Pd** Palladium 106.4 2 8 18 18	47 **Ag** Silver 107.87 2 8 18 18 1	48 **Cd** Cadmium 112.41 2 8 18 18 2	49 **In** Indium 114.82 2 8 18 18 3	50 **Sn** Tin 118.69 2 8 18 18 4	51 **Sb** Antimony 121.75 2 8 18 18 5	52 **Te** Tellurium 127.60 2 8 18 18 6	53 **I** Iodine 126.90 2 8 18 18 7	54 **Xe** Xenon 131.30 2 8 18 18 8
6	55 **Cs** Cesium 132.91 2 8 18 18 8 1	56 **Ba** Barium 137.33 2 8 18 18 8 2	57 **La** Lanthanum 138.91 2 8 18 18 9 2	72 **Hf** Hafnium 178.49 2 8 18 32 10 2	73 **Ta** Tantalum 180.95 2 8 18 32 11 2	74 **W** Tungsten 183.85 2 8 18 32 12 2	75 **Re** Rhenium 186.21 2 8 18 32 13 2	76 **Os** Osmium 190.2 2 8 18 32 14 2	77 **Ir** Iridium 192.22 2 8 18 32 15 2	78 **Pt** Platinum 196.09 2 8 18 32 17 1	79 **Au** Gold 196.97 2 8 18 32 18 1	80 **Hg** Mercury 200.59 2 8 18 32 18 2	81 **Tl** Thallium 204.37 2 8 18 32 18 3	82 **Pb** Lead 207.2 2 8 18 32 18 4	83 **Bi** Bismuth 208.98 2 8 18 32 18 5	84 **Po** Polonium (209) 2 8 18 32 18 6	85 **At** Astatine (210) 2 8 18 32 18 7	86 **Rn** Radon (222) 2 8 18 32 18 8
7	87 **Fr** Francium (223) 2 8 18 32 18 8 1	88 **Ra** Radium 226.03 2 8 18 32 18 8 2	89 **Ac** Actinium (227) 2 8 18 32 18 9 2	104 **Rf** Rutherfordium (261) 2 8 18 32 32 10 2	105 **Db** Dubnium (262) 2 8 18 32 32 11 2	106 **Sg** Seaborgium (263) 2 8 18 32 32 12 2	107 **Bh** Bohrium (264) 2 8 18 32 32 13 2	108 **Hs** Hassium (265) 2 8 18 32 32 14 2	109 **Mt** Meitnerium (266) 2 8 18 32 32 15 2	110 **Ds** Darmstadtium (269) 2 8 18 32 32 16 2	111 **Rg** Roentgenium (272)	112 **Cn** Copernicium (285)	113 **Nh** Nihonium (286)	114 **Fl** Flerovium (289)	115 **Mc** Moscovium (290)	116 **Lv** Livermorium (293)	117 **Ts** Tennessine (294)	118 **Og** Oganesson (294)

Lanthanide Series	58 **Ce** Cerium 140.12 2 8 18 20 8 2	59 **Pr** Praseodymium 140.91 2 8 18 21 8 2	60 **Nd** Neodymium 144.24 2 8 18 22 8 2	61 **Pm** Promethium (145) 2 8 18 23 8 2	62 **Sm** Samarium 150.4 2 8 18 24 8 2	63 **Eu** Europium 151.96 2 8 18 25 8 2	64 **Gd** Gadolinium 157.25 2 8 18 25 9 2	65 **Tb** Terbium 158.93 2 8 18 27 8 2	66 **Dy** Dysprosium 162.50 2 8 18 28 8 2	67 **Ho** Holmium 164.93 2 8 18 29 8 2	68 **Er** Erbium 167.26 2 8 18 30 8 2	69 **Tm** Thulium 168.93 2 8 18 31 8 2	70 **Yb** Ytterbium 173.04 2 8 18 32 8 2	71 **Lu** Lutetium 174.97 2 8 18 32 9 2
Actinide Series	90 **Th** Thorium 232.04 2 8 18 32 18 10 2	91 **Pa** Protactinium 231.04 2 8 18 32 20 9 2	92 **U** Uranium 238.03 2 8 18 32 21 9 2	93 **Np** Neptunium 237.06 2 8 18 32 22 9 2	94 **Pu** Plutonium (244) 2 8 18 32 24 8 2	95 **Am** Americium (245) 2 8 18 32 25 8 2	96 **Cm** Curium (247) 2 8 18 32 25 9 2	97 **Bk** Berkelium (247) 2 8 18 32 27 8 2	98 **Cf** Californium (251) 2 8 18 32 28 8 2	99 **Es** Einsteinium (254) 2 8 18 32 29 8 2	100 **Fm** Fermium (257) 2 8 18 32 30 8 2	101 **Md** Mendelevium (258) 2 8 18 32 31 8 2	102 **No** Nobelium (259) 2 8 18 32 32 8 2	103 **Lr** Lawrencium (260) 2 8 18 32 32 9 2

Fig. 4.2. Modern Periodic Table

It may be mentioned here that **all the elements belonging to a particular group have same number of valence electrons and hence exhibit similar properties**. All the elements belonging to the same group constitute a family. For example, elements of group 1 are known as **alkali metals**. Similarly, elements of group-2 are known as **alkaline earth metals**, elements of group 16 are known as **chalcogens**, elements of group 17 are known as **halogens** and elements of group 18 are **noble gases**.

4.2.1.2. Periods

A horizontal row of a periodic table is called a **period**. A period constitutes *a series of elements having same valence shell.* There are **seven periods** in all, which are numbered as 1, 2, 3, 4, 5, 6 and 7.

Each period begins with the filling of new energy shell. In fact, *the number of period also represents the number of the valence shell of elements present in it.*

Different periods can accommodate different number of elements. It may be noted that periods 1, 2 and 3 contain 2, 8, 8 elements respectively and are called **short periods**. There are 18 elements in 4th and 5th periods and they are called the **long periods**. Sixth period containing 32 elements is the **longest period**. Seventh period also contains 32 elements.

Occurrence of 2, 8, 8, 18, 18 and 32 elements in the first, second, third, fourth, fifth, sixth and seventh period can be explained by keeping the following points in mind:

(a) All the elements in the periodic table are arranged in the order of their increasing atomic number.

(b) Each period in the periodic table starts with the beginning of a new energy level. For example, in case of first period the K-shell starts filling while in case of second period L-shell starts filling.

(c) Filling of electron shells takes place according to a well defined set of rules.

The first period begins with the first energy shell. First energy shell can accommodate only 2 electrons in it. **Hence, there are only 2 elements in the first period**.

The second period begins with the filling of second energy level. This shell can accommodate only eight electrons. **Hence, there are only eight elements in the second period**.

After the addition of eight electrons to the second shell, the second shell becomes fully filled. Now, the next electron enters the third shell and the elements now go to third period. The third shell has capacity to accommodate 18 electrons. However, after the third shell receives 8 electrons, the filling of fourth shell begins (because the outermost shell cannot have more than eight electrons). **Therefore, third period has only eight elements**.

With the beginning of the filling of fourth energy shell, the elements start entering the fourth period. Similarly, we can explain why there are 18, 18 and 32 elements in fourth, fifth and sixth period respectively.

There are two rows of elements at the bottom of the periodic table. The first row contains elements with atomic numbers 58 to 71. These 14 elements which follow lanthanum are called **lanthanoids**. The second row contains elements with atomic numbers 90 to 103. These 14 elements which follow actinium are known as **actinoids**.

Different elements belonging to a particular period have different electronic configurations and have different number of valence electrons. That is why, elements belonging to a particular period have different properties.

Example 4.1: *State one reason for keeping fluorine and chlorine in the same group of periodic table.*

Solution: Fluorine and chlorine have similar valence shell electronic configurations. Hence, they have similar properties and are placed in the same group of the periodic table.

$$_{9}\text{F} \;:\; 2,\ 7$$

$$_{17}\text{Cl} \;:\; 2,\ 8,\ 7$$

Both fluorine and chlorine have 7 electrons in the valence shell.

Example 4.2: *Name two other elements which are in the same group as:*

(i) *Carbon* (ii) *Fluorine* (iii) *Sodium, respectively.*

Solution:

(i) The two other elements which are in the same group as carbon are: *silicon (Si)* and *germanium (Ge).*

(ii) The two other elements which are in the same group as fluorine are: *chlorine (Cl)* and *bromine (Br).*

(iii) The two other elements which are in the same group as sodium are: *lithium (Li)* and *potassium (K).*

4.2.2. Metals, Non-metals and Metalloids

In addition to the classification of elements into *s*-, *p*-, *d*- and *f*-blocks, it is possible to divide them into metals, non-metals and metalloids. More than 78% of the elements are metals. **Metals** are present on the left side and the centre of the periodic table. Metals are the elements which are malleable and ductile, possess lustre, are good conductors of heat and electricity and have high densities. Metals usually have high melting and boiling points, and are generally solids at room temperature. Mercury is the only metal which is liquid at room temperature. Gallium (303 K) and caesium (302 K) also have very low melting points.

Non-metals are much less in number than metals. There are only about 20 non-metals. Non-metals are located at the top right hand side of the Periodic Table. Non-metals have low melting and boiling points. They usually are solids or gases at room temperature. Non-metals are neither malleable nor ductile. They are poor conductors of heat and electricity. In a period, the non-metallic character increases as we move from left to right. In a group, the non-metallic character decreases and metallic character increases on going down a group. There is no sharp line dividing metals from non-metals. A zig-zag line separates metals from non-metals as shown in Fig. 4.3. The borderline elements such as silicon, germanium, arsenic, antimony and tellurium exhibit characteristic properties of metals as well as non-metals. These elements are called **semi-metals** or **metalloids**.

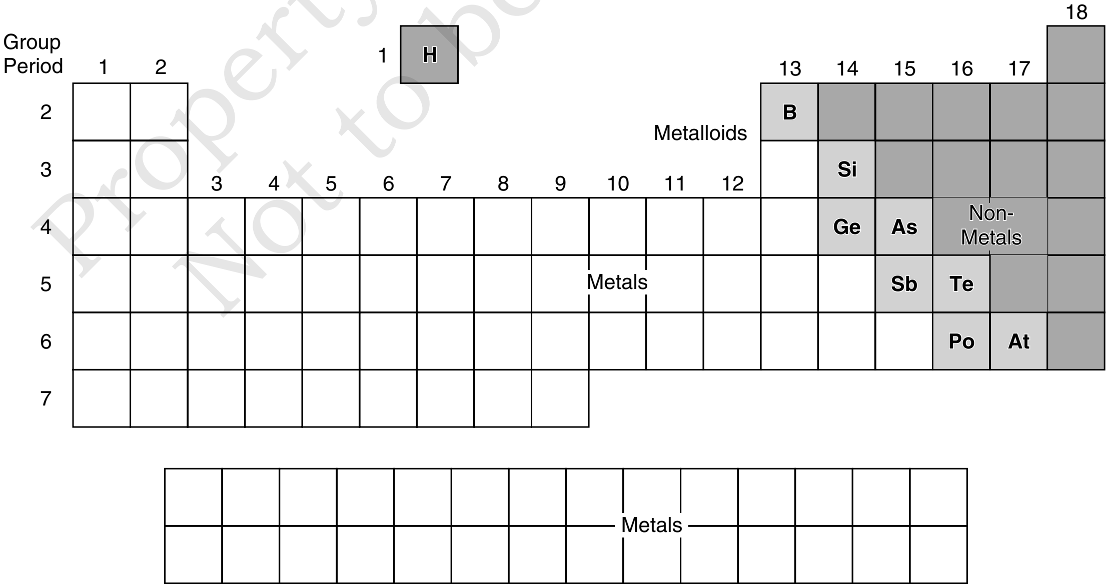


Fig. 4.3. Position of metals, non-metals and metalloids in the periodic table

4.3. TRENDS IN PERIODIC PROPERTIES

Most of the properties of the elements such as electronegativity, ionization energy, electron affinity, atomic radius, metallic character, etc., are directly related to the electronic configuration of the atoms. These properties undergo periodic variation with the change in the atomic number within a period or a group. These properties indirectly control the physical properties such as melting point, boiling point, density, etc. Let us now study the variation of some of the atomic properties in the periodic table.

4.3.1. Electronegativity

Electronegativity *may be defined as the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.*

The main factors on which the electronegativity depends are nuclear charge and atomic radius.

- *Greater the nuclear charge, greater is the electronegativity.*
- *Smaller the atomic radius, greater is the electronegativity.*

4.3.1.1. Variation across a Period

In a period, electronegativity increases in moving form left to right. This is due to the reason that nuclear charge increases whereas atomic radius decreases as we move from left to right in a period.

4.3.1.2. Variation down a Group

In a group, electronegativity decreases on moving down the group. This is due to the effect of increased atomic radius.

4.3.2. Ionization Energy

Ionization energy *may be defined as the amount of energy required to remove the most loosely bound electron from the isolated gaseous atom in its ground state.*

$$\mathbf{A(g) + Energy\ (I.E.) \longrightarrow A^{+}\ (g) + e^{-}}$$

Ionization energy is expressed in terms of *kilo joules per mole of atoms* ($kJ\ mol^{-1}$).

4.3.2.1. Variation in a Period

In general, *the ionization energy increases with the increase in atomic number across the period.* This can be attributed to the fact that moving across the period from left to right,

(i) nuclear charge increases regularly;

(ii) atomic size decreases.

Thus, due to the gradual increase in nuclear charge and simultaneous decrease in atomic size, the attractive force between the nucleus and the electron cloud increases.

4.3.2.2. Variation in a Group

The ionization energies of elements decrease regularly with the increase in atomic number within a group.

The decrease in the value of ionization energy within the group can be explained on the basis of net effect of the following factors:

As we move down the group there is:

(i) a gradual increase in the atomic size.

(ii) increase in the shielding effect on the outermost electron due to increase in the number of inner electrons.

Example 4.3: *From each set, choose the atom which has the largest ionization energy and explain your answer*

(i) *F, O, N* (ii) *Mg, P, Ar.*

Solution:

(i) F has the highest ionization energy among F, O and N because it has smallest size and highest nuclear charge. In general, ionization energy increases as we go from left to right in a period.

(ii) Ar (a noble gas) has the highest ionization energy among the elements Mg, P and Ar because it has stable electronic configuration and maximum nuclear charge.

4.3.3. Electron Affinity

The tendency of a gaseous atom to form anion is expressed in terms of *electron affinity*.

Electron affinity *may be defined as the energy change taking place when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.*

The values of electron affinity are expressed in *kilo joules per mole of atoms*. For example, electron affinity of chlorine is – 348 kJ mol^{-1}.

$$\mathbf{Cl(g) + e^- \longrightarrow Cl^-(g);\ E_{ea} = -348\ kJ\ mol^{-1}}$$

Depending on the element, the process of adding an electron can be either exothermic or endothermic.

The magnitude of electron affinity measures the tightness with which the atom can hold the additional electron. ***The large negative value of electron affinity reflects the greater tendency of an atom to accept the electron***.

4.3.3.1. Factors Affecting Electron Affinity

Following factors affect the electron affinity of an atom:

1. **Nuclear Charge.** Greater the magnitude of nuclear charge greater will be the attraction for the incoming electron and as a result, larger will be the negative value of electron affinity.
2. **Atomic Size.** Larger the size of an atom more will be the distance between the nucleus and the additional electron and smaller will be the negative value of electron affinity.
3. **Electronic Configuration.** Elements having the stable electronic configuration have tendency to accept the electron and larger will be the positive value of its electron affinity.

4.3.3.2. Variation across a Period

On moving across the period, *the atomic size decreases* and *nuclear charge increases*. Both these factors result into greater attraction for the incoming electron, therefore, *electron affinities tend to become more negative as we go from left to right across a period.*

4.3.3.3. Variation down a Group

On moving down a group, *the atomic size* as well as nuclear charge increases. But the effect of increase in atomic size is much more pronounced than that of nuclear charge. Consequently, *electron affinity becomes less negative on going down the group*.

4.3.4. Atomic Radius

The atomic size is very important property of the atoms because it is related to many other chemical and physical properties. In dealing with atomic size, the atom is assumed to be a sphere and its radius determines the size. In general, **atomic radius** is defined as ***the distance of closest approach, to another identical atom***.

4.3.4.1. Variation of Atomic Radii in the Periodic Table

Atomic radii usually depend upon *nuclear charge* and *number of main energy levels* of an atom. The periodic trends in atomic radii have been described as follows:

Variation in a Period

In general, *the atomic radii decrease with the increase in the atomic number in a period.* For example, atomic radii decrease from lithium to fluorine in second period.

The decrease of atomic radii along a period can be explained on the basis of nuclear charge. In moving from left to right across the period, the nuclear charge increases progressively by one unit but the additional electron goes to the same principal shell. As a result, the electron cloud is pulled closer to the nucleus by the increased effective nuclear charge. This causes the decrease of atomic size.

Variation in a Group

In general, *the atomic radii increase from top to bottom within a group of the periodic table.*

In moving down a group, the nuclear charge increases with increase in atomic number but at the same time, there is a progressive increase in the principal energy, shells. The number of electrons in the outermost shell, however, remains the same. Since, the effect of additional energy level is more pronounced than the effect of increased nuclear charge, therefore, *atomic size goes on increasing as we move down a group.*

Atomic raddi increase down the group. Atomic radii decrease across the period.

4.3.5. Electropositivity and Metallic Character

Tendency of atoms of an element to lose electrons and form positive ion is known as ***electropositivity****.*

A more electropositive element has more metallic character.

Whether an element behaves as a metal or a non-metal is directly related to its ionization energy. The elements having low values of ionization energies are metals whereas elements having high values of ionization energies are non-metals. The border line elements behave as metalloids.

4.3.5.1. Variation across a Period

Metallic character decreases across a period from left to right. On the other hand, non-metallic character increases with increase in atomic number across a period.

4.3.5.2. Variation down a Group

On going down a group from top to bottom, the metallic character of elements increases.

In each group, the first element is least metallic while the last element is most metallic.

Metals are located on the left hand side and the centre of the periodic table. Non-metals are located at the top right hand side of the periodic table. A zig-zag line separates metals from non-metals. The borderline elements such silicon, germanium, arsenic, antimony, etc., behave as metalloids. (Fig. 4.4)

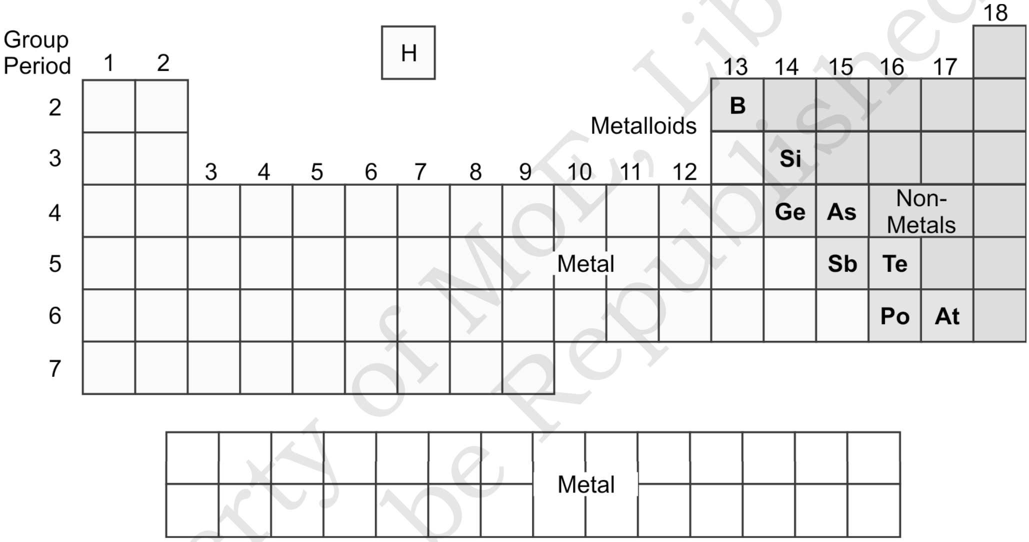


Fig. 4.4. Position of metals, non-metals and metalloids in the periodic table

4.3.6. Ionic to Covalent Bonding in Compounds

Because of the nature of ionic and covalent bonds, the materials produced by those bonds tend to have quite different macroscopic properties. The atoms of covalent materials are bound tightly to each other in stable molecules, but those molecules are generally not very strongly attracted to other molecules in the material. On the other hand, the atoms (ions) in ionic materials show strong attractions to other ions in their vicinity. This generally leads to low melting points for covalent solids, and high melting points for ionic solids. For example, the molecule carbon tetrachloride is a non-polar covalent molecule, CCl_4. It's melting point is –23°C. By contrast, the ionic solid NaCl has a melting point of 800°C.

You can anticipate some things about bonds from the positions of the constituents in the periodic table. Elements from opposite ends of the periodic table will generally form ionic bonds. They will have large differences in electronegativity and will usually form positive and negative ions. The elements with the largest electronegativities are in the upper right of the periodic table, and the elements with the smallest electronegativities are on the bottom left. If these extremes are combined, such as in RbF, the dissociation energy is large. As can be seen from the illustration below, hydrogen is the exception to that rule, forming covalent bonds.

Elements which are close together in electronegativity tend to form covalent bonds and can exist as stable free molecules. Carbon dioxide is a common example.

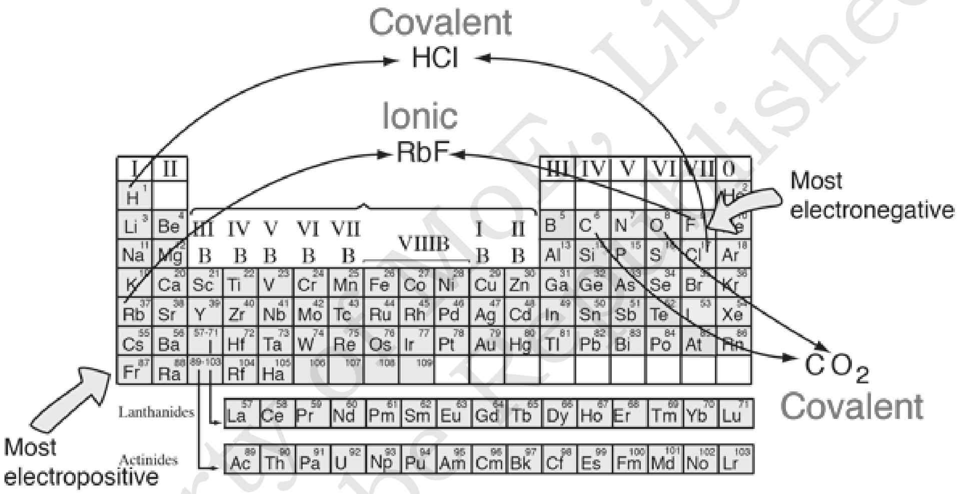


Fig. 4.5.

4.3.7. Non-Metallic Character

Non-metallic elements have strong tendency to gain electrons. Therefore, electronegativity is directly related to the non-metallic character of elements. We can also say that the electronegativity is inversely related to the metallic character of elements.

4.3.7.1. Variation across a Period

In a period, non-metallic character increases from left to right. This is due to the increase in electronegativities across a period.

4.3.7.2. Variation down a Period

In a group, non-metallic character decreases on moving down a group. This is due to the decrease in electronegativities down a group.

4.3.8. Lattice Energy

The lattice energy is the energy change occurring when one mole of a solid ionic compound forms in its gaseous state. It also refers to the energy required to disassociate one mole of a solid compound into its component gaseous ions. Lattice energy can be released (exothermic) or absorbed (endothermic) depending on whether the compound forms or disassociates.

General Reaction during Formation of Solid Ionic Compound:

$$A^+ + B^- \rightarrow AB + \text{Lattice energy}$$

General Reaction during Dissociation of Solid Ionic Compound:

$$AB\,(s) \rightarrow A^+(g) + B^-(g)$$

The following factors affect lattice energy:

1. **Ionic radius:** As the ionic radius increases, the lattice energy decreases. In other words, the bond between opposite ions is strongest when the ions are small. The following table shows the lattice energy values (in kJ/mol) for the ionic bond formed between alkali metals and halogens. It is clear that the bond between Li^+ and F^- (LiF) has the highest lattice energy and that between Cs^+ and I^- (CsI) has the lowest.

	F^-	Cl^-	Br^-	I^-
Li^+	1036	853	807	757
Na^+	923	787	747	704
K^+	821	715	682	649
Rb^+	785	689	660	630
Cs^+	740	659	631	604

2. **Ionic charge:** As the ionic charge increases, the lattice energy increases. In other words, the ionic bond becomes stronger as the charge on the ions becomes large. The lattice energy is proportional to the product of the two ionic charges. The following table shows the lattice energies for salts of OH^- and O^{2-}. It is clear that the bond between Na^+ and OH^- (NaOH) has the smallest lattice energy, and that between Al^{3+} and O^{2-} (Al_2O_3) has the greatest.

	OH^-	O^{2-}
Na^+	900	2481
Mg^{2+}	3006	3791
Al^{3+}	5627	15916

From the above tables, one can observe that the lattice energy increases with atomic charge and decreases with ionic radius. Across a period, the atomic charge increases, and down a group, the ionic radius increases. Hence, the lattice energy increases from left to right across a period and decreases from top to bottom down a group.

4.4. MAIN GROUP ELEMENTS

The main group elements are any of the chemical elements belonging to the *s* and *p* blocks of the periodic table. The *s*-block elements are group 1 (alkali metals) and group 2 (alkaline earth metals). The *p*-block elements are groups 13-18 (basic metals, metalloids, non-metals, halogens, and noble gases). The *s*-block elements usually have one oxidation state (+1 for group 1 and +2 for group 2). The *p*-block elements may have more than one oxidation state, but when this happens, the most common oxidation states are separated by two units. Specific examples of main group elements include helium, lithium, beryllium, boron, carbon, nitrogen, oxygen, and fluorine.

Let us discuss properties of some main group elements.

4.4.1. Helium (He)

Helium is a member of the noble gas family. The noble gases are the elements in Group 18 (VIIIA) of the periodic table. Helium was first discovered in the Sun. In 1868 Pierre Janssen (1824-1907), a French astronomer, studied light from the Sun during a solar eclipse. He found proof that a new element existed in the Sun. He called the element helium.

4.4.1.1. Physical Properties of Helium

Following are some physical properties of helium:

- It is a colourless, odourless, tasteless and inert gas.
- It is completely inert (i.e.,it doesn't react with any other element).
- It has a very low boiling point (4.2 K).
- It is less dense than any other known gas except hydrogen.

4.4.1.2. Chemical Properties of Helium

Following are some chemical properties of helium:

- It is non-toxic.
- It is less soluble in water than any other gas.
- It is not flammable.

4.4.2. Lithium (Li)

Lithium is part of the Group 2 Alkaline Earth Metals. It is found in beryl and emerald, minerals that were known to the ancient Egyptians. It is widely distributed in earth's crust and is estimated to occur in Earth's igneous rocks to the extent of 0.0002 percent.

4.4.2.1. Physical Properties of Lithium

Following are some physical properties of lithium:

- Lithium is a silvery white metal.
- It is the lightest of all solid elements.
- It is a good conductor of heat and electricity.
- Its vapours impart calamine red colour to the flame.
- It gives alloys with number of metals and forms amalgam.

4.4.2.2. Chemical Properties of Lithium

Following are some chemical properties of lithium:

1. **Reaction with air:** Lithium is not affected by dry air but in moist air it is readily oxidized. When heated in air above 450K, it burns to give lithium monoxide and lithium nitride.

$$4\ Li + O_2 \rightarrow 2\ Li_2O$$

$$6\ Li + N_2 \rightarrow 2\ Li_3N$$

2. **Reactions with Water:** It decomposes cold water forming lithium hydroxide and hydrogen.

$$2\ Li + 2\ H_2O \rightarrow 2\ LiOH + H_2$$

4.4.3. Beryllium (Be)

Beryllium is part of the Group 2 Alkaline Earth Metals. It is found in nature and is combined with other elements in minerals, including beryl and chrysoberyl. In its purest form, beryllium is a steel-gray and lightweight alkaline earth metal.

4.4.3.1. Physical Properties of Beryllium

Following are some physical properties of beryllium:

- Beryllium is a silvery-white metal.
- It has low density.
- It is a non-magnetic, hard and brittle metal.
- It is a toxic element.

4.4.3.2. Chemical Properties of Beryllium

Following are some chemical properties of beryllium:

1. **Reaction with water:** Beryllium metal does not react with water or steam, even if the metal is heated to red heat.
2. **Reaction with acids:** The surface of beryllium metal is covered with a thin layer of oxide that helps protect the metal from attack by acids, but powdered beryllium metal dissolves readily in dilute acids such as sulphuric acid (H_2SO_4), hydrochloric acid (HCl), or nitric acid (NO_3), to form solutions containing the aquated Be(II) ion together with hydrogen gas (H_2).

$$Be(s) + H_2SO_4(aq) \rightarrow Be_2+(aq) + SO_4^{2-}(aq) + H_2(g)$$

4.4.4. Boron (B)

Boron is part of the Group 13 of the periodic table. Boron is not present in nature in elemental form. It is found combined in borax, boric acid, kernite, ulexite, colemanite and borates. Vulcanic spring waters sometime contains boric acids.

5
B
Boron
10.811

4.4.4.1. Physical Properties of Boron

Following are some physical properties of boron:

- It is a hard and black-coloured non-metallic solid.
- It has an unusually high melting point.
- At room temperature, it is a poor electrical conductor, but it is a good conductor at high temperatures.

4.4.4.2. Chemical Properties of Boron

Following are some chemical properties of boron:

1. **Reaction with air:** Boron do not react with oxygen at room temperature but at a higher temperature it reacts to form boron trioxide (B_2O_3).

$$4\ B(s) + 3\ O_2(g) \rightarrow 2\ B_2O_3(s)$$

2. **Reaction with water:** Boron does not react with water under normal conditions.
3. **Reaction with halogens:** When boron undergoes halogenation, the product formed is boron trihalides. The reaction with bromine (Br) is given below.

$$2B\ (s) + 3Cl_2(g) \rightarrow 2BCl_3(l)$$

4.4.5. Carbon (C)

6
C
Carbon
12.011

Carbon is part of the Group 14 of the periodic table. It is the seventeenth most abundant element found on earth. It is found in the minerals of most metals in the form of carbonates.

4.4.5.1. Physical Properties of Carbon

Following are some physical properties of carbon:

- Carbon is a unique element. It occurs in many forms. Some examples of the pure form of carbon are coal and soot.
- It is soft, dull grey or black non-metal.
- It occurs in a number of allotropic forms such as diamond and graphite.
- It is available in various shapes.

4.4.5.2. Chemical Properties of Carbon

The chemical properties of carbon are observed during the chemical reactions. Carbon takes part in four main reactions:

1. **By heating carbon in limited supply of oxygen:** Carbon monoxide is formed by incomplete combustion of carbon or carbon containing compounds (such as hydrocarbons) in the limited supply of oxygen.

$$C(s) + \tfrac{1}{2}\,O_2(g) \rightarrow \underset{\text{Carbon monoxide}}{CO(g)}$$

2. **Reaction with carbon dioxide:** When carbon reacts with carbon dioxide it produces carbon monoxide.

$$C(s) + CO_2(g) \rightarrow \underset{\text{Carbon monoxide}}{2\ CO\ (g)}$$

3. **Reaction with iron oxide:** Carbon is more reactive than iron, so it can displace iron from iron oxide.

$$2Fe_2O_3 + 3C \rightarrow 4Fe(l) + 3CO_2(g)$$

4. **Reaction with conc. H_2SO_4:** When carbon react with conc. sulphuric acid it produces carbon dioxide, sulphur dioxide and water.

$$C(s) + 2H_2SO_4\ (aq) \rightarrow CO_2\ (g) + 2SO_2\ (g) + 2H_2O(l)$$

4.4.6. Nitrogen (N)

Nitrogen is part of the Group 15 of the periodic table. It is essential to life on Earth. It is a component of all proteins, and it can be found in all living systems. Nitrogen is crucial to life, but in excess it can also be harmful to the environment.

7
N
Nitrogen
14.007

4.4.6.1. *Physical Properties of Nitrogen*

Following are some physical properties of nitrogen:

- It is colourless, tasteless and odourless gas.
- It is non-toxic in nature.
- It is almost insoluble in water (23.2° cm^3 per litre of water at 273 K and 1 bar pressure).
- Its freezing point and boiling point are 63 K and 77.2 K respectively.

4.4.6.2. *Chemical Properties of Nitrogen*

Following are some chemical properties of nitrogen:

1. **Reaction with highly electropositive metals like lithium, calcium magnesium etc.** These metals burn in the atmosphere of dinitrogen of form their respective nitrides

$$6Li + N_2 \rightarrow 2Li_3N$$
$$3Ca + N_2 \rightarrow Ca_3N_2$$
$$3Mg + N_2 \rightarrow Mg_3N_2$$

2. **Reaction with non-metal like dihydrogen and dioxygen:**

$$N_2 + 3H_2 \xrightarrow[\text{750 K pressure}]{\text{Fe/MO}} \underset{\text{Ammonia}}{2NH_3}$$

$$N_2 + O_2 \xrightarrow[\text{elertric arc}]{\text{3000 K}} \underset{\text{Nitric oxide}}{2NO}$$

3. **Reaction with compounds like Al_2O_3 and calcium carbide (CaC_2):**

$$\underset{\text{Alumina}}{Al_2O_3} + N_2 + 3C \xrightarrow{2100\,K} \underset{\text{Aluminium nitride}}{2AlN} + 3CO$$

$$\underset{\text{Calcium carbide}}{CaC_2 + N_2} \xrightarrow{1300\,K} \underset{\text{Calcium cyanamide}}{CaNCN} + \underset{\text{Coke}}{C}$$

Both AlN and $CaCN_2$ hydrolyse with boiling water to give ammonia.

4.4.7. Oxygen (O)

8
O
Oxygen
15.999

Oxygen is part of the Group 16 of the periodic table. It is the third most abundant element on the earth. About 50% of the earth's crust consists of oxygen. In free state it constitutes 21% of atmosphere by volume. In the combined state oxygen is present in water, minerals and the bodies of plants and animals. The human body has about 65% oxygen by weight. It is present in the bodies of animals and plants. Sand contains nearly 56% oxygen.

4.4.7.1. Physical Properties of Oxygen

Following are some physical properties of oxygen:

- Oxygen is a colourless, odourless and tasteless gas.
- Its density is higher than air.
- Oxygen is a very poor conductor of heat and electricity.
- Oxygen is soluble in some liquids such as water, alcohol, etc.

4.4.7.2. Chemical Properties of Oxygen

Oxygen is a highly reactive element. It is easily capable of combining with other elements. The most important chemical property of oxygen is that it supports combustion. It also combines with elements at room temperature, for example, the formation of rust. Decaying is an example of oxygen reacting with compounds. Carbon dioxide and water are the main products of decay.

1. **Reaction of Oxygen with Metals:** Metals react with oxygen to form basic oxides which dissolve in water to form basic oxide (alkalies) and turn red litmus blue. For example,

   Sodium burns brightly in oxygen with a golden yellow flame forming sodium oxide.

$$4Na + O_2 \rightarrow 2Na_2O$$

Sodium oxide thus formed dissolves in water forming NaOH (Sodium hydroxide) which turns red litmus blue.

2. **Reaction of Oxygen with Non-Metals:** Non-metals burn in oxygen to form acidic oxides. The acidic oxides dissolve in water and turn blue litmus red. For example,

   Burning sulphur combined with oxygen forms SO_2 gas which has a pungent and suffocating smell.

   $$S + O_2 \rightarrow SO_2$$

   SO_2 dissolves in water to form sulphurous acid which turns blue litmus red.

   $$SO_2 + H_2O \rightarrow H_2SO_3 \text{ (sulphurous acid)}$$

4.4.8. Fluorine (F)

Fluorine is part of the Group 17 of the periodic table. Fluorine occurs naturally in the crust of the earth where it is present in rocks, coal, and clay. Through wind-blown soil, fluorides are released into the air. Fluorine is the 13th most abundant element in the crust of the Earth.

9
F
Fluorine
18.998

4.4.8.1. Physical Properties of Fluorine

Following are some physical properties of fluorine:

- Fluorine can be found in nature as a gas.
- It is a light gas with a pale yellow colour and a faint smell.
- It is a flammable gas.

4.4.8.2. Chemical Properties of Fluorine

Following are some chemical properties of fluorine:

1. **Reaction of fluorine with hydrogen:** Fluorine reacts quickly with hydrogen to form hydrogen fluoride. The reaction can be explosive under the right conditions.

   $$H_2(g) + F_2(g) \rightarrow 2\ HF(g)$$

2. **Reaction of fluorine with metals/metal ions:** Fluorine reacts with sodium to form sodium fluoride.

   $$2\ Na(s) + F_2(g) \rightarrow 2\ NaF(s)$$

   The reaction with metals is a general reaction for most metals.

   Similarly Tellurium reacts with excess fluorine to form tellurium(VI) fluoride.

   $$Te(s) + 3\ F_2(g) \rightarrow TeF_6(s)$$

3. **Reaction of fluorine with noble gasses:** Krypton will react with fluorine, F_2, when cooled to –196 °C (liquid nitrogen) and zapped with an electric discharge or X-rays, forming krypton(II) fluoride (KrF_2).

$$Kr(s) + F_2(s) \rightarrow KrF_2(s)$$

This compound decomposes when heating to room temperature.

4.5. PERIOD THREE COMPOUNDS

4.5.1. Properties of Period three Compounds

Properties of some period 3 compounds are given below.

1. Hydrides

The hydrides of period 3 are given below:

Hydrides	NaH	MgH_2	AlH_3	SiH_4	PH_3	H_2S	HCl
Bonding	Ionic	Ionic/ Covalent	Covalent/ Polymer	Covalent	Covalent	Covalent	Covalent
Physical State	Solid	Solid	Solid	Gas	Gas	Gas	Gas

The hydrides of Na and Mg are ionic while that of other elements of period 3 are covalent. Aluminium hydride is polymeric $(AlH_3)_n$. Hydrides of Na, Mg and Al are solids while that of other elements of period 3 are gases. Argon does not form hydride.

NaH, MgH_2 and AlH_3 react vigorously with water and yield basic solutions.

$$NaH + H_2O \longrightarrow NaOH + H_2$$

$$MgH_2 + 2H_2O \longrightarrow Mg(OH)_2 + 2H_2$$

$$2AlH_3 + 3H_2O \longrightarrow 2Al(OH)_3 + 3H_2$$

SiH_4, PH_3 and H_2S are not much soluble in water and their aqueous solutions are almost neutral. HCl is highly soluble in water and its aqueous solution is strongly acidic. It neutralizes bases to form salts.

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

2. Oxides

The oxides of third period elements are given below in tabular form:

Oxide	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10} P_4O_6	SO_3 SO_2	Cl_2O_7 Cl_2O
Bonding	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Nature of Oxide	Alkaline	Alkaline	Alkaline	Acidic	Acidic	Acidic	Acidic

The oxides of metallic elements are ionic solids. SiO_2 is a giant covalent network solid. The remaining non-metallic oxides are covalent molecular compounds.

Acid-Base Behaviour of Some Period 3 Oxides

On moving from left to right among the oxides we observe that metal oxides form strongly basic oxides on the left-hand side to strongly acidic ones on the right, via an amphoteric oxide (aluminium oxide) in the middle. An *amphoteric oxide* is one which shows both acidic and basic properties.

Sodium oxide is a simple strongly basic oxide. Sodium oxide reacts exothermically with cold water to produce sodium hydroxide solution.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

As a strong base, sodium oxide also reacts with acids. For example, it would react with dilute hydrochloric acid to produce sodium chloride solution.

$$Na_2O + 2HCl \longrightarrow 2NaCl + H_2O$$

Magnesium oxide is not as strongly basic as sodium oxide.

Magnesium oxide has a slight reaction with the water to produce hydroxide ions in solution.

$$MgO + H_2O \longrightarrow Mg(OH)_2$$

Magnesium oxide reacts with warm dilute hydrochloric acid to give magnesium chloride solution.

$$MgO + 2HCl \longrightarrow MgCl_2 + H_2O$$

3. Hydroxides

The acid-base character of hydroxides of elements of period 3 changes from basic to amphoteric to acidic across the period from left to right.

Sodium hydroxide and magnesium hydroxide are basic.

Both react with acids to form salts. For example, with dilute hydrochloric acid, you get colourless solutions of sodium chloride or magenesium chloride.

$$\mathrm{NaOH + HCl \longrightarrow NaCl + H_2O}$$
$$\mathrm{Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O}$$

Aluminium hydroxide is amphoteric.

Like sodium or magnesium hydroxides, it reacts with acids. It shows the basic nature of aluminium hydroxide.

$$\mathrm{Al(OH)_3 + 3HCl \longrightarrow AlCl_3 + 3H_2O}$$

But aluminium hydroxide also reacts with sodium hydroxide solution

$$\mathrm{Al(OH)_3 + NaOH \longrightarrow NaAl(OH)_4}$$

4. Chlorides

The chlorides of period 3 elements are:

Chlorides	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_5 PCl_3	S_2Cl_2
Bonding	Ionic	Ionic	Ionic covalent	Covalent	Covalent	Covalent

Sodium chloride and magnesium chloride are ionic and consist of giant ionic lattices at room temperature. Aluminium chloride and phosphorus (V) chloride change their structure from ionic to covalent when the solid turns to a liquid or vapour.

The others are simple covalent molecules.

Sodium and magnesium chlorides are solids with high melting and boiling points because of the large amount of heat which is needed to break the strong ionic attractions.

The rest are low melting point solids or liquids.

Reactions with Water

The simple ionic chlorides (sodium and magnesium chloride) dissolve in water. The other chlorides all react with water in different ways to form a variety of products. The reaction with water is known as **hydrolysis**.

Solid aluminium chloride reacts with the water rather than just dissolving in it. In the first instance, hexaaquaaluminium ions are formed together with chloride ions.

$$\mathrm{AlCl_3}(s) + \mathrm{6H_2O}(l) \longrightarrow [\mathrm{Al(H_2O)_6}]^{3+}(aq) + \mathrm{3Cl^-}(aq)$$

Silicon tetrachloride is a colourless liquid at room temperature which fumes in moist air.

It fumes in moist air because it reacts with water in the air to produce hydrogen chloride.

$$\mathrm{SiCl_4 + 2H_2O \longrightarrow SiO_2 + 4HCl}$$

4.5.2. Thermal stability of CO_3^{2-}, NO_{3-} of Li, Na, K, Mg and Ca

The trioxocarbonate(IV) of Group-1 metal are known to be relatively more stable to heat as compared to those of group-2 metals which are less stable.

Carbonates and Hydrogen Carbonates

The trioxocarbonate(IV) of alkali metals except lithium trioxocarbonate(IV) are stable to heat. The trioxocarbonate(IV) of group-2 metals *i.e.*, magnesium and calcium and that of lithium decompose on heating, forming an oxide along with the evolution of carbon dioxide. For example,

$$Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$$

$$MgCO_3 \xrightarrow{\Delta} MgO + CO_2$$

$$Na_2CO_3 \xrightarrow{\Delta} \text{no effect.}$$

Reason. *The stability of carbonate towards heat depends upon the relative stability of the resulting metal oxide. More is the stability of the resulting metal oxide lesser is the stability of the carbonate towards heat and vice versa. Now the stability of resulting metal oxides decreases down the group due to decrease in lattice enthalpy, (because of bigger size) therefore the stability of carbonates towards heat increases.*

Li_2CO_3 decomposes on heating because it gives lithium oxide. The small O^{2-} anion is very strongly attracted to Li^+ ion resulting in an ionic compound with high lattice energy and therefore it is more stable than Li_2CO_3. Greater is the stability of resulting oxide more is its tendency of formation and hence lower is the thermal stability of the carbonate.

In period 3, magnesium trioxocarbonate(IV) decomposes more readily than sodium trioxocarbonate(IV). It is because the resulting MgO is more stable than $MgCO_3$ because of strong ionic bond between Mg^{2+} and O^{2-} ions. Thus its lattice energy is high and hence the stability. The stabilities of carbonates of group-2 metals increase on moving down the group. For example, $BeCO_3$ decomposes at 373 K, $MgCO_3$ at 813 K, $CaCO_3$ at 1173 K, $SrCO_3$ at 1563 K and $BeCO_3$ at 1633 K.

Lithium and group-2 metals do not form solid hydrogencarbonates, although they exist in solution. On heating these solutions, the hydrogencarbonates decompose to form carbonates and CO_2 gas is liberated. The solid hydrogencarbonates of alkali metals decompose between 375 and 575 K.

$$2NaHCO_3(s) \longrightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Nitrates

Trioxonitrate(V) of alkali metals (Na, K, etc.), except $LiNO_3$, decompose on strong heating forming nitrites and oxygen. For example,

$$2KNO_3(s) \xrightarrow{\text{heat}} 2KNO_2(s) + O_2(g)$$

Trioxonitrate(V) of Mg and Ca metals and $LiNO_3$ decompose on heating to form oxides, nitrogen dioxide and oxygen.

$$2LiNO_3(s) \xrightarrow{\Delta} Li_2O(s) + 2NO_2(g) + O_2(g)$$

$$2Ca(NO_3)_2(s) \xrightarrow{\Delta} 2CaO(s) + 4NO_2(g) + O_2(g)$$

Thermal stabilities of trioxonitrate(V) of group-1 *i.e.*, Li, Na, K, Rb, Cs, etc., and group-2 metals Be, Mg, Ca, Sr, Ba etc., increase on moving down the group from top to bottom.

EXPERIMENT 4.1

Objectives:

To demonstrate thermal stabilities of some trioxocarbonated(IV) in the laboratory.

Requirements:

The experiment requires, three test tubes, gas holder, lime water.

Procedure:

1. Take the three test tubes and to each of it add about 1g each of Li_2CO_3, Na_2CO_3 and $CuCO_3$.
2. To the mouth of the test tubes attach gas detector containing lime water. The simple experimental set up is shown in Fig. 4.6.

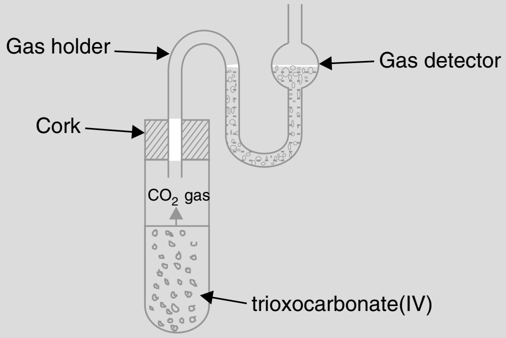


Fig. 4.6.

Observations:

The trioxocarbonate(IV)s of lithium and copper will decompose on heating and will turn lime water milky.

$$Li_2CO_3(s) \xrightarrow{\Delta} Li_2O(s) + CO_2(g)$$

$$CuCO_3(s) \xrightarrow{\Delta} CuO(s) + CO_2(g)$$

The trioxocarbonate of copper decomposes at much slower rate than that of lithium.

Na_2CO_3 will not decompose to produce carbon dioxide gas.

4.6. PERIOD FOUR METALS

Metals in the period 4 are K, Ca, Sc, Ti, etc. In this section, we will discuss about K and Ca.

4.6.1. Potassium

Potassium is the first element in the periodic table's fourth period. The name potassium is derived from the mineral Potash. For hundreds of years, the element has been used. It, along with lithium, rubidium, sodium, caesium, and francium, is an alkali metal. Potassium has an atomic mass of 39.098 atomic mass units. It is represented by the letter 'K.'

Physical Properties of Potassium

Following are some physical properties of potassium:

- Potassium is an alkali metal.
- It is a highly reactive element and does not occur in a free state.
- It is a soft, silvery-white metal.

- Potassium has a density less than that of water (0.89 g/cm^3). Hence, it can float on the water surface.
- It is malleable in nature.
- Potassium has a melting point of 63.5 °C and a boiling point of 759 °C.

Chemical Properties of Potassium

Following are some chemical properties of potassium:

- It gives out hydrogen gas when reacts with water. The reaction is volatile and can cause an explosion.

$$2K + 2H_2O \rightarrow 2KOH + H_2\uparrow$$

- It is highly reactive with nitrogen, phosphorous, sulphur, and fluorine.
- It rapidly gets dissolved when reacted with dilute sulphuric acid and gives up potassium ions along with hydrogen gas.

$$2K + H_2SO_4 \rightarrow 2K^+ + SO_4 + H_2\uparrow$$

- Potassium forms potassium halides when gets reacted with halogens.

$$2K + Cl_2 \rightarrow 2KCl$$

- Potassium forms potassium halides when gets reacted with halogens.

$$2K + Cl_2 \rightarrow 2KCl$$

Uses of Potassium

Potassium is widely used in our day to day life and some of which are mentioned below:

- Industries use potassium for making soaps, detergents, dyes, gunpowder, etc.
- Potassium is used for muscle contraction.
- Excess potassium diet helps to reduce blood pressure and prevents heart strokes.
- Potassium carbonate is used for the production of glass.
- It has a high demand for fertilizers.
- It can also be used as a medium of heat exchange and is used in nuclear power plants.

4.6.2. Calcium

Calcium is the second element in the fourth period of the periodic table. Calcium (Ca) is an essential mineral that helps our bones stay strong and capable of bearing our weight. Calcium is also employed by our nervous system to aid in the transmission of impulses throughout our bodies.

Physical Properties of Calcium

Following are some physical properties of calcium:

- Calcium doesn't occur naturally in the free-state.
- It is used as an alloying agent for aluminum, lead, copper, and other base metals.
- It is a form of soft metal.
- Calcium is a good conductor of electricity.
- It's malleable and ductile in nature.

Chemical Properties of Calcium

Following are some chemical properties of calcium:

- The dissolved form of calcium bicarbonate is found in hard water.

$$CaCO_3 + CO_2 \rightarrow Ca(HCO_3)_2 + H_2O$$

- When calcium comes in contact with air, it forms a coating of nitride and oxide.

$$2Ca + O_2 \rightarrow 2CaO$$

$$3Ca + N_2 \rightarrow Ca_3N_2$$

- Compounds of calcium are highly reactive to acids.

$$CaCO_3 + HCl \rightarrow CaCl_2 + HO + CO_2\uparrow$$

Uses of Calcium

Calcium can be used for many purposes. Some of them are mentioned below:

- Calcium helps to maintain strong bones to perform many necessary functions.
- It is needed for nerves to carry messages between the brain and every body part.
- It can be used as a reducing agent in the metal extraction process.
- It is also used as an alloying agent for the production of some metals.
- Calcium concatenate is used as a food additive.
- Calcium carbide is used for the production of plastics and acetylene gas.

4.7. GROUP 7 ELEMENTS

The Group 7 elements are placed in the vertical column, second from the right-hand side of the periodic table. All Group 7 elements have 7 electrons in their outer shell. Fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At), all belong to Group 7. All these elements are called halogens. Let us discuss about these in detail.

4.7.1. Physical Properties of Group 7 Elements

Following are some physical properties of group 7 elements:

1. **Physical State.** The halogens are all diatomic and exist as F_2, Cl_2, Br_2 and I_2. The intermolecular forces are very weak in halogens. The nature of forces is van der Waals' and their magnitude increases down the group.

   Thus, F_2 and Cl_2 are gases, bromine is a volatile liquid and iodine is a volatile solid.

2. **Colour.** Halogens are coloured.

   The colour of different halogens are as follows:

Halogen	*Fluorine*	*Chlorine*	*Bromine*	*Iodine*
Colour	Pale yellow	Greenish yellow	Reddish orange	Dark violet

   Thus, the colour deepens down the group.

3. **Melting and Boiling Points.** Melting and boiling points increase with increase in atomic number. This indicates that the strength of intermolecular forces of attraction between the molecules increases with the increase in atomic number.

   The intermolecular forces in halogens are van der Waal's forces which increase with the size of the molecule.

4. **Ionization Energies.** Ionization energies of all the halogens are very high. Therefore, they have a less tendency to lose electron. However, this tendency increases down the group.

5. **Electronegativity.** The halogens have very high electronegativity. Halogens are the most electronegative elements in their respective periods. Electronegativity decreases on descending the group. Fluorine is the most electronegative element in the Periodic Table.

6. **Non-metallic Character.** All the halogens have very high values of ionization energies and exhibit non-metallic character. The non-metallic character, however, decreases down the group.

Iodine shows some distinct metallic properties *e.g.*, it possesses metallic lustre and forms positive ions like I^+, I^{3+}, etc.

4.7.2. Chemical Properties of Group 7 Elements

Following are some properties of Group 7 elements:

1. **Oxidizing Power:** Elements of this group are great oxidizing agents. *Fluorine* can oxidize all halide particles to halogen in a solution. However, oxidizing power decreases as we move down the group. *Chlorine* can oxidize bromide to bromine and iodide to iodine.

   $$Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$$
   $$Cl_2 + 2I^- \rightarrow I_2 + 2Cl^-$$

   Bromine can oxidize iodide to iodine.

   $$Br_2 + 2I^- \rightarrow I + 2Br^-$$

2. **Reaction with Hydrogen:** Acidic hydrogen halides are formed when halides react with hydrogen. The reactivity of halogen towards halogen decreases as we move down group 17. Therefore, their acidity also decreases as we move down the group.

   In dark: $H_2 + F_2 \rightarrow 2HF$

   In sunlight: $H_2 + Cl_2 \rightarrow 2HCl$

3. **Reaction with Metals:** Halogens react with metals instantly due to their high reactivity to form metal halides.

   Sodium reacts with chlorine to form sodium chloride which releases a large amount of heat energy and yellow light as it is an exothermic reaction.

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

   Metal halides are ionic in nature due to the high electronegativity of halogen and electro positivity of metals. The ionic character decreases down the group.

4.7.3. Uses of Group 7 Elements

Uses of some group 7 elements are discussed below:

Fluorine

It is mainly used for the manufacture of UF_6 for nuclear power generation. It is also used for the preparation of many fluorinating agents. The important organic chemicals derived from fluorine are the *chlorofluorocarbons and polytetrafluoroethylene* (teflon). Chlorofluorocarbons known as *freons* are used as refrigerants and in aerosols. Teflon is a plastic which is not attacked by chemical reagents and is heat-resistant. It is used for many special applications. An important inorganic chemical made from HF is cryolite (Na_3AlF_6) which is used for the production of aluminium. Some of the uses of HF are in the glass industry as an *etching agent* and in the manufacture of fluoride salts. Prominent among the fluorides is NaF used for the fluorination of water; one part per million level fluoride in drinking water prevents tooth decay.

Chlorine

The chief uses of chlorine are:

(i) In the production of organic compounds like polyvinyl chloride, chlorinated hydrocarbons, pharmaceuticals, herbicides, pesticides, etc.

(ii) It is used to bleach paper pulp and textiles and as disinfectant for sterilizing drinking water.

(iii) It is used in the preparation of solvents such as CCl_4, $CHCl_3$, CH_2Cl_2 and trichloroethene.

(iv) In the production of inorganic compounds like HCl, PCl_3, PCl_5 sodium hypochlorite (NaOCl), bleaching powder ($CaOCl_2$) etc.

Bromine

Some important uses of bromine are:

(i) Bromine is used in the preparation of ethylene bromide which is used as an additive to *leaded* petrol.

(ii) Bromine is used to make AgBr for photography.

(iii) It is also used in the preparation of dyes, fire retardants disinfectants, fumigants and medicines.

Iodine

Some important uses of iodine are:

(i) Iodine is used as an anticeptic in the form of an alcoholic solution, which is known as *tincture of iodine*.

(ii) It is also used for the preparation of iodoform and potassium iodide.

(iii) Iodide ion is necessary for the normal functioning of the thyroid gland. Insufficient iodide in the diet leads to *goitre* (enlargement of thyroid gland). Hence, sodium or potassium iodide is added to table salt and this type of salt is known as “*iodized*” salt.

4.8. ELEMENTS OF THE FIRST TRANSITION SERIES

In general, any element which corresponds to the *d-block* of the modern periodic table (which consists of groups 3-12) is considered to be a transition element. The elements of the first transition series are scandium (Sc), titanium (Ti), vanadium (V), chromium (Cr), manganese (Mn), iron (Fe), cobalt (Co), nickel (Ni), copper (Cu), and zinc (Zn).

4.8.1. Properties of the Elements of the First Transition Series

Following are some common characteristics of elements of the first transition series:

- **Physical States:** Except for mercury, all transition elements exist in solid state.
- **Metallic Property:** All the transition elements are metals; this is because the number of electrons in outermost shell is only 2.
- **Magnetic Property:** Most of the transition metals are paramagnetic in nature. This is due to the presence of unpaired electrons in the transition elements.
- **Variable Oxidation State:** Transition metals can form compounds with a wide range of oxidation states. Some of the observed oxidation states of the elements of the first transition series are shown in Table. 4.3.

Table 4.3. Oxidation States of Elements of the First Transition Series

^{21}SC	**^{22}Ti**	**^{23}V**	**^{24}Cr**	**^{25}Mn**	**^{26}Fe**	**^{27}Co**	**^{28}Ni**	**^{29}Cu**	**^{30}Zn**
								1+	
		2+	2+	2+	2+	2+	2+	2+	2+
3+	3+	3+	3+	3+	3+	3+	3+	3+	
	4+	4+	4+	4+					
		5+							
			6+	6+	6+				
				7+					

As we move from left to right across the first transition series, we see that the number of common oxidation states increases at first to a maximum towards the middle of the table, then decreases.

- **Formation of Coloured Compounds:** Many compounds of transition elements are coloured in contrasts to those of *s* and *p* block elements. In compound state due to the surrounding groups (ligands), the *d*-orbitals of transition elements are not degenerate but split into two groups of different energy.
- **Catalytic Property:** Many transition metals and their compounds have catalytic properties. For example, V_2O_5, Fe, $FeCl_3$, Ni, Pd, etc. This is due to following reasons:
    - (i) *Variable oxidation state*: Due to variable oxidation state they form unstable intermediate compounds and provide a new path with lower activation energy for the reaction (Intermediate compound formation theory).
    - (ii) *Large Surface area*: Finely divided transition metals or their compounds provide a large surface area for adsorption and the adsorbed reactants react faster due to the closer contact.
- **Formation of Alloys:** Alloys are homogenous solid solutions of two or more metals obtained by melting the components and then cooling the melt. These are formed by metals whose atomic radii differ by not more than 15% so that the atoms of one metal can easily take up the positions in the crystal lattice of the other. Since transition metals have similar atomic radii, they form alloys very readily.

GLOSSARY

- **Covalent Radius:** Half of the inter-nuclear distance between two atoms of the element held by a single covalent bond.
- **Electron Affinity:** The energy change taking place when an electron is added to an isolated gaseous atom of the element.
- **Electronegativity:** It is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.
- **Group:** A vertical column of elements in the periodic table.
- **Ionization Energy:** The energy required to remove the outermost electron from an isolated gaseous atom of the element.
- **Metallic Radius:** Half of the inter-nuclear distance between two nearest atoms in the metallic lattice.
- **Period:** A horizontal row of elements in the periodic table.

- **Periodic Table:** Arrangement of elements in the increasing order of atomic number such that elements with similar properties fall under same vertical column.
- **Triads:** The elements arranged according to their increasing atomic mass, in a group of three.

SUMMARY

- The periodic table is the arrangement of all the known elements according to their properties so that similar elements fall within the same vertical column and dissimilar elements are separated.
- According to the law of octaves, *"If the elements are arranged in the order of increasing atomic masses, the eighth element, starting from a given one is a kind of repetition of the first—like the eighth note in an octave of music."*
- The modern periodic law states that: "The physical and chemical properties of the elements are the periodic function of their atomic numbers."
- Modern periodic table arranges the elements in the order of their atomic numbers. There are 118 elements in the Modern periodic table.
- A horizontal row of a periodic table is called a period. There are seven periods in all, which are numbered as 1, 2, 3, 4, 5, 6 and 7.
- A vertical column of elements in the periodic table is called a group. There are 18 groups in the modern periodic table.
- Electronegativity is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.
- Electron affinity is the energy change taking place when an isolated gaseous atom accepts an electron to form a monovalent gaseous anion.
- The lattice energy is the energy change occurring when one mole of a solid ionic compound forms in its gaseous state.
- The main group elements are any of the chemical elements belonging to the *s* and *p* blocks of the periodic table. Specific examples of main group elements include helium, lithium, beryllium, boron, carbon, nitrogen, oxygen, and fluorine.
- Group 7 elements are called halogens.
- Any element which corresponds to the *d*-block of the modern periodic table (which consists of groups 3-12) is considered to be a transition element.

EVALUATION

I. Multiple Choice Questions

1. In Mendeleev's periodic table, the elements are arranged according to their
(a) Atomic size
(b) Atomic number
(c) Atomic mass
(d) None of these

2. In the modern periodic table the elements are arranged according to their
(a) Atomic size
(b) Atomic number
(c) Atomic mass
(d) None of these

3. Elements of Group-2 are called
(a) Alkali metals
(b) Alkaline earth metals
(c) Atomic mass
(d) None of these

4. Elements of which group are called chalcogens?
(a) Group-1
(b) Group-2
(c) Group-16
(d) Group-17

5. The tendency of gaseous atom to form anion is expressed in terms of
(a) Ionization energy
(b) Electron affinity
(c) Electronegativity
(d) None of these

6. On moving down a group, the atomic size
(a) increases
(b) decreases
(c) becomes stable
(d) None of these

7. On moving across a period from left to right, the metallic character
(a) increases
(b) decreases
(c) becomes stable
(d) None of these

8. The energy change occurring when one mole of a solid ionic compound forms in its gaseous state.
(a) Ionization energy
(b) Lattice energy
(c) Electronegativity
(d) None of these

9. Which of the following is not a Group-7 element?
(a) Fluorine
(b) Aluminium
(c) Bromine
(d) Iodine

II. Fill in the Blanks

1. A horizontal row of periodic table is known as _______ .
2. The elements in a _______ have same valence shell.
3. A vertical column of elements in periodic table is known as _______ .
4. The elements in a _______ have same outer shell configuration.
5. Group 1 elements are known as _______ .
6. Group 17 elements are known as _______ .
7. Noble gases are the elements belonging to group _______ .
8. The elements of group 16 are known as _______ .
9. Alkali metals occur at the _______ in Lothar Meyer's atomic volume curve.
10. Transition metals belong to _______ block of the periodic table.

III. Answer the Following Questions

1. What is a group and a period of periodic table?
2. What is the number of groups and number of periods in the long form of periodic table?
3. What property did Mendeleev use to arrange elements in the periodic table?
4. State the periodic law on which the modern periodic table based. Who proposed this law?
5. What are the common features of the electron structures of elements in: (i) Group 2 (ii) Period 3.
6. The atomic number of carbon is 6 and that of silicon is 14. Write the electronic configurations of carbon and silicon atoms. In which group or groups do these elements occur?
7. On the basis of electronic configuration how will you identify:
    - (i) chemically similar elements
    - (ii) the first element of a period
    - (iii) the last element of a period.
8. Consider the following elements:

    Ca, Na, Mg, Al, Be, Si
    - (i) Which of these elements belong to the same period?
    - (ii) Which of these elements belong to the same group?
9. Define electronegativity. How does it vary along a period and along a group?
10. What are the elements of the first transition series? Write their common properties.

SEMESTER–II (Period–IV)

C10CH5

5
TOPIC

Chemical Bonding

Learning Objectives

Upon completion of this topic, learners will be able to:

- Distinguish the different types of chemical bonding
- Discuss hybridization of atomic orbitals
- Demonstrate the types of bonding of molecules by using the Lewis Structure
- Discuss electronegativity relative to the concept of bond polarity
- Discuss the molecular geometry
- Distinguish between inter-atomic bonding and intermolecular bonding and
- Explain coordinate covalent (dative) bond.

Introduction

Matter is made up of one or different type of elements. Under normal conditions no other element exists as an independent atom in nature, except noble gases. However, a group of atoms is found to exist together as one species having characteristic properties. Such a group of atoms is called a **molecule**. Obviously there must be some force which holds these constituent atoms together in the molecules. *The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a* ***chemical bond***.

5.1. BONDING TYPES

The strength and properties of the chemical bonds formed vary. There are four primary types of chemical bonding. These are:

1. Ionic Bonding
2. Covalent Bonding
3. Metallic Bonding
4. Coordinate Covalent Bonding

5.1.1. Ionic Bonding

The compounds which are made up of ions are known as ionic compounds. In an ionic compound, the positively charged ions (cations) and negatively charged ions (anions) are held together by the strong electrostatic forces of attraction. The forces which hold the ions together in an ionic compound are known as **ionic bonds** or **electrovalent bonds**. Since an ionic bond consists of an equal number of positive ions and negative ions, the overall charge on an ionic compound is zero. For example, sodium chloride (NaCl) is an ionic compound which is made up of equal number of positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-). Some of the common ionic compounds, their formulae and the ions present in them are given in Table 5.1.

Table 5.1. Formulae and Nomenclature of Some Ionic Compounds

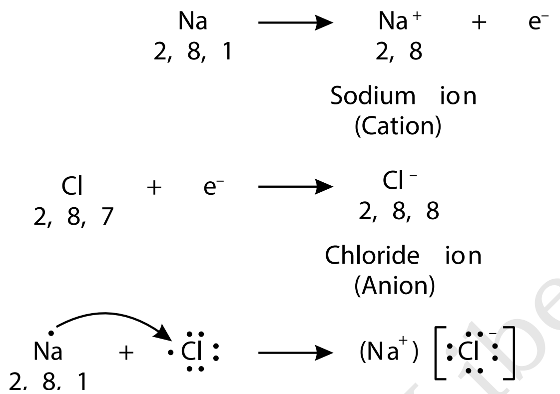
Nomenclature	Formula	Ions present
Aluminium oxide	Al_2O_3	Al^{3+} and O^{2-}
Ammonium chloride	NH_4Cl	NH_4^+ and Cl^-
Calcium hydroxide	$Ca(OH)_2$	Ca^{2+} and OH^-
Calcium nitrate	$Ca(NO_3)_2$	Ca^{2+} and NO_3^-
Calcium oxide	CaO	Ca^{2+} and O^{2-}
Copper sulphate	$CuSO_4$	Cu^{2+} and SO_4^{2-}
Magnesium chloride	$MgCl_2$	Mg^{2+} and Cl^-
Potassium chloride	KCl	K^+ and Cl^-
Potassium hydroxide	KOH	K^+ and OH^-
Sodium carbonate	Na_2CO_3	Na^+ and CO_3^{2-}
Sodium hydroxide	NaOH	Na^+ and OH^-

Ionic compounds are made up of a metal and a non-metal (except ammonium chloride which is an ionic compound made up of only non-metals). So, whenever a bond involves a metal and a non-metal, we call it ionic bond.

5.1.2. Formation of Ionic Bonding

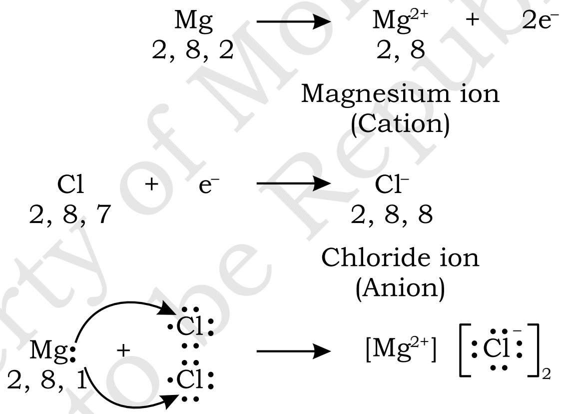
An ionic bond changes the electronic configurations of the atoms. Metal atoms lose their outermost electron(s), forming cations. Non-metal atoms gain electron(s) to fill their outermost shell, forming anions. The electrostatic force of attraction between the oppositely charged ions holds the ions together. For example,

(a) When a hot sodium atom is placed in chlorine gas, a reaction takes place resulting in formation of sodium chloride.



Formation of Sodium Chloride

(b) When a magnesium atom comes in contact with chlorine gas, it forms magnesium chloride.



Formation of Magnesium Chloride

5.1.2.1. Factors Influencing the Formation of Ionic Bonding

The following three factors influence the formation of ionic bonding:

1. **Ionization energy:** It is defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom of an element. The lesser the ionization energy, the greater is the ease of the formation of a cation.
2. **Electron affinity:** It is defined as the amount of energy released when an electron is added to an isolated gaseous atom of an element. The higher the energy released during this process, the easier will be the formation of an anion.

3. **Lattice energy:** It is defined as the amount of energy released when cations and anions are brought from infinity to their respective equilibrium sites in the crystal lattice to form one mole of the ionic compound. The higher the lattice energy, the greater is the tendency of the formation of an ionic bond. The higher the charges on the ions and smaller the distance between them, the greater is the force of attraction between them.

5.1.3. Properties of Ionic Compounds

ACTIVITY 5.1

Illustrating Physical Properties of Ionic Compounds

1. Take a sample of sodium chloride or any other salt from the science laboratory.

   What is the physical state of this salt?

2. Take a small amount of a sample on a metal spatula and heat directly on the flame as shown in Fig. 5.1.

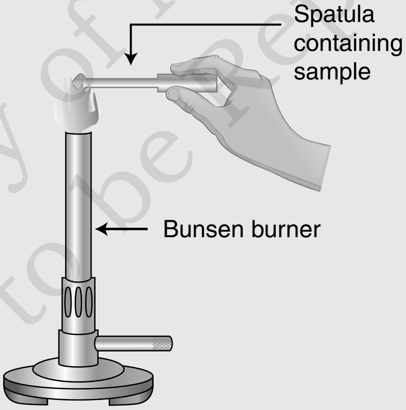


Fig. 5.1. Testing melting point of sodium chloride

   What did you observe? Did the sample impart any colour to the flame? Does this compound melt?

3. Try to dissolve the sample in water, petrol and kerosene. Is it soluble?
4. Make a circuit as shown in Fig. 5.2 and insert the electrodes into a solution of salt. What did you observe?

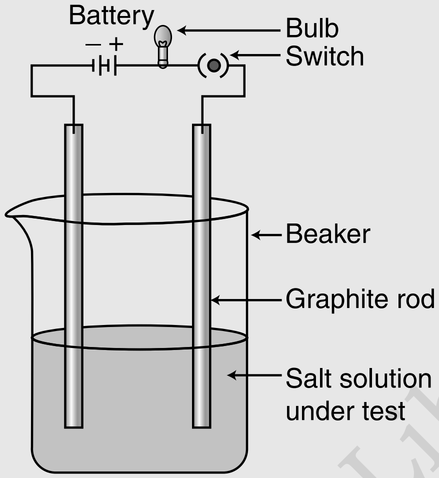


Fig. 5.2. Testing electrical conductivity of salt solution

- What is your inference about the nature of this compound?

You may have observed the following general properties for ionic compounds:

- Ionic compounds are usually crystalline solids.
- Ionic compounds have high melting and high boiling points.
- Ionic compounds are usually soluble in water but insoluble in organic solvents like petrol and kerosene.
- Ionic compounds conduct electricity when dissolved in water or when melted. When we dissolve the ionic solid in water or melt it, the crystal structure is broken down to form ions. These ions help in conducting electricity.

5.2. HYBRIDIZATION OF ATOMIC ORBITALS

Hybridization is a theoretical concept which has been introduced to explain some structural properties such as shapes of molecules or equivalency of bonds, etc., which cannot be explained by simple theories of valency.

According to this concept, valence orbitals of the atom intermix to give rise to another set of equivalent orbitals before the formation of bonds. These orbitals are called **hybrid orbitals** or **hybridized orbitals** and the phenomenon is referred to as **hybridization**. Thus, ***hybridization*** may be defined as *the phenomenon of intermixing of atomic orbitals of slightly different energies of the atom (by redistributing their energies) to form new set of orbitals of equivalent energies and identical shape.*

There are many different types of hybridization depending upon the type of orbitals involved in mixing such as sp^3, sp^2, sp, etc. Let us discuss various types of hybridization alongwith some examples.

sp^3-Hybridization

The type of hybridization involves the mixing of one orbital of *s*-sub-level and three orbitals of *p*-sub-level of the valence shell to form four sp^3 hybrid orbitals of equivalent energies and shape. Each sp^3 hybrid orbital has 25% *s*-character and 75% *p*-character. These hybridized orbitals tend to lie as far apart in space as possible so that the repulsive interactions between them are minimum. The fours sp^3 hybrid orbitals are directed towards the four corners of a tetrahedron. The angle between the sp^3 hybrid orbitals is 109.5° (Fig. 5.3).

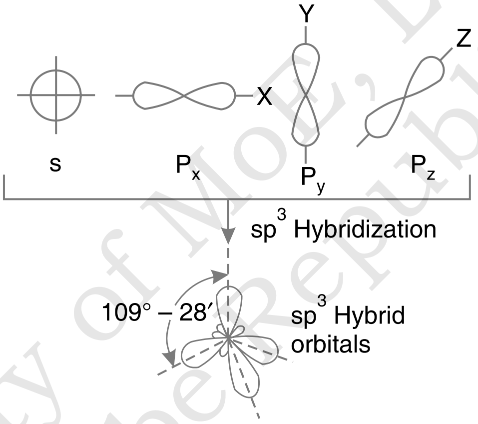


Fig. 5.3. Tetrahedral orientations of sp^3 hybrid orbitals

sp^3 hybridization is also known as **tetrahedral hybridization**. The molecules in which central atom is sp^3 hybridized and is linked to four other atoms directly, have **tetrahedral shape**.

Let us study an example of a molecule which involves sp^3 hybridization.

1. **Formation of methane (CH_4).** In methane molecule carbon atom acquires sp^3 hybrid state.

   Carbon atom has ground state configuration of $2s^2$ $2b_x$, $2b_y$ and in excited state its configuration is $2s$, $2p_x$, $2p_y$, sp_z. One orbital of 2*s*-sub-shell and three orbitals of 2*p*-sub-shell of excited carbon atom hybridise to form four sp^3 hybird orbitals.

   The sp^3 hybrid orbitals of carbon atom are directed towards the corners of regular tetrahedron. Each of the sp^3 hybrid orbitals overlaps axially with half-filled 1*s*-orbital of hydrogen atom constituting a sigma bond (Fig. 5.4).

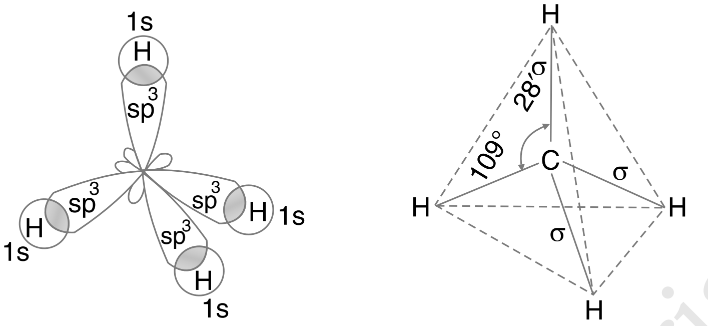


Fig. 5.4. Orbital picture of methane

Because of sp^3 hybridization of carbon atom, $\mathbf{CH_4}$ molecule has **tetrahedral shape**.

sp^2-Hybridization

This type of hybridization involves the mixing of one orbital of *s*-sub-level and two orbitals of *p*-sub-level of the valence shell to form three sp^2 hybrid orbitals. These sp^2 hybrid orbitals lie in a plane and are directed towards the corners of equilateral triangle (Fig. 5.5).

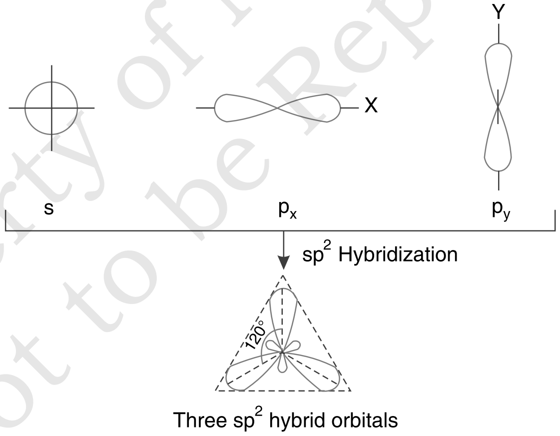


Fig. 5.5. Trigonal planer orientation of sp^2 hybrid orbitals

Each sp^2 hybrid orbital has one-third *s*-character and two-third *p*-character. sp^2 hybridization is also called **trigonal hybridization**. The molecules in which central atom is sp^2 hybridized and is linked to three other atoms directly have triangular planar shape.

Let us study an example of a molecule which involves sp^2 hybridization.

1. **Formation of boron trichloride ($BC1_3$).** Boron ($_5$B) atom has ground state configuration as $1s^2$, $2s^2$, $2p^1$. But in the excited state its configuration is $1s^2$, $2s^1$, $2p_x^{\ 1}$, $2p_y^{\ 1}$. One $2s$-orbital of

boron intermixes with two 2*p*-orbitals of excited boron atom to form three sp^2 hybrid orbitals.

The sp^2 hybrid orbitals of boron are directed towards the corners of equilateral triangle and lie in a plane. Each of the sp^2 hybrid orbitals of boron overlaps axially with 3*p*-half-filled orbital of chlorine atom to form three B-Cl sigma bonds as shown in Fig. 5.6.

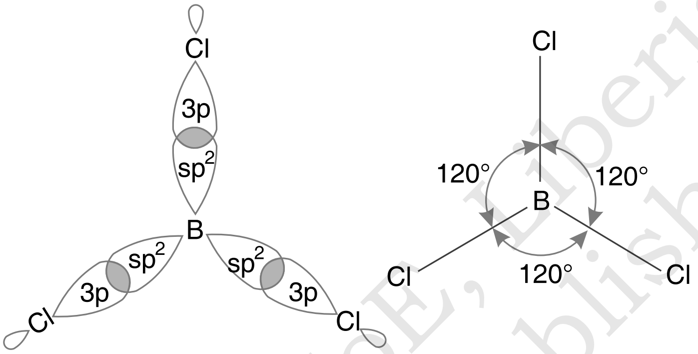


Fig. 5.6. Orbital diagram of $BC1_3$

Because of sp^2 hybridization of boron, **BCl_3** molecule has **triangular planar shape**.

sp-Hybridization

This type of hybridization involves the mixing of one orbital of *s*-sub-level and one orbital of *p*-sub-level of the valence shell of the atom to form two *sp*-hybridized orbitals of equivalent shapes and energies. These *sp*-hybridized orbitals are oriented in space at an angle of 180° (Fig. 5.7). This hybridization is also called **diagonal hybridization**. Each *sp* hybrid orbital has equal *s* and *p* character, *i.e.*, 50% *s*-character and 50% *p*-character. The molecules in which the central atom is *sp*-hybridized and is linked to two other atoms directly have linear shape.

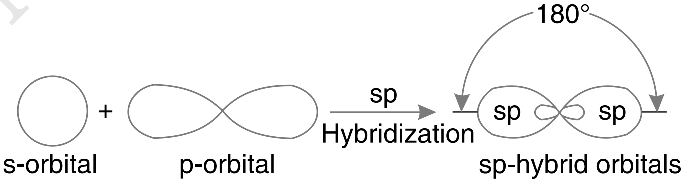


Fig. 5.7. Formation of *sp* hybrid orbitals

Let us study the **formation of beryllium fluoride ($BeCl_2$).** Beryllium ($_4Be$) atom has a ground state configuration as ls^2, $2s^2$. In the excited state however, its configuration becomes υ^2, 2^1, $2p^1x$. One 2*s*-orbital and one 2*p*-orbital of excited beryllium atom undergo 2*p*-hybridization to form two *sp*-hybridized orbitals.

The two *sp*-hybird orbitals are linear and oriented in opposite directions at an angle of 180°. Each of the *sp*-hybridized orbital overlap axially with 3*p*-half-filled orbital of chlorine atom to form two Be—Cl sigma bonds (Fig. 5.8).

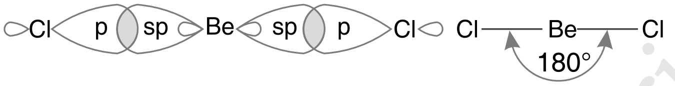


Fig. 5.8. Orbital diagram of $BeCl_2$

Due to the *sp*-hybridised state of beryllium, **$BeCl_2$** molecule has **linear shape**.

5.2.1. Covalent Bonding

The chemical bond formed by sharing of electrons between two atoms is known as a **covalent bond**. The compounds containing covalent bonds are known as covalent compounds. A covalent bond is formed when both the reacting atoms need electrons to achieve the inert gas electron arrangement. Now, the non-metals have usually 5, 6 or 7 electrons in the outermost shells of their atoms. So, all the non-metal atoms need electrons to achieve the inert gas structure. They get these electrons by mutual sharing. Thus, whenever a non-metal combines with another non-metal, covalent bond is formed.

Table 5.2. Formulae and Nomenclature of Some Covalent Compounds

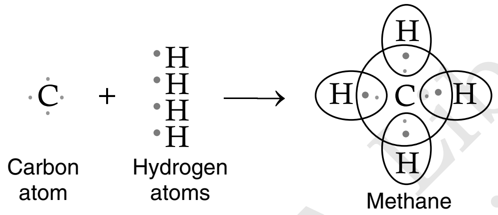
Nomenclature	Formula	Elements present
Methane	CH_4	C and H
Ethane	C_2H_6	C and H
Ethene	C_2H_4	C and H
Ethyne	C_2H_2	C and H
Ammonia	NH_3	N and H
Alcohol (Ethyl alcohol)	C_2H_5OH	C, H and O
Hydrogen sulphide gas	H_2S	H and S
Carbon dioxide	CO_2	C and O
Carbon disulphide	CS_2	C and S
Carbon tetrachloride	CCl_4	C and Cl

Glucose	$C_6H_{12}O_6$	C, H and O
Cane sugar	$C_{12}H_{22}O_{11}$	C, H and O
Urea	$CO(NH_2)_2$	C, O, N and H

5.2.1.1. Formation of Covalent Bonding

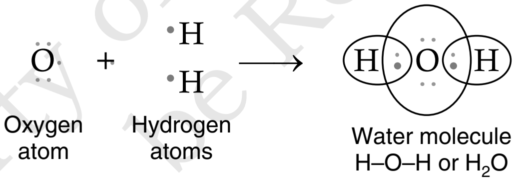
Covalent bonding between atoms of different elements.

(i) Carbon atom shares four electrons to form methane.



Covalent bonding or sharing of electrons only takes place in outermost shells of atoms to attain inert gas electron arrangement.

(ii) As in water molecule, 2 hydrogen atoms share electrons with oxygen atom.



5.2.1.2. Factors that Influence the Formation of Covalent Bonding

Since the covalent bond is formed by mutual sharing of electrons, therefore, the factors which favour the formation of covalent bonds are as follows:

1. **Electron Affinity:** A covalent bond is generally favoured between the two atoms if both the atoms have high electron affinity.
2. **Ionisation Energy:** The ionisation energy of both the atoms participating in bonding should be high.
3. **Atomic Size:** Atomic size of the atoms forming covalent bond should be smaller. Smaller the atomic radii of atoms, stronger the covalent bond will be. For example, H−H bond is stronger than Cl−Cl bond which in tum is stronger than Br-Br bond.

4. **Electronegativity:** The electronegativities of both then atoms should be high. The difference of electronegativities between the two atoms should be minimum.

5.2.1.3. Properties of Covalent Bonding

ACTIVITY 5.2

Illustrating Physical Properties of Covalent Compounds

Let us test some covalent compounds in different ways:

1. Take sample of cooking oil. Try to dissolve it in water and ethanol. Does it dissolve?
2. Have you ever observed a burning candle wax? If not, take a candle wax and observe it burning. How much time does it take to melt down?
3. Take a pan and add water to it. Let it boil. Do you know the boiling point of water?
4. Now add two electrodes to the water pan making a circuit. What did you observe? What would have happened if you would have added NaCl salt in the pan?

   What can you now say about these covalent compounds?

You have observed the following properties of covalent compounds:

- Covalent compounds are usually liquids, gases or solids. For example, alcohol, benzene, water and cooking oil are liquids. Methane, ethane and chlorine are gases. Glucose, urea, and wax are solid covalent compounds.
- Covalent compounds have usually low melting points and low boiling points.
- Covalent compounds are usually insoluble in water, but they are soluble in organic solvents. Some of the covalent compounds like glucose, sugar and urea are soluble in water.
- Covalent compounds do not conduct electricity because they do not contain ions.

ACTIVITY 5.3

Detecting an Ionic Bond or Covalent Bond

1. Take the sample such as common salt (NaCl) provided.
2. Try to dissolve it in water.
3. If it dissolves, chances are it is likely to be an ionic compound. But, you already know some covalent compounds like glucose, urea and sugar are soluble in water.
4. Now, perform electrical conductivity test (Fig. 5.9).

5. If the NaCl sample dissolves in water, arrange a circuit with two electrodes and a bulb.
6. Figure out whether the bulb glows or not. According to your observation conclude the bond present in the sample.
7. Make a report on the properties of ionic and covalent compounds.

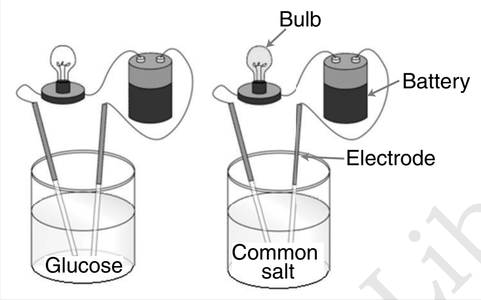
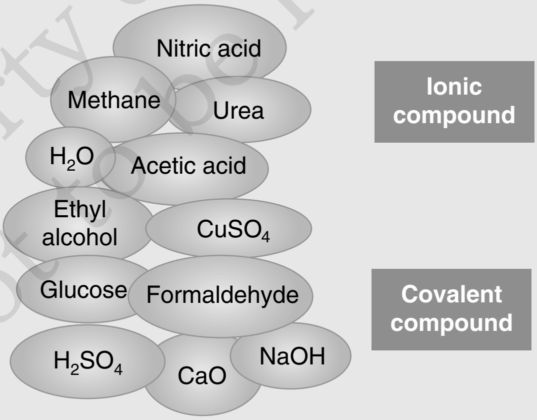


Fig. 5.9. Electrical conductivity test

ACTIVITY 5.4

Identifying Ionic and Covalent Compounds

Choose the ionic as well as covalent compounds from the bubbles and make a table in your exercise notebook.

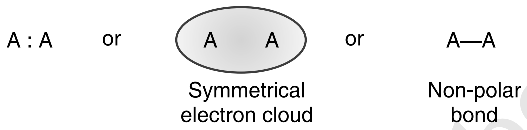


5.2.2. Bond Polarity

Covalent bonds between different atoms may acquire some ionic characters which can be explained on the basis of concept of electronegativity. **Electronegativity** of an element is defined as *the power of its atoms to attract the bonding or shared pair of electrons towards itself.*

5.2.2.1. Polar Character of Covalent Bond

When covalent bond is formed between two similar atoms, the shared pair of electrons lies midway between the nuclei of the two atoms, because both the atoms have same attraction for the bonding electrons. The electron cloud constituting the covalent bond is symmetrically distributed around the atoms. Such a covalent bond is called a **non-polar covalent bond**.



For example, molecules like H_2, Cl_2, O_2 and N_2 contain non-polar bonds.

In case, the covalent bond is formed between two **dissimilar atoms**, one of which has a larger value of electronegativity, the bonding pair of electrons is displaced towards the more electronegative atom. In other words, electron cloud containing the bonding electrons gets distorted and the charge density concentrates around the more electronegative atom.

Due to the unequal distribution of electron charge density, the more electronegative atom acquires a partial negative charge (indicated as δ^-) whereas the less electronegative atom acquires a partial positive charge indicated as δ^+. Thus, *a covalent bond develops a partial ionic character as a result of the difference of electronegativities of the atoms comprising the bond. Such a bond is called* **polar covalent bond** (as shown in Fig. 5.10).

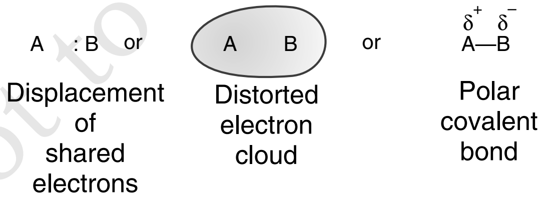
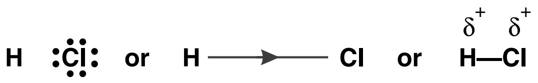


Fig. 5.10. Polarity in covalent bond

For example, the bond between hydrogen and chlorine atoms in HCl molecule is polar because the shared electron pair is displaced towards chlorine atom which is more electronegative.



The extent of ionic character in a covalent bond depends upon the difference of electronegativities of the two atoms forming a bond. The difference of electronegativities, greater is the percentage of ionic character in a bond. For example, hydrogen fluoride is more polar than hydrogen chloride because the difference of between H and F is more than that between H and Cl.

	$\overset{\delta+}{\text{H}}—\overset{\delta-}{\text{F}}$	$\overset{\delta+}{\text{H}}—\overset{\delta-}{\text{Cl}}$
Electronegativity	2.1 4.0	2.1 3.0
Difference	1.9	1.9
	More ionic than H—Cl	More covalent than HF

It has been observed that the bond has 50% ionic character and 50% covalent character if the difference of electronegativities of the participating atoms is 1.7. On the other hand, the covalent character dominates if the difference of electronegativities is less than 1.7 while ionic character dominates if the difference of electronegativities is greater than 1.7.

5.2.2.2. Dipole Moment (μ)

Degree of polarity in a bond is measured in terms of **dipole moment** which is defined as *the product of the magnitude of charge (e) on any one of the atoms and the distance (d) between them.* Dipole moment is represented by a Greek letter 'μ'. It can be expressed mathematically, as

$$\mu = e \times d$$

Dipole moment is expressed in a unit called *Debye* and is represented by D.

$$\mathbf{1\ D = 1 \times 10^{-18}\ e.s.u.–cm. = 3.335 \times 10^{-30}\ C\ m}$$

For example, dipole moment of HC1 is 1.03×10^{-18} *e.s.u.-cm* and is expressed as 1.03 D. Dipole moment can be determined experimentally and its-value can give an idea of the polar character of a molecule. It is a *vector quantity* as it has a direction as well as magnitude. The direction of dipole moment is usually represented by an arrow → pointing from positive end towards the negative end.

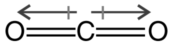
5.2.2.3. Polar and Non-Polar Molecules

Polar character of bonds also affects the polarity in a molecule. In **diatomic** molecules, the two atoms are bonded by a covalent bond.

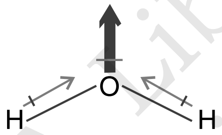
In such molecules, the dipole movement of bond also represents the dipole moment of the molecule, some examples are:

H—H	Cl—Cl	H→Cl	H→F
$\mu = 0$	$\mu = 0$	$\mu = 1.03D$	$\mu = 1.78D$

In polyatomic molecules having more than one bonds, the idea of a dipole can be applied to individual bonds within the molecule. The dipole moment of individual bond in a polyatomic molecule is referred to as **bond dipole**. The dipole moment of the molecule depends upon the orientations of various bond dipoles. For example, carbon dioxide (CO_2) and water (H_2O) are both triatomic molecules but dipole moment of carbon dioxide is zero whereas that of water is 1.83 D. This can be explained on the basis of their structures.



Bond dipoles cancel
Resultant $\mu = 0$



Bond dipoles do not cancel Resultant
$\mu = 1.83$ D

Carbon dioxide is a **linear molecule** in which the two C = O bonds are oriented in the opposite directions at an angle of 180°. The bond dipole of two C = O bonds cancel each other. Therefore, the resultant dipole moment of the molecule is zero. Hence, CO_2 is a non-polar molecule. On the other hand, water molecule has a bent structure in which two O—H bonds are oriented at an angle 104.5°. Therefore, the bond dipoles of two O—H bonds do not cancel each other. Hence, the resulting dipole moment of H_2O molecule is not zero but has a value equal to 1.83 D.

It may be noted that the molecules having **zero dipole moment (μ = 0)** are called **nonpolar molecules**. On the other hand the molecules having dipole moment greater than zero ($\mu > 0$) are **polar molecules**.

5.3. MOLECULAR GEOMETRY

Atoms bind together to form molecules, and atoms of each element are different. They are different sizes, they have different electron configurations, and their chemical behaviour means that they act differently inside a molecule. In this way, the variety of atoms and the behaviour of their electrons means that molecules form in certain ways, and their shapes can be predicted using molecular geometry. Thinking of atoms as balls and atomic bonds as sticks, one can construct "toothpick models" that are approximations of molecular shapes.

Molecular geometry is explained primarily through Valence Shell Electron Pair Repulsion (VSEPR) theory. While electrons can coexist within an orbital in pairs, they are still all negatively charged, so those pairs will repel each other. Most molecular geometry can be explained by counting the valence electrons in a molecule and accounting for how they want to be as far away from each other as possible. As molecules get larger, they develop stable configurations wherein electron pairs find a "comfortable" position apart from each other, and so familiar geometric shapes are formed.

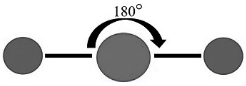
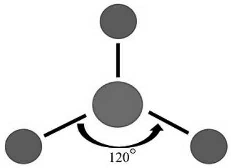
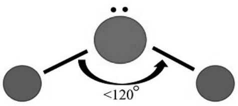
Valence Shell Electron Pair Repulsion (VSEPR) theory

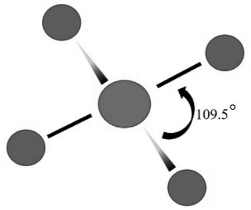
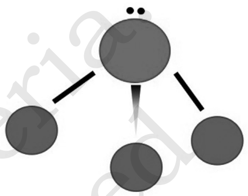
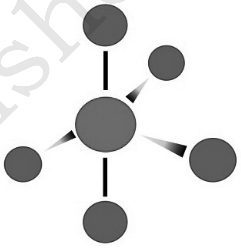
The valence shell electron pair repulsion (VSEPR) theory is a model used to predict 3-D molecular geometry based on the number of valence shell electron bond pairs among the atoms in a molecule or ion. This model assumes that electron pairs will arrange themselves to minimize repulsion effects from one another. In other words, the electron pairs are as far apart as possible.

The VSEPR model is useful for predicting and visualizing molecular structures. The structures are: linear, trigonal planar, angled, tetrahedral, trigonal pyramidal, trigonal bipyramidal, disphenoidal (seesaw), t-shaped, octahedral, square pyramidal, square planar, and pentagonal bipyramidal.

5.3.1. Shapes of Some Simple Molecules

Simple molecules have simple shapes, but things get increasingly more complicated when more atoms are involved. Some of the major molecular shapes are listed below.

S.No.	Shape	Structure
1.	**Linear:** In a linear molecule, the electron pairs take up opposite sides of the central atom.	
2.	**Trigonal Planar:** In a trigonal molecule, three pairs are evenly spaced around the central atom.	
3.	**Angular:** In an angular (also called bent) molecule, a lone pair of electrons takes what would be a flat molecule and pushes the arms further away from it and toward each other in the "down" direction.	

4.	**Tetrahedral:** In a tetrahedral molecule, the four electron pairs get further away from each other by leaving the flat plane and separating in 3 dimensions.	
5.	**Trigonal Pyramidal:** Much like an angular molecule, the multiple arms in a pyramidal molecule have been pushed out of the flat plane by the repulsion from a nonbonding pair.	
6.	**Trigonal Bipyramidal:** In a trigonal bipyramidal molecule, three of the pairs are on a flat plane, as in a trigonal planar molecule. Two more arms are aimed away from the flat plane, one “up” and one “down.”	

5.3.1.1. How to Calculate Molecular Geometry

In short, count the valence electrons to determine the shape of the molecule. Follow these steps to calculate molecular geometry:

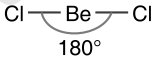
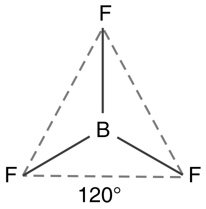
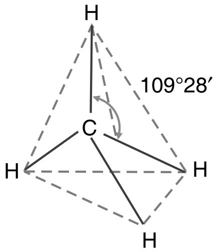
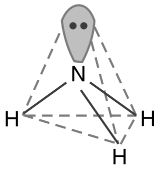
1. Determine the Lewis structure of the atom.
2. Count the valence electrons of the central atom in the molecule. The number of valence electrons, for the most part, can be determined by the element's group on the periodic table.
3. Determine the number of bonding electrons, and the number of electrons in lone pairs.
4. The number of bonds coming from the central atom and the number of lone pairs will determine the shape of the molecule. The bond angle is the degree measurement of the angle between any two bonds.

No. of Bonds on Central Atom	No. of Lone Pairs	Molecular Shape	Bond Angle	Examples
2	0	Linear	180	BeF_2, $BeCl_2$, CO_2
3	0	Trigonal planar	120	BF_3, $AlCl_3$

2	1	Angular (bent)	<120	O_3, $SnCl_2$
4	0	Tetrahedral	109.5	CH_4, SiF_4
3	1	Trigonal pyramidal	<109.5	PCl_3, NCl_3, NH_3
5	0	Trigonal bipyramidal	90, 120, 180	PF_5, PCl_5, $SbCl_5$

Example 5.1: *Describe the shapes of the following molecules on the basis of VSEPR Model:* $BeCl_2$, BF_3, CH_4, NH_3, H_3O^+, CO_2, SF_6

Solution: The Lewis dot structures and shapes of various species are given below in tabular form:

Species	Lewis dot structure	Electron pairs around central atom		Arrangement of electron pairs	Shape
		bp	lp		
$BeCl_2$	Cl : Be : Cl	2	0		Linear
BF_3	F:B:F ·· F	3	0		Trigonal planar
CH_4	H ·· H:C:H ·· H	4	0		Tetrahedral
NH_3	·· H:N:H ·· H	3	1		Pyramidal
CO_2	O::C::O	2	0	O = C = O	Linear

5.4. METALLIC BONDING

The force which binds various metal atoms together is called ***metallic bond****.* The metallic bond is neither a covalent bond nor an ionic bond because these bonds are not able to explain properties of metals.

For example, metals are very good conductors of electricity but in solid state. Both ionic and covalent compounds cannot do so with the exception of graphite.

5.4.1. Formation of Metallic Bonding

Loreutz proposed the theory of electron gas model or **electron sea model** for metallic bonding.

In this model, the metal is pictured as an array of metal cations in a "sea" of electrons. The atoms in a metallic solid contribute their valence electrons to form a "sea" of electrons that surrounds metallic cations. Delocalised electrons are not held by any specific atom and can move easily throughout the solid. A metallic bond is the attraction between these electrons and the metallic cation.

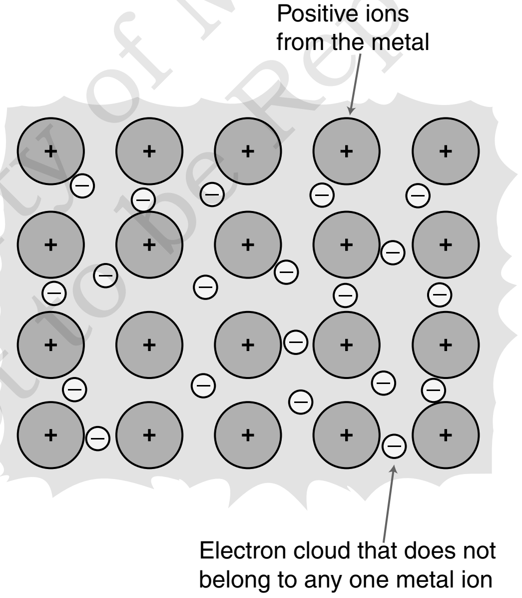


Fig. 5.11. Formation of metallic bond

5.4.1.1. Factors that influence the formation of metallic bonding

The factors influencing the formation of metallic bonds are:

- The *ionisation energy* of the metal should be low so that the valence electrons are loosely held by the nucleus and thus become mobile.

- The metals should have high *electropositivity*, so they do not accept electrons and form anions.
- The *vacant orbital* should be more than the number of valence electrons so that the mobile electrons can easily move in the vacant orbitals.

5.4.2. Properties of Metals

ACTIVITY 5.5

Illustrating Conductivity of Heat and Electricity of Metals

1. Take an aluminium or copper wire. Clamp this wire on a stand, as shown in Fig. 5.12.
2. Fix a pin to the free end of the wire using wax.
3. Heat the wire with a spirit lamp, candle or a burner near the place where it is clamped.
   What do you observe after some time?
4. Repeat the same with carbon or sulphur.
5. Note your observations.
   Does the element melt?

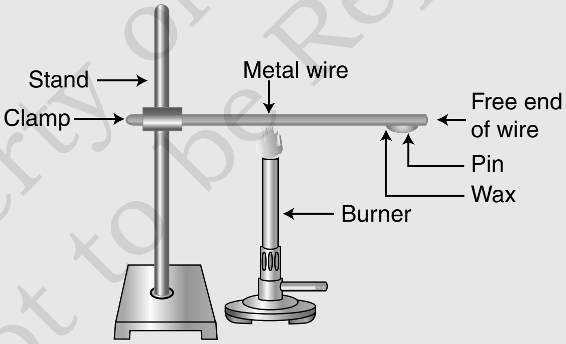


Fig. 5.12.

6. Consider elements aluminium, copper, sulphur and carbon.
7. Set up an electric circuit as shown in Fig. 5.13.
8. Place the element to be tested in the circuit between terminals A and B as shown.
   Does the bulb glow? What does this indicate?
9. Compile your observations regarding properties of elements in your exercise book.

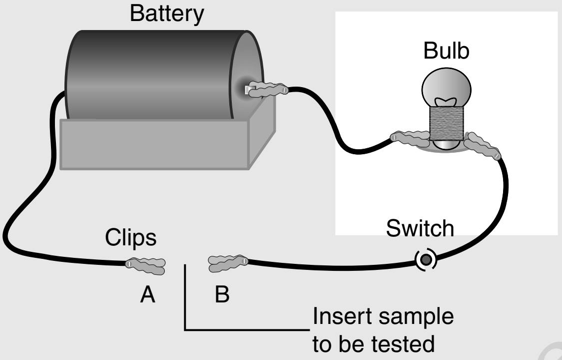


Fig. 5.13.

- Metals are good conductors of heat and electricity. This means that metals allow heat and electricity to pass through them easily. Silver metal is the best conductor of heat. Copper metal is a better conductor of heat than aluminium metal.
- Metals are lustrous (or shiny). This means that metals have a shiny appearance.
- Metals are usually **strong**. For example, iron metal (in the form of steel) is very strong when freshly cut and is used in the construction of bridges, buildings and vehicles. Some metals are not strong. For example, sodium and potassium.
- Metals are **ductile**. This means that metals can be drawn (or stretched) into thin wires.
- Gold and silver are among the best ductile metals.
- Metals are **malleable**. This means that metals can be hammered into thin sheets.

5.5. INTERMOLECULAR BONDING

We know that in covalent molecules, the constituent atoms are held by the covalent bonds. These molecules are further held by some attractive forces in different chemical entities. For example, H_2O molecule is a descrete unit which contains two O–H covalent bonds. But a sample of water contains large number H_2O units held by some attractive forces. *The attractive forces operating between the molecules of a given substances* (*solid, liquid or gas*) *are called **intermolecular forces**.*

The different types of intermolecular forces among the covalent molecular are

1. van der Waal forces or London dispersion forces.
2. Dipole-Dipole interactions
3. Hydrogen Bonds.

The magnitude of intremolecular forces control most of the physical properties of the substance such as *melting point, boiling point, viscosity, surface tension, molar heat of vaporisation, extent of evaporation*, etc. Higher the magnitude of the interparticle forces, higher will be boiling point. In many cases they also affect the chemical properties of the substances.

Intremolecular forces on the other hand, mostly control the chemical properties of the substance.

Let us now study the origin and nature of various intermolecular forces.

5.5.1. London Dispersion Forces or van der Waal Forces

These are the intermolecular forces among the **monoatomic** or **nonpolar molecules** such as N_2, He, H_2, CO_2, etc. Non-polar molecules are electrically symmetrical because their electronic charge cloud is symmetrically distributed. However, an **instantaneous dipole** may develop in such molecules. Let us understand this by assuming two atoms of neon 'A' and 'B' in the close vicinity of each other (Fig. 5.14 (a)).

Rapid movement of electrons in A may cause its temporary polarization due to momentary accumulation of electron density on one side and making the charge distribution unsymmetrical (Figs. 5.14 (b) and (c)).

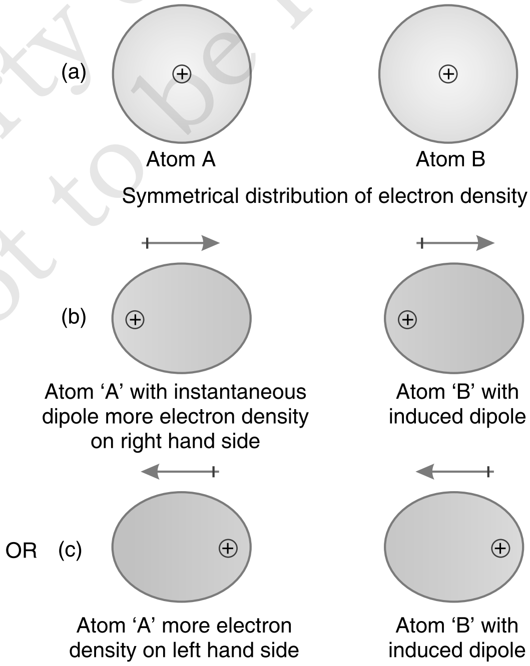


Fig. 5.14. Dispersion forces or London forces between atoms

This will result in the development of temporary instantaneous dipole on the atom 'A' for a very short time. This instantaneous or transient dipole distorts the electron density of the atom 'B', which is close to it. In other words, a dipole is induced in the atom 'B' also.

The **temporary dipoles** of atom 'A' and 'B' attract each other. Similar temporary dipoles are induced in non-polar molecules also. Magnitude of such a force of attraction was first calculated by the German physicist **Fritz London**. For this reason, force of attraction between two temporary dipoles is known as **London force**. Another name for this force is **dispersion force**. These forces are always attractive and interaction energy is proportional to $\left(\frac{1}{r^6}\right)$ where r is the distance between two interacting particles. These forces are important only at short distances (~ 500 pm) and their magnitude depends on the ability of the particles to undergo **polarization**.

5.5.2. Dipole-Dipole Forces

Dipole-dipole forces are *attractive forces operating between the polar molecules possessing permanent dipole (μ > 0).* Ends of the dipole possess "*partial charges*" and these charges are shown by Greek letter delta (δ). Partial charges are always less than the unit charge (1.6×10^{-19} C) because of electron sharing effect. The polar molecules interact with neighbouring molecules through their oppositely charged ends. This interaction is weak as compared to ion-ion interactions. The interaction energy decreases with the increase of distance between the dipoles. Now, greater the dipole moment of polar molecule, greater is the magnitude of dipole-dipole forces. Figure 5.15 shows electron cloud distribution and attractive interactions between H—Cl dipoles.

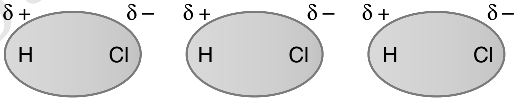


Fig. 5.15. Attractive interactions between H—Cl dipoles

In general, for the molecules of similar molecular masses, the magnitude of dipole-dipole forces is higher in case of more polar molecules. For example, molecular masses of PH_3 and H_2S is same but H_2S is more polar (μ = 1.10D) than PH_3 (μ = 0.55D). Hence boiling point of H_2S (187K) is higher than that of PH_3 (137 K).

5.5.3. Hydrogen Bond

When a H-atom is bonded to a highly electronegative atom like F, O or N (say, Z) by a covalent bond, the bond pair of electrons is displaced towards the electronegative atom. When solitary electron of hydrogen atom lies away from it, its nucleus gets exposed and behaves as a bare proton. Such a bare hydrogen nucleus exerts a strong electrostatic attraction on the electronegative atom of the adjacent molecule. This

$$\ldots\ldots\overset{\delta^+}{H}—\overset{\delta^-}{Z}\ldots\ldots\overset{\delta^+}{H}—\overset{\delta^-}{Z}\ldots\ldots\overset{\delta^+}{H}—\overset{\delta^-}{Z}\ldots\ldots$$

interaction between hydrogen atom of one molecule and the electronegative atom of the other molecule is referred to as **hydrogen bond**. Thus, **hydrogen bond** *is defined as the electrostatic force of attraction which exists between the covalently bonded hydrogen atom of one molecule and the electronegative atom of the other molecule*. The hydrogen bond is represented by dotted line (.......). For example, in case of hydrogen fluoride the hydrogen bond exists between H atom of one molecule and fluorine atom of another molecule as shown.

$$\ldots\ldots H—F\ldots\ldots H—F\ldots\ldots H—F\ldots\ldots$$

It may be noted that the hydrogen atom acts as a bridge between the electronegative atoms of two different molecules, to one atom it is linked through a covalent bond while to the other it is linked through a hydrogen bond. However, the hydrogen atom does not lie in the centres of the two electronegative atoms because hydrogen bond and covalent bond do not have same strength.

5.5.3.1. Examples of Hydrogen Bonding

Let us study some examples of the compounds which exhibit hydrogen bonding.

(i) **Hydrogen fluoride.** In hydrogen fluoride, hydrogen atom is bonded to highly electronegative atom fluorine (electronegativity = 4). It has been found that in solid state hydrogen fluoride consists of long zig-zag chains of H-F molecules associated by H-bonds as shown in Fig. 5.16 (a). On heating, progressively, the length of the chain shortens, and associated units become quite small.

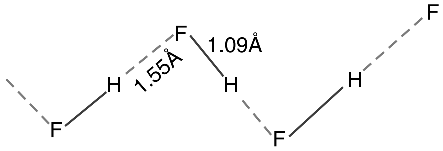


Fig. 5.16. (a) Zig-zag chains of H-F molecules

(ii) **Water.** In water molecule, oxygen atom is bonded to two hydrogen atoms. Due to large electronegativity, oxygen atom forms the negative centre whereas each of the hydrogen atom acquires a partial positive charge. Each O atom can form two hydrogen bonds as shown in Fig. 5.16 (b).

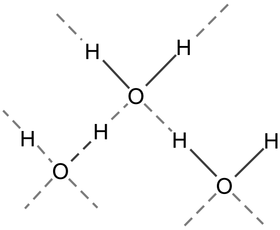


Fig. 5.16. (b) H-bonds in H_2O molecules

(iii) **Ammonia.** In ammonia molecule, nitrogen, an electronegative atom is bonded to three hydrogen atoms. The nitrogen atom forms a negative site of the molecule whereas each of three H-atoms acquires a partial positive charge. The ammonia molecules are associated by H-bonds as shown in Fig. 5.16 (c).

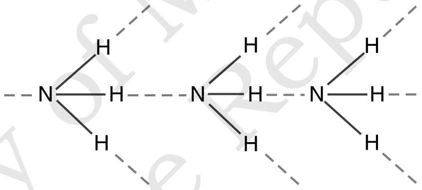
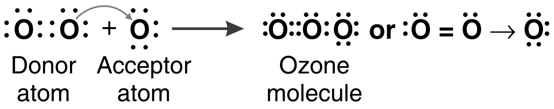


Fig. 5.16. (c) H-bonds in NH_3 molecules

5.6. COORDINATE COVALENT (DATIVE) BONDING

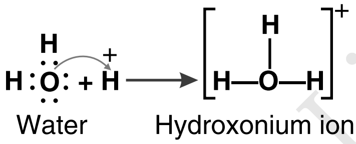
It is a special case of covalent bond the formation of which was postulated by ***Perkins*** (1921). It is formed by *mutual sharing of electrons between the two atoms but the shared pair of electrons is contributed only by one of the two atoms, the other atom simply participates in sharing.* The atom which donates an electron pair for sharing is called **donor** and it must have already completed its octet. On the other hand, the atom which accepts the electron pair in order to complete its octet is called **acceptor**. The bond is represented by an arrow (→) pointing from the donor towards the acceptor. Let us consider the formation of ozone molecule. A molecule of oxygen contains two oxygen atoms which share four electrons and complete their octets. In the formation of O_3 molecule, one of the oxygen atom of O_2 molecules acts as donor atom and shares two of its electrons with another oxygen atom. It can be represented as follows:



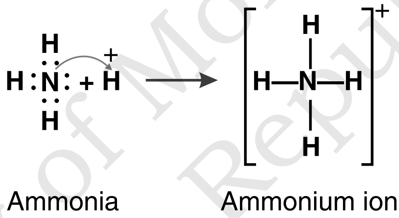
It is important to note that co-ordinate bond once formed, cannot be distinguished from covalent bond.

Some more examples of molecules/polyatomic ions having coordinate bond are as follows:

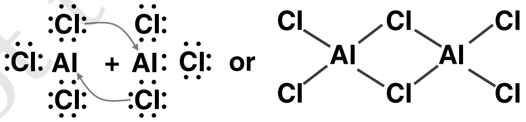
(i) **Hydroxonium ion.** Here, oxygen atom of H_2O molecule is donor while H^+ is acceptor.



(ii) **Ammonium ion.** Here, N atom of ammonia molecule is donor while H^+ is acceptor.

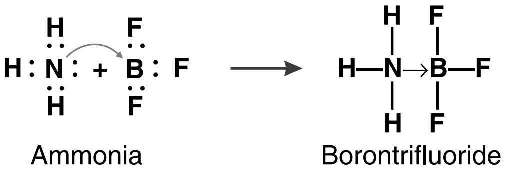


(iii) **Aluminium chloride.** Aluminium chloride exists in the form of dimer Al_2Cl_6. Here, chlorine atom of $AlCl_3$ unit acts as donor while electron deficient Al atom of the other $AlCl_3$ unit acts as acceptor as shown below.



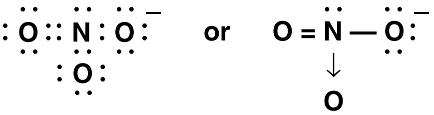
Formation of Al_2Cl_6

(iv) **Combination of borontrifluoride and ammonia.** Boron atom of boron trifluoride being electron deficient acts as acceptor while nitrogen atom of ammonia molecule is donor. Their combination by donor acceptor mechanism is shown on follows.



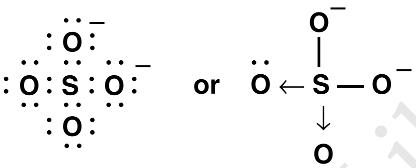
(v) **Lewis structures of nitrate (NO_3^-) and sulphate (SO_4^{2-}) ions**

Nitrate ion (NO_3^-)



(Nitrate ion)

Sulphate ion (SO_4^{2-})



(Sulphate ion)

ACTIVITY 5.6

In small groups, discuss the formation of covalent bond, dative bond and write the Lewis structures of simple covalent molecules/poly atomic ions. Watch video clips about formation of covalent and coordinate bond.

Make presentations by writing Lewis structures of the following molecules/ions.

- Water
- Sulphate ion
- Hydrogen peroxide
- Hydrogen chloride
- Carbon dioxide
- Carbon monoxide
- Methane
- Boron trifluoride
- Ammonia
- Nitrate ion
- Sulphur hexafluoride
- Ammonium ion

5.6.1. Comparison of All Bond Types

A comparison chart of bond types is summarized in Table 5.3.

Table 5.3. Comparison of All Bond Types

	Ionic	Covalent	Metallic
State at room temperature	Crystalline	Liquid, gas or low-melting solid	Solid state (Exception in mercury)
Occurs between	A metal and a non-metal	Two non-metals	Two and more metals
Character of bond	Non-directional	Directional	Non-directional

Conductivity	Conducts electricity when molten and dissolves in water	Do not conduct electricity	Good conductor of electricity
Solubility	Soluble in water but insoluble in organic solvents like petrol and kerosene.	Insoluble in water, but soluble in organic solvents like petrol and kerosene.	Insoluble in water or organic solvents, unless they undergo a chemical reaction with them.
Melting Point	High	Low to moderate	Low
Boiling Point	High	Low to moderate	Low
Ductility	Non-ductile	Non-ductile	Ductile
Malleability	Non-malleable	Non-malleable	Malleable
Examples	NaCl, $MsCl_2$, CaO, etc.	CH_4, NH_3, CO_2, etc.	Na, Mg, Al, etc.

GLOSSARY

- **Anion:** a negatively charged ion.
- **Boiling Point:** The temperature at which vapour pressure of a liquid becomes equal to the atmospheric pressure.
- **Cation:** a positively charged ion.
- **Covalent bonding:** A chemical bond formed when electrons are shared between two atoms.
- **Dative Bond** or **Co-ordinate Bond:** The bond is formed by sharing of electrons in which the shared pair of electrons is contributed by one of the atom called DONOR while the other atom is called ACCEPTOR.
- **Dipole Moment (μ):** A vector quantity defined by the product of charge developed on any of the atom and distance between the atoms; creating a dipole.
- **Ductile:** Able to be drawn out into a thin wire.
- **Electronegativity:** It is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.
- **Hybridization:** The concept of mixing two atomic orbitals to give rise to a new type of hybridized orbitals.
- **Hydrogen Bond:** The electrostatic force of attraction between covalently bonded H-atom of one molecule and the electronegative atom (F or N or O) of the other molecule.

- **Intermolecular Forces:** The attractive forces which hold molecules of the substance together in different states of matter.
- **Ionic Bonding:** A chemical bond formed when one atom gives up one or more electrons to another atom.
- **Ionization Energy:** The energy required to remove the outermost electron from an isolated gaseous atom of the element.
- **Malleable:** Able to be hammered or pressed into shape without breaking or cracking.
- **Melting Point:** The temperature at which a given solid material changes from a solid state to a liquid, or melts.
- **Metallic Bonding:** A type of chemical bonding between two or more metal atoms.
- **Noble Gas:** The gaseous elements helium, neon, argon, krypton, xenon, and radon.
- **Van der Waals' Forces:** Dipole-dipole forces, dipole-induced dipole forces and dispersion forces are collectively called van der Waal forces.

SUMMARY

- The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a chemical bond.
- The compounds which are made up of ions are known as ionic compounds.
- The factors influencing the formation of ionic bonding are (i) ionization energy, (ii) electron affinity, and (iii) lattice energy.
- Ionic compounds are usually crystalline solids and have high melting and high boiling points.
- Hybridization is the phenomenon of intermixing of atomic orbitals of slightly different energies of the atom (by redistributing their energies) to form new set of orbitals of equivalent energies and identical shape.
- The chemical bond formed by sharing of electrons between two atoms is known as a covalent bond.
- The factors that influence the formation of covalent bonding are (i) electron affinity, (ii) ionization energy, (iii) atomic size, and (iii) electron egativity.
- Covalent compounds are usually liquids, gases or solids, and have low melting points and low boiling points.
- The force which binds various metal atoms together is called metallic bond.

- Degree of polarity in a bond is measured in terms of dipole moment.
- The VSEPR theory is a model used to predict 3-D molecular geometry based on the number of valence shell electron bond pairs among the atoms in a molecule or ion.
- Metals are good conductors of heat and electricity. They are strong, ductile and malleable.
- The attractive forces which hold together various molecules of the species in solid or liquid states are called intermolecular forces. The magnitude of these forces control the melting, boiling points of the substance.
- Different intermolecular forces are, van der Waal forces, hydrogen bonds, dipole-dipole forces and London dispersion forces.

EVALUATION

I. Multiple Choice Questions

1. Which of these is an ionic compound?

(a) NaCl
(b) H_2O
(c) NH_3
(d) CO_2

2. Which of the following is not a property of ionic compound?

(a) Ionic compounds are usually crystalline solids.
(b) Ionic compounds conduct electricity in solid-state.
(c) Ionic compounds are insoluble in organic solvents.
(d) Ionic compounds have high melting and boiling points.

3. Which type of hybridization occurs in methane?

(a) sp^3-hybridization
(b) sp^2-hybridization
(c) sp-hybridization
(d) None of these

4. Which type of hybridization occurs in beryllium chloride?

(a) sp^3-hybridization
(b) sp^2-hybridization
(c) sp-hybridization
(d) None of these

5. Which of these is NOT a covalent compound?

(a) Carbon dioxide
(b) Methane
(c) Ammonia
(d) None of these

6. Covalent bond occurs between:

(a) two metals
(b) two non-metals
(c) a metal and a non-metal
(d) none of these

7. This is the force which binds various metal atoms together.
   (a) ionic bond (b) covalent bond
   (c) metallic bond (d) none of these
8. Molecular geometry is explained primarily through _______ theory.
   (a) Valence Bond (b) VSEPR
   (c) Molecular Orbital (d) None of these
9. Which of the following is a directional bond?
   (a) Coordinate (b) Covalent
   (c) Metallic (d) All these
10. This type of bonding is formed by mutual sharing of electrons between the two atoms but the shared pair of electrons is contributed only by one of the two atoms, the other atom simply participates in sharing.
   (a) Coordinate (b) Covalent
   (c) Metallic (d) Coordinate Covalent

II. State True or False

1. Ionic compounds are insoluble in water.
2. Ionic bond involves a metal and a non-metal.
3. The ionic bond formed by sharing of electrons between two atoms.
4. Covalent compounds have usually low melting points and low boiling points.
5. Covalent compounds conduct electricity.
6. Molecular shape of NH_3 is trigonal pyramidal.
7. Metals are ductile.
8. Metals are not lustrous.
9. The magnitude of intermolecular forces control the melting, boiling points of the substance.
10. Dipole moment of carbon dioxide is 1.83D.

III. Answer the Following Questions

1. What do you mean by chemical bonding? What are its different type?
2. What is ionic bonding? Give at least two examples of ionic compounds.
3. List some factors that influence the formation of ionic bonding.
4. List some properties of ionic compounds.
5. Why do ionic compounds conduct electricity when dissolved in water?
6. What do you mean by hybridization?

7. What do you mean by covalent bonding? Give at least two examples of covalent compounds.
8. List the properties of covalent compounds.
9. Why do covalent compounds not conduct electricity?
10. List some factors that influence the formation of covalent bonding.
11. What do you mean by metallic bonding?
12. List some factors that influence the formation of metallic bonding.
13. Illustrate some physical properties of metals.
14. Give two examples each of polar and non-polar molecules.
15. What is a polar covalent bond? Give an example of a polar diatomic molecular. Draw its structure.
16. In a molecule AX3, there are 3 bond pairs and one lone pair around A. What shape will it adopt? Draw it.
17. What is Dative bond? How it is different from covalent bond?
18. Show how the coordinate covalent bond is formed between ammonia and the hydrogen ion.
19. Differentiate between ionic, covalent and metallic bonding.

PROJECT WORK

Make a chart to compare ionic covalent and metallic bond in terms of structure, heat and electrical conductivities, boiling & melting points, solubility in water and non-polar solvents and hardness. Submit it in the class as an assignment.

SEMESTER–II (Period–V)

C10CH6

6

TOPIC

Atoms, Molecules, Ions and Stoichiometry

Learning Objectives

Upon completion of this topic, learners will be able to:

- Distinguish the fundamental chemical laws
- Distinguish atoms, molecules and ions
- Discuss the mole concept
- Determine the percent of elements in compounds
- Determine the formula of a compound
- Discuss the kinds and types of chemical reactions
- Analyze the techniques in balancing chemical reactions and
- Determine the limiting reagent/reactant of a chemical reaction.

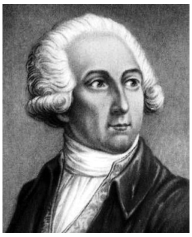
Introduction

The structure of matter has been a subject of speculation from very early times. In the fourth century BC, Greek philosopher, Democritus, suggested that if we go on dividing matter into smaller parts, a stage would be reached when particles obtained cannot be divided further. He called these particles '*atoms*' meaning indivisible. All these concepts were based on philosophical considerations and not much experimental work was done to confirm these ideas. After Democritus' death, little more was done with atomic theory until the end of the eighteenth century, when Antoine Lavoisier introduced modern chemistry to the world. He put forward two important laws of chemical combination which formed the basis of Dalton's atomic theory which was published in 1808.

6.1. FUNDAMENTAL CHEMICAL LAWS

By studying the results of quantitative measurements of many reactions it was observed that whenever substances react, they follow certain laws. These laws are called the **laws of chemical combination**. These laws formed the basis of Dalton's atomic theory of matter.

6.1.1. Law of Conservation of Mass



Antoine Lavoisier
(1743–1794)

This law was stated by the French chemist **Antonie Lavoisier** (1774). This law states that:

In every chemical reaction, the total mass before and after the reaction remains constant.

In other words, *mass can neither be created nor destroyed in a chemical reaction.*

Lavoisier showed that when mercuric oxide was heated, it produced free mercury and oxygen. The sum of masses of mercury and oxygen was found to be equal to the mass of mercuric oxide.

$$\begin{array}{ccccc} \text{Mercuric oxide} & \longrightarrow & \text{Mercury} & + & \text{Oxygen} \\ 100\text{ g} & & 92.6\text{ g} & & 7.4\text{ g} \end{array}$$

Law of conservation of mass is also known as *law of indestructibility of matter.*

ACTIVITY 6.1

Demonstration of Law of Conservation of Mass

1. Prepare separately a 5% solution of barium chloride and a 5% solution of sodium sulphate.
2. Take about 20 mL of barium chloride solution in a conical flask.
3. Take sodium sulphate solution in a small test tube. Hang the test tube in the conical flask with the help of a thread. Close the mouth of the flask with a cork (Fig. 6.1).

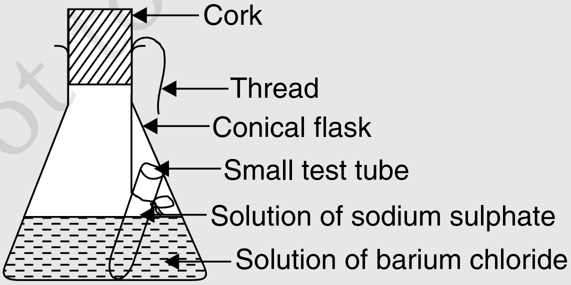


Fig. 6.1. Study of reaction between barium chloride and sodium sulphate.

4. Weigh the flask along with its contents.
5. Now tilt the flask so that the two solutions get mixed.
6. Weigh the flask again along with its contents.

   What do you observe?

It is observed that on mixing the two solutions a chemical reaction takes place which is indicated by the formation of a white precipitate.

$$\text{Barium chloride + Sodium sulphate} \longrightarrow \underset{\text{(White ppt.)}}{\text{Barium sulphate}} + \text{Sodium chloride}$$

The mass of the flask and its contents remains constant. Thus, during a chemical reaction mass is neither created nor destroyed.

This activity can also be carried out with the following pairs:

(a) Silver nitrate and sodium chloride.
(b) Copper sulphate and sodium carbonate.

Example 6.1: *8.4 g of magnesium carbonate on heating gave 4.4 g of carbon dioxide and 4.0 g of magnesium oxide. Show that these observations are in agreement with law of conservation of mass.*

Solution: Mass of the reactants = 8.4 g

Mass of the products = 4.4 + 4.0 = 8.4 g

Since the total mass of the products formed is equal to the total mass of the reactants undergoing reaction, so the data is in agreement with law of conservation of mass.

6.1.2. Law of Definite Proportion

This law deals with the composition of chemical compounds. It was discovered by the French chemists,

A. Lavoisier and Joseph Proust (1799). This law states that:

A pure chemical compound always contains same elements combined together in same proportion by mass.

It implies from this law that in a chemical compound the elements are present in *fixed* and not *arbitrary* ratio by mass. For example, pure water obtained from different sources such as *river, well, spring, sea,* etc., always contains hydrogen and oxygen combined together in the ratio 1 : 8 by mass. Similarly, carbon dioxide can be obtained by different methods such as:

(i) By burning of carbon,
(ii) By heating lime stone, or
(iii) By the action of dilute hydrochloric acid on marble pieces.

It can be shown experimentally that samples of carbon dioxide obtained from different sources contain carbon and oxygen in the ratio of 3 : 8 by mass.

Example 6.2: *A sample of ascorbic acid (vitamin C), synthesized in laboratory contains 15.0 g of carbon in 35.0 g of ascorbic acid. Another sample of ascorbic acid, isolated from lemons contains 42.9% carbon. Show that the data is in accordance with law of constant composition.*

Solution: In the first sample of ascorbic acid,

$$\text{Mass of carbon} = 15.0 \text{ g}$$

$$\text{Mass of ascorbic acid} = 35.0 \text{ g}$$

$$\% \text{ of carbon} = \frac{15.0}{35.0} \times 100 = 42.9\%$$

In the second sample of ascorbic acid,

$$\% \text{ of carbon} = 42.9\%$$

Since both the samples contain same percentage of carbon, the given data is in accordance with law of constant composition.

Example 6.3: *When 3.0 g of carbon is burnt in 8.00 g oxygen, 11.00 g of carbon dioxide is produced. What mass of carbon dioxide will be formed when 3.0 g of carbon is burnt in 50.00 g of oxygen? Which law of chemical combination will govern your answer?*

Solution: Carbon and oxygen combine in the ratio 3: 8 to form carbon dioxide. Even when we use excess of oxygen, carbon and oxygen would combine in the same ratio to form carbon dioxide. Excess oxygen would be left unreacted. Therefore, 3 g of carbon when burnt in 50 g of oxygen would give only 11.0 g of carbon dioxide by reacting with 8 g of oxygen.

The **law of constant composition** governs this answer.

6.1.3. Law of Multiple Proportions

This law was proposed by John Dalton (1803). This law states that:

When two elements combine with each other to form two or more than two compounds, the masses of one element which combine with fixed mass of the other, bear a simple whole number ratio to one another.

For example, carbon and oxygen combine with each other to form carbon monoxide (CO) and carbon dioxide (CO_2).

In carbon monoxide (CO), 12 parts by mass of carbon combine with 16 parts by mass of oxygen.

In carbon dioxide (CO_2), 12 parts by mass of carbon combine with 32 parts by mass of oxygen.

Ratio of the masses of the oxygen which combine with fixed mass of carbon (12 parts) in these compounds is 16 : 32 or 1 : 2, which is a simple whole number ratio.

Example 6.4: *Hydrogen and oxygen are known to form two compounds. The hydrogen content in one of these is 5.93% while in the other it is 11.2%. Show that this data illustrates the law of multiple proportions.*

Solution: *In the first compound*:

Hydrogen = 5.93%
Oxygen = (100 – 5.93)% = 94.07%.

In the second compound:

Hydrogen = 11.2%
Oxygen = (100 – 11.2)% = 88.8%.

In the first compound, the number of parts by mass of oxygen that combine with one part by mass of hydrogen = $\frac{94.07}{5.93}$ = 15.86 parts.

In the second compound, the number of parts by mass of oxygen that combine with one part by mass of hydrogen = $\frac{88.8}{11.2}$ = 7.9 parts.

The ratio of masses of oxygen that combine with fixed mass (1 part) by mass of hydrogen is 15.86 : 7.9 or 2 : 1.

Since this ratio is a simple whole number ratio, hence the given data illustrates the law of multiple proportions.

6.2. ATOMS, MOLECULES AND IONS

6.2.1. Atoms

Atoms are the building blocks of all matter. An atom is the smallest particle of an element which can ever exist. A chemical reaction involves redistribution of atoms among various species. An atom is neither created nor destroyed during chemical reactions. In other words, in an ordinary chemical reaction, no atom of any element disappears or is changed into an atom of another element. *It is the smallest particle of an element that can take part in chemical reactions.*

Atoms are very small in size. The size of an atom is expressed in terms of atomic radius. Atomic radius is measured in nanometres (nm).

$$1 \text{ nm} = \frac{1}{10^9} \text{ m} = 10^{-9} \text{ m}$$

Radii of most of the atoms are of the order of 0.1 nm or 10^{-10} m. For example, atomic radius of hydrogen is 0.037 nm while that of gold atom is 0.144 nm.

Atoms cannot be viewed by simple optical microscopes. However, through modern techniques such as scanning tunneling microscope (STM) it is possible to produce magnified images of surfaces of elements showing atoms.

6.2.2. Molecules

As already mentioned an atom is the smallest particle of an element that takes part in chemical reactions. An atom may or may not have independent existence. The smallest particle of an element or a compound which can exist independently is called a **molecule**. Thus,

*A **molecule** is the smallest particle of an element or compound that has independent existence.* A molecule may contain one or more than one atoms.

6.2.2.1. Molecules of Elements

The molecules of elements contain atoms of only one kind.

The number of atoms in a molecule of an element is known as **atomicity of the element**.

Noble gases (He, Ne, Ar, Kr and Xe) have monoatomic molecules.

Hydrogen (H_2), oxygen (O_2), nitrogen (N_2), chlorine (Cl_2), etc., have diatomic molecules. Ozone (O_3) has triatomic molecule.

For phosphorus (P_4) atomicity is four and for sulphur (S_8) atomicity is eight.

Many elements such as carbon, silicon and metals have complex molecular structures consisting of a very large, indefinite number of atoms. These elements are represented by their atomic symbols.

6.2.2.2. Molecules of Compounds

Molecules of compounds contain atoms of two or more different elements. For example, a molecule of ammonia (NH_3), a molecule of methane (CH_4), a molecule of glucose ($C_6H_{12}O_6$).

The molecular formula of a substance tells us about the number of atoms of various elements present in a molecule of the substance.

6.2.3. Ions (or Radicals)

Most of the inorganic compounds which are composed of metals and non-metals, contain oppositely charged parts called ions or radicals. The positively charged part is called **cation** or **electropositive radical** or **basic radical** whereas negatively charged part is called **anion** or **electronegative radical** or **acidic radical**. For example, sodium chloride, is made up of Na^+ ions and Cl^- ions similarly nickel sulphate is made up of Ni^{2+} ions and SO_4^{2-} ions. An ion behaves as single unit in reactions. Thus:

*An **ion** is an atom or group of atoms, carrying positive or negative charge, that behaves as a single unit in reactions.*

The positively charged ion is called **cation**. Whereas the negatively charged ion is called **anion**.

An ion containing only one atom is known as **simple ion** whereas an ion containing two or more than two atoms is known as **polyatomic ion** or **compound ion**. Thus, Na^+ is a simple ion whereas SO_4^{2-} is a compound ion.

The charge on the ion is known as valency of the ion.

6.3. THE MOLE CONCEPT

6.3.1. Mole and Particles

While carrying out investigations we are often interested in knowing the number of atoms and molecules we are dealing with. Sometimes, we have to take the atoms or molecules of different reactants in a definite ratio. For example, let us consider the following reaction:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

In this reaction, one molecule of nitrogen reacts with three molecules of hydrogen. So it would be desirable to take the molecules of nitrogen and hydrogen in the ratio 1 : 3, so that the reactants are completely consumed during the reaction. But we know atoms and molecules are so small in size that it is not possible to count them individually. Moreover, the smallest sample of the substance that we may take contains a very large number of atoms or molecules. In order to deal with such huge numbers a unit of similar magnitude is required. In SI system, a unit

mole was introduced to count entities, at microscopic level, such as atoms, molecules, ions, etc. According to this concept, the number of particles of the substance is related to mass of the substance. For the sake of counting elementary particles the number of atoms in 12 g of carbon-12 isotope has been taken as standard. This number has been experimentally found to be 6.0221367×10^{23} and is called **one mole**. The **mole** may be defined as ***"the amount of the substance that contains as many specified elementary particles as the number of atoms in 12 g of carbon-12 isotope."***

The elementary particles may be atoms, molecules or other discrete particles such as ions, electrons, etc. Mole is simply a unit for counting atoms, molecules, ions, etc. It stands for 6.022×10^{23} particles irrespective of their nature, just as we use one dozen for twelve objects and one score for twenty items irrespective of their nature.

6.3.2. Avogadro Number

The number of entities in one mole is so significant that it has been given a separate name and symbol. It is known as **Avogadro number** or **Avogadro constant** and is denoted as $\mathbf{N_A}$.

$$\boxed{\textbf{One Mole} = 6.022 \times 10^{23} \text{ or } N_A \textbf{ Particles}}$$

- 1 mole oxygen atoms = 6.022×10^{23} O atoms
- 1 mole oxygen molecules = 6.022×10^{23} O_2 molecules
- 1 mole chloride ions = 6.022×10^{23} Cl^- ions
- 1 mole electrons = 6.022×10^{23} electrons

Now, if we multiply the number moles (**n**) of the substance with Avogadro number ($\mathbf{N_A}$) we get total number of particles (**N**) in the given amount of the substance.

$$\boxed{\mathbf{N = n \times N_A}}$$

6.3.3. Molar Mass

When we deal with very large number of particles, it is easier to weigh them instead of counting them. Therefore, it is more appropriate to relate the mass of the substance to the number of particles.

It can be easily shown that if we take weights of two elements in the ratio of their atomic masses then they will have equal numbers of atoms. Suppose we take 12.0 g of carbon and 16.0 g of oxygen separately then both the samples will have same number of atoms because the ratio of their weights is same as the ratio of their atomic masses. Now we know that 12.0 g of carbon contains 6.022 × 10^{23} atoms, therefore, 16.0 g of oxygen will also contain 6.022 × 10^{23} atoms of oxygen.

Similarly, 32.0 g of oxygen gas contains 6.022 × 10^{23} oxygen (O_2) molecules and 17.00 g of ammonia will contain 6.022 × 10^{23} ammonia (NH_3) molecules. From this we can conclude that mass of element in grams which is numerically equal to its atomic mass contains 6.022 × 10^{23} (N_A) atoms and the mass of substance in grams which is numerically equal to its molecular mass contains 6.022 × 10^{23} (N_A) molecules. This mass is referred to as the *molar mass*. **Molar mass** of the substance may ***thus, be defined as the mass of one mole, i.e., 6.022 × 10^{23} particles of that substance***. Molar mass is represented by **M** and is expressed in the units **g mol^{-1}**. Molar mass in grams is numerically equal to the atomic mass/molecular mass/formula mass expressed in unified mass units (u).

Examples:

- **Atomic oxygen:** Atomic mass = 16 u

  Molar mass of oxygen atoms = **16 g mol^{-1}** and it contains 6.022 × 10^{23} O atoms
- **Water (H_2O):** Molecular mass = 18 u

  Molar mass of water = **18 g mol^{-1}** and it contains 6.022 × 10^{23} H_2O molecules.
- **Sodium chloride (NaCl):**

  Formula mass = 58.5 u

  Molar mass of NaCl = **58.5 g mol^{-1}** and it contains 6.022 × 10^{23} NaCl units.

In short, the relationship between moles, mass and number of particles can be summed up as follows:

Moles	Number of Particles	Mass (g)
1.0	N_A or 6.022 × 10^{23}	M (molar mass)
n	$n \times N_A$	$n \times M$

6.3.4. Mole and Volume

From Avogadro's Law it is known that equal volumes of all gases contain equal number of molecules under similar conditions of temperature and pressure. Conversely, we can say that different samples of gases containing equal number of molecules occupy same volume under similar conditions. We know one mole molecules of all gases contain same number (6.022×10^{23}) of molecules, therefore, they occupy same volume under similar conditions of temperature and pressure. The volume occupied by **one mole** molecules of a gaseous substance is called **Molar Volume**. ***One mole molecules of all gases occupy 22.4 litres at 273 K and 1 atm pressure (S.T.P.)***. Hence, molar volume of all gases at S.T.P. is 22.4 litres.

Thus, for the gaseous substances at S.T.P. (273.15 K, 1 atm) conditions,

$$\text{One mole} = 22.4 \text{ L (Molar volume)}$$

It may be noted that 1 mole of gas at S.T.P. conditions (273.15 K, 1 bar) occupies a volume of 22.7 litres.

Remember

The following formulae are quite helpful in solving the problems on mole concept. In these formulae, N_A represents $\mathbf{6.022 \times 10^{23}}$.

1. Mass of **1 atom** of element

$$= \frac{\textbf{Molar mass of atoms}}{\mathbf{N_A}}$$

2. Mass of **1 molecule** of substance $= \dfrac{\textbf{Molar mass}}{\mathbf{N_A}}$
3. **Number of molecules** in **n** moles of substance = $\mathbf{n \times N_A}$
4. **Mass of n moles = n × Molar mass**
5. **Number of moles** in Wg of a substance $= \dfrac{\mathbf{W}}{\textbf{Molar Mass}}$
6. 1 **u** or (1 **amu**) = $\mathbf{1/N_A}$ **g**.
   It is also called **1 avogram** or **1 dalton**.

Summary of various relationships of mole has been illustrated in Fig. 6.2.

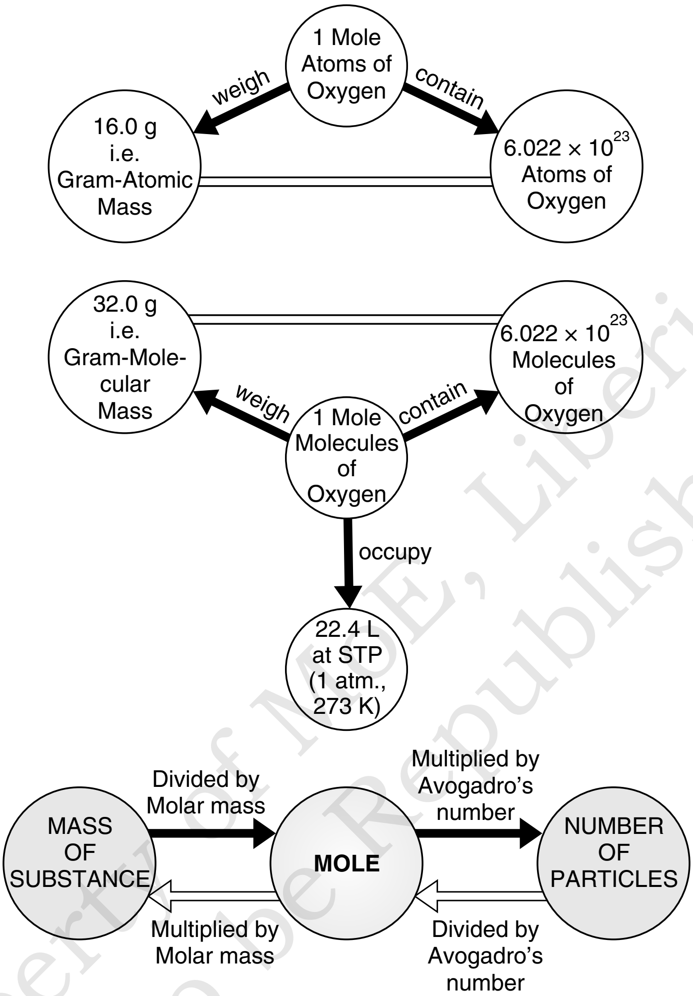


Fig. 6.2. Relationship between mole, number of particles and mass and interconversion of one into the other.

Example 6.5: *(a) Calculate the mass of 2.5 moles of calcium. Atomic mass of calcium is 40. (b) Calculate the mass of 1.5 mole of water (H_2O).*

Solution:

(a) Mass of 1 mole of calcium

$$= \text{molar mass of calcium} = 40 \text{ g mol}^{-1}$$

Mass of 2.5 moles of calcium

$$= 40 \times 2.5 = \mathbf{100\ g}.$$

(b) Molecular mass of water (H_2O)

$$= 1 \times 2 + 16 = 18\ u$$

$\therefore$ Molar mass of H_2O = 18 g mol^{-1}.

Mass of 1.5 mole of H_2O

$$= 1.5 \times 18 = \mathbf{27\ g.}$$

Example 6.6: *Calculate the molar mass of glucose ($C_6H_{12}O_6$).*

Solution: Molecular mass of glucose ($C_6H_{12}O_6$)

$$= 6 \times (12.011\ u) + 12 \times (1.008\ u) + 6 \times (16.00\ u)$$

$$= 72.066\ u + 12.096\ u + 96.00\ u$$

$$= \mathbf{180.162\ u}$$

$\therefore$ Molar mass of glucose = 180.162 g mol^{-1}

Example 6.7: *Calculate mass of the following:*

(i) *one atom of calcium*

(ii) *one molecule of sulphur dioxide (SO_2).*

Solution:

(i) Mass of 6.022×10^{23} atoms of calcium

= gram atomic mass of calcium = 40 g

$\therefore$ Mass of 1 atom of calcium

$$= \frac{40\,\text{g}}{6.022 \times 10^{23}} = \mathbf{6.6 \times 10^{-23}\ g.}$$

(ii) Mass of 6.022×10^{23} molecules of SO_2

= molar mass of SO_2 = 64 g.

$\therefore$ Mass of 1 molecule of SO_2

$$= \frac{64\,\text{g}}{6.022 \times 10^{23}} = \mathbf{1.06 \times 10^{-22}\ g}.$$

Example 6.8: *How many atoms of oxygen are present in 300 g of $CaCO_3$?*

Solution: Molar mass of $CaCO_3$ = 100 g

Now, 1 mole of $CaCO_3$ contains

= 3 mole of oxygen atoms.

or 100 g of $CaCO_3$ contain

= $3 \times 6.022 \times 10^{23}$ oxygen atoms

∴ 300 g of $CaCO_3$ contain oxygen atoms

$$= \frac{3 \times 6.022 \times 10^{23}}{100} \times 300$$

$$= \mathbf{5.4198 \times 1024\ oxygen\ atoms}.$$

6.4. PERCENTAGE COMPOSITION

Percentage composition of a compound refers to the amount of various constituent elements present per 100 parts by mass of the substance. It can be calculated from the formula of the compound.

Knowing the molecular formula of a compound, we can calculate its percentage composition by mass. To do this, we calculate the molecular mass of the compound. From this we can find out mass of one mole of the compound, which is equal to its gram molecular mass. Then we calculate mass of each element in one mole of the compound. The mass percentage of each element is then calculated by the following formula:

Mass per cent of the element X

$$= \frac{\text{Mass of X in one mole}}{\text{Molar mass of the compound}} \times 100$$

Similarly, mass percentages of other elements can be calculated as illustrated in the solved examples given below:

Example 6.9: *Calculate the mass per cent of different elements present in ethanol (C_2H_5OH).*

Solution:

Molar mass of ethanol $= 2 \times 12.01 + 6 \times 1.008 + 16.00$

$= 46.068$ g

Mass per cent of carbon $= \frac{24.02}{46.068} \times 100 = \mathbf{52.14\%}$

Mass per cent of hydrogen $= \frac{6.048}{46.068} \times 100 = \mathbf{13.13\%}$

Mass per cent of oxygen $= \frac{16.00}{46.068} \times 100 = \mathbf{34.73\%}$.

6.5. DETERMINATION OF FORMULA OF A COMPOUND

The **molecular formula** of a compound may be defined as the *formula which gives the actual number of atoms of various elements present in the molecule of the compound*. For example, the molecular formula of the compound glucose can be represented as $C_6H_{12}O_6$. A molecule of glucose contains six atoms of carbon, twelve atoms of hydrogen and six atoms of oxygen.

In order to find out molecular formula of a compound, the first step is to determine its empirical formula from the percentage composition.

The **empirical formula** of a compound may be defined as *the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound*. The empirical formula of the compound glucose ($C_6H_{12}O_6$), is CH_2O which shows that C, H and O are present in the simplest ratio of 1 : 2 : 1.

Empirical formula mass of substance is equal to the sum of atomic masses of all the atoms in the empirical formula of the substance.

6.5.1. Relation between the Two Formulae

Molecular formula is whole number multiple of empirical formula. Thus,

$$\text{Molecular formula} = \text{Empirical formula} \times n$$

where $n = 1, 2, 3\ldots$

$$n = \frac{\text{Molecular formula}}{\text{Empirical formula}} = \frac{\text{Molecular mass}}{\text{Empirical formula mass}}$$

6.5.2. Steps for Writing the Empirical Formula

The percentage of the elements in the compound is determined by suitable methods and from the data collected, the empirical formula is determined by the following steps:

(i) Divide the percentage of each element by its atomic mass. This will give the relative number of moles of various elements present in the compound.

(ii) Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.

(iii) Multiply the figures, so obtained by a suitable integer if necessary in order to obtain whole number ratio.

(iv) Finally write down the symbols of the various elements side by side and put the above numbers as the subscripts to the lower right hand corner of each symbol. This will represent the **empirical formula** of the compound.

6.5.3. Steps for Writing the Molecular Formula

The following steps can be followed while writing the molecular formula:

(i) Calculate the empirical formula as described above.

(ii) Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of the compound.

(iii) Divide the molecular mass (determined experimentally by some suitable method) by the empirical formula mass and find out the value of *n*.

(iv) Multiply the empirical formula of the compound with n so as to get the **molecular formula** of the compound.

The determination of the empirical formula and molecular formula of a compound involving the above steps is illustrated by the following examples.

Example 6.10: *Write the empirical formula for each of the compounds having molecular formulae*:

(i) C_6H_6 (ii) C_6H_{12} (iii) H_2O_2
(iv) H_2O (v) Na_2CO_3 (vi) B_2H_6
(vii) N_2O_4 (viii) H_3PO_4 (ix) Fe_2O_3
(x) C_2H_2.

Solution: Empirical formula is the simplest whole number ratio of atoms of different elements in the molecule. Therefore, the empirical formula of given compounds are:

(i) CH (ii) $\mathrm{CH_2}$ (iii) HO
(iv) $\mathrm{H_2O}$ (v) $\mathrm{Na_2CO_3}$ (vi) $\mathrm{BH_3}$
(vii) $\mathrm{NO_2}$ (viii) $\mathrm{H_3PO_4}$ (ix) $\mathrm{Fe_2O_3}$
(x) CH.

Example 6.11: *What is the simplest formula of the compound which has the following percentage composition: Carbon 80%, Hydrogen 20%? If the molecular mass is 30, calculate its molecular formula.*

Solution: Calculation of empirical formula:

Element	*Percentage*	*At. mass*	*Relative no. of moles*	*Simple ratio*	*Simplest whole no. ratio*
C	80	12	$\frac{18}{12} = 6.66$	$\frac{6.66}{6.66} = 1$	1
H	20	1	$\frac{20}{1} = 20$	$\frac{20}{6.66} = 3$	3

Hence, the empirical formula is CH_3.

6.6. KINDS OF CHEMICAL REACTIONS

Generally, chemical reactions are of two types: (i) Reversible reaction and (ii) Irreversible reaction. Let us discuss these in detail.

6.6.1. Reversible Reaction

*The reactions in which products under suitable conditions react to give back reactants are called **reversible reactions**.*

A reversible reaction is represented by putting two arrows pointing in opposite directions between the formulae of the reactants and the products as shown below:

6.6.2. Irreversible Reaction

When a piece of sodium is dropped into water, a violent reaction occurs resulting in the formation of sodium hydroxide and hydrogen gas.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

However, it is not possible to carry out the reverse reaction under any known experimental conditions, i.e., reduction of aqueous sodium hydroxide by hydrogen to form sodium and water cannot be achieved.

*The reactions in which the products do not react under any condition to give back reactants are called **irreversible reactions**.*

6.7. TYPES OF REACTIONS

There are millions of known chemical reactions. Many of these chemical reactions have common aspects, so they can be grouped into specific classification. The majority of chemical reactions (not all) fall into the following major categories:

- Combination reactions
- Decomposition reactions
- Single replacement reactions
- Double displacement reactions
- Combustion reactions
- Oxidation reduction reactions

6.7.1. Composition/Combination Reactions

ACTIVITY 6.2

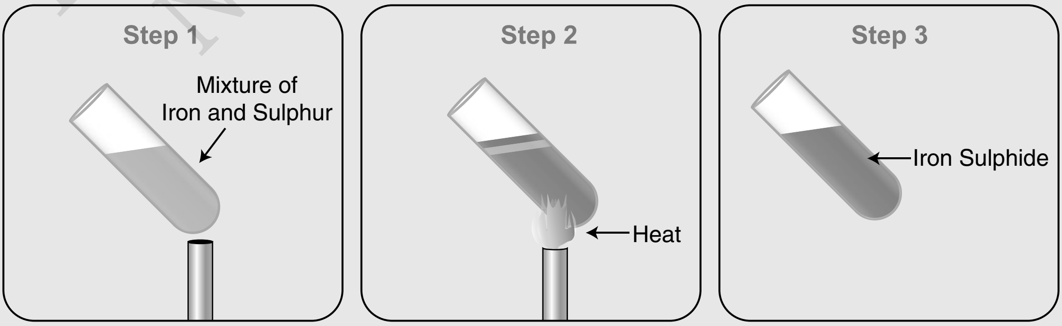
Combination of Iron and Sulphide (Combination Reaction)

Materials Required

Iron powder (filings) (7 grams), sulphur (4 grams), test tube, Bunsen burner, and pair of test tube tongs.

Procedure

1. Prepare a mixture containing iron powder and sulphur powder in the ratio 7:4 by mass.
2. Note the appearance of the pure elements and the mixture.
3. Take about 0.5 g of the mixture into a hard glass test tube.
4. Heat the mixture at the base of the test tube—gently at first and then more strongly (use a blue flame throughout).
5. Allow the test tube to cool down.
6. Observe the product.



Can you write the chemical equation for this reaction?

Safety

- It is advisable to wear protective gloves and heat the mixture using tongs.
- Eye protection required.

In this activity, you will observe that when we start heating a mixture of iron filings and sulphur, the sulphur melts and reacts with the iron filings to form the compound iron (II) sulphide. In this reaction, two different elements combine to form a single product. This is an example of **combination reaction**.

Combination reactions are those reactions in which a single product is produced from two (or more) reactants. The general equation for a combination reaction involving two reactants is

The reactants X and Y can be elements or compounds, or a compound and an element. The product XY is always a compound.

Combination reactions may involve:

- the combination of two elements to form a compound.
- the combination of a compound and an element to form a new compound.
- the combination of two compounds to form a new compound.

Example of some combination reactions are :

(i)	$Fe\ (s)$ Iron	+	$S\ (s)$ Sulphur	$\xrightarrow{\text{Combination}}$	$FeS\ (s)$ Iron sulphide
(ii)	$2Na(s)$ Sodium	+	$Cl_2(g)$ Chlorine	$\xrightarrow{\text{Combination}}$	$2NaCl(s)$ Sodium chloride
(iii)	$H_2\ (s)$ Hydrogen	+	$Cl_2(g)$ Chlorine	$\xrightarrow{\text{Combination}}$	$2HCl(s)$ Hydrogen chloride

6.7.2. Single Replacement Reactions

ACTIVITY 6.3

Reaction of Iron Nails with Copper Sulphate

1. Take three iron nails and clean them by rubbing with sand paper.
2. Take two test tubes marked as (A) and (B). In each test tube, take about 10 ml of copper sulphate solution.
3. Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes [Fig. 6.3 (a)]. (If the solution is concentrated, leave it for 2 hours) keep one iron nail aside for comparison.

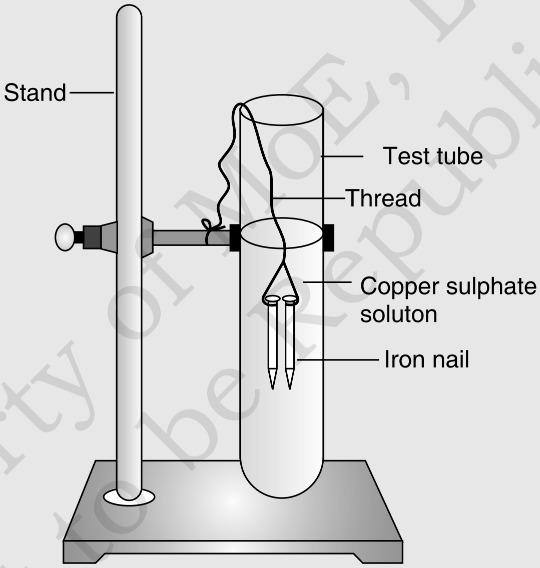


Fig. 6.3. (a) Iron nails dipped in copper sulphate solution

4. After 20 minutes, take out the iron nails from the copper sulphate solution.
5. Compare the intensity of the blue colour of copper sulphate solutions in test tubes (A) and (B), [Fig. 6.3 (b)].
6. Also, compare the colour of the iron nails dipped in the copper sulphate solution with the one kept aside [Fig. 6.3 (b)].

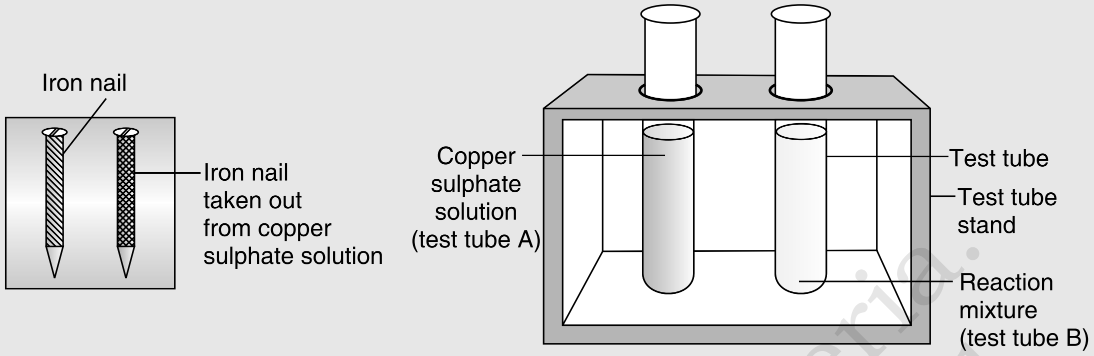


Fig. 6.3. (b) Iron nails and copper sulphate solutions compared before and after the experiment

Single replacement reactions are those reactions in which one element displaces another element from a compound. The general equation for this type of reaction is

$$X + YZ \longrightarrow XZ + Y$$

In reactants, X is an element and YZ is a compound. In products, XZ is a compound and Y is an element.

In single replacement reactions, more active metals displace less active metals (or hydrogen) from their compounds. Some common elements are arranged in decreasing order of their ability to replace element (metal ion) in aqueous solution. This series is known as activity series.

$\mathbf{K} > Na > Ca > Mg > Al > Zn > Cr > Fe > Ni > Sn > Pb > H > Cu > Ag > \mathbf{Au}$

Note: Potassium (K) is the *most reactive* metal and gold (Au) is the *least reactive* metal.

Example of some single replacement reactions are:

(i) $$\underset{\substack{\text{Copper sulphate}\\\text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Iron}\\\text{(Grey)}}}{Fe(g)} \longrightarrow \underset{\substack{\text{Iron sulphate}\\\text{(Greenish solution)}}}{FeSO_4(aq)} + \underset{\substack{\text{Copper}\\\text{(Red-brown)}}}{Cu\,(s)}$$

(ii) $$\underset{\substack{\text{Silver nitrate}\\\text{(Colourless solution)}}}{2AgNO_3\,(aq)} + \underset{\substack{\text{Copper}\\\text{(Red-brown)}}}{Cu(s)} \longrightarrow \underset{\substack{\text{Copper nitrate}\\\text{(Blue solution)}}}{Cu(NO_3)_2(aq)} + \underset{\substack{\text{Silver}\\\text{(Greyish white)}}}{2Ag\,(s)}$$

(iii) $$\underset{\substack{\text{Copper sulphate}\\\text{(Blue solution)}}}{CuSO_3\,(aq)} + \underset{\substack{\text{Zinc}\\\text{(Silvery-white)}}}{Zn(s)} \longrightarrow \underset{\substack{\text{Zinc sulphate}\\\text{(Colourless solution)}}}{ZnSO_4\,(aq)} + \underset{\substack{\text{Copper}\\\text{(Red-brown)}}}{Cu\,(s)}$$

6.7.3. Double Replacement Reactions

ACTIVITY 6.4

Formation of Precipitate

1. Take about 3 ml of sodium sulphate solution in a test tube.
2. In another test tube, take about 3 ml of barium chloride solution.
3. Mix the two solutions (Fig. 6.4).

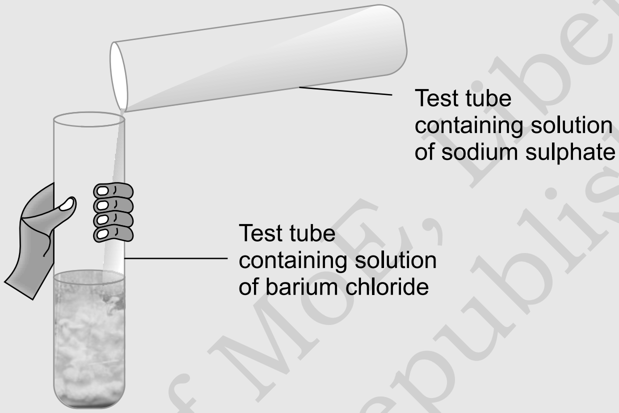


Fig. 6.4. Formation of Precipitate

What do you observe?

In Activity 6.4, you will observe that a white substance, which is insoluble in water, is formed.

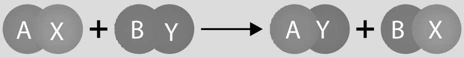
This insoluble substance formed is known as a *precipitate*. Any reaction that produces a precipitate can be called a ***precipitation reaction***.

Reaction:

$$\underset{\substack{\text{Sodium}\\\text{Sulphate}}}{Na_2SO_4\,(aq)} + \underset{\substack{\text{Barium}\\\text{chloride}}}{BaCl_2(aq)} \longrightarrow \underset{\substack{\text{Barium}\\\text{sulphate}}}{BaSO_4(s)} + \underset{\substack{\text{Sodium}\\\text{chloride}}}{2NaCl\,(aq)}$$

Double displacement reactions are those reactions in which two compounds react by exchange of ions to form two new compounds. These reactions are also known as **metathesis reactions** or **exchange reactions**.

The general equation for a double displacement reaction is.



where A and B are positive ions (cations).

X and Y are negative ions (anions).

Another example of double displacement reaction is when we mix a solution of silver nitrate ($AgNO_3$) and sodium chloride, silver chloride and sodium nitrate is formed.

$$AgNO_3\ (aq) + NaCl\ (aq) \longrightarrow AgCl\ (s) + NaNO_3\ (aq)$$

Double displacement reactions result in the removal of ions from the solutions.

Examples of some double replacement reactions are:

(i)	$BaCl_2(aq)$ Barium chloride	+	$CuSO_4(aq)$ Copper sulphate	$\longrightarrow$	$BaSO_4(s)$ Barium sulphate (White ppt.)	+	$CuCl_2(aq)$ Copper chloride
(ii)	$CuSO_4(aq)$ Copper sulphate	+	$H_2S\ (g)$ Hydrogen sulphide	$\longrightarrow$	$CuS(s)$ Copper sulphide (Black ppt.)	+	$H_2SO_2(aq)$ Sulphuric acid
(iii)	$Ba(OH)_2(aq)$ Base	+	$2HCl(aq)$ Acid	$\longrightarrow$	$BaCl_2(aq)$ Salt	+	$2H_2O(l)$ Water
(iv)	$2NaOH(aq)$ Base	+	$H_2SO_4(aq)$ Acid	$\longrightarrow$	$Na_2SO_4(aq)$ Salt	+	$H_2O(l)$ Water

6.7.4. Decomposition Reactions

ACTIVITY 6.5

Decomposition of Silver Chloride

1. Take about 2 g of silver chloride in a China dish.
2. Note the colour of silver chloride.

3. Place the China dish in sunlight for 10–30 minutes.
4. Observe the colour of silver chloride after 30 minutes.

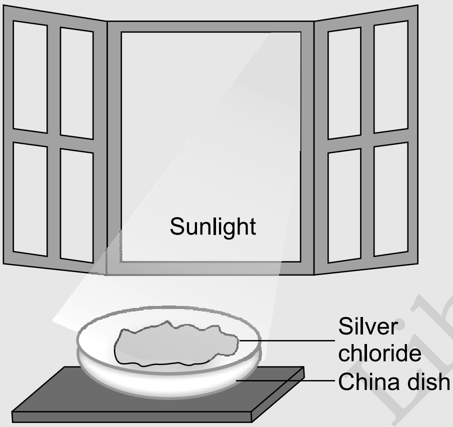


Fig. 6.5. Decomposition of $AgCl_2$

In Activity 6.5, you will observe that white silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.

Reaction:

$$2AgCl(s) \xrightarrow{\text{Sunlight}} 2Ag(s) + Cl_2(g)$$

Decomposition reactions are those reactions in which a single reactant breaks down into two (or more) simpler substances (elements or compounds). The general equation for a decomposition reaction in which there are two products, is

$$XY \longrightarrow X + Y$$

The reactant XY is always a compound. The products X and Y may be elements or compounds.

In other words, decomposition reactions are opposite of combination reactions. These reactions often involve an energy source such as heat, light, or electricity which breaks apart the bonds of compounds.

The products of decomposition reactions may be

- two elements
- one (or more) elements and one (or more) compounds
- two (or more) compounds.

Examples of some decomposition reactions are :

(i) $$\underset{\text{Water}}{2H_2O(l)} \xrightarrow[\text{Decomposition}]{\text{Electricity}} \underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

(ii) $$2H_2O_2(l) \xrightarrow{\text{light}} 2H_2O(l) + O_2(g)$$

(iii) $$\underset{\text{Hydrogen peroxide}}{2H_2O_2(l)} \xrightarrow{MnO_2} \underset{\text{Steam}}{2H_2O(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

6.7.5. Combustion Reactions

ACTIVITY 6.6

Burning of Magnesium Ribbon

1. Take a piece of magnesium ribbon and hold it with a pair of tongs. Light magnesium ribbon.
2. The magnesium ribbon starts burning with a dazzling flame.

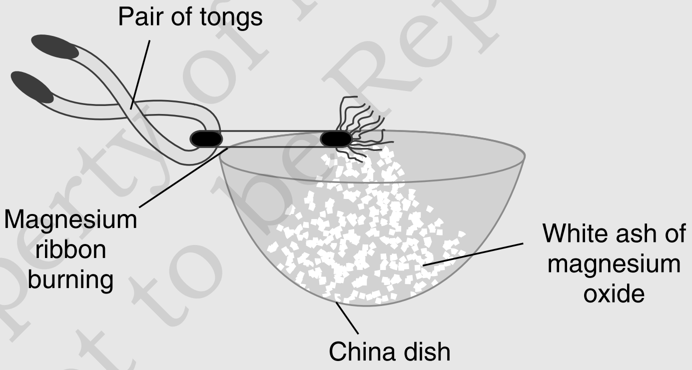


Fig. 6.6. Burning of magnesium ribbon in air

Explanation

You would observe that magnesium ribbon soon changes into white powder. This white powdery substance is magnesium oxide which is formed as a result of combination reaction.

Reaction

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

This type of reaction refers to the reaction of an element or compound with oxygen. Combustion usually releases a lot of heat energy. It is also referred to as *burning*.

Note: All combustion reactions are exothermic reactions.

6.7.5.1. Elements on Combustion

When elements undergo combustion, generally only one product is formed.

For example,

- Burning of coal

$$C\ (s) + O_2\ (g) \longrightarrow CO_2(g)$$

- Formation of water

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

- Burning of magnesium

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

6.7.5.2. Compounds on Combustion

When compounds undergo combustion, two or more products are formed. When carbon-hydrogen (hydrocarbon) or carbon-hydrogen-oxygen compounds undergo combustion in an excess of oxygen, the products are carbon dioxide and water. For example,

- Burning of natural gas

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

- Glucose reacts with oxygen

$$\underset{\text{(glucose)}}{C_6H_{12}O_6\ (aq)} + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

ACTIVITY 6.7

Oxidation and Reduction Reactions

1. Take 1g of copper powder in the China dish.
2. Put the tripod stand over the burner and the wire gauze on top of it.
3. Now put the China dish containing copper powder over the wire gauze.

4. Turn on the burner and heat the China dish.
5. Take the black copper(II) oxide powder and heat it again with hydrogen gas.

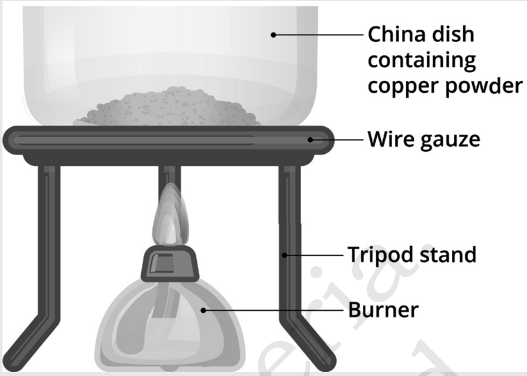


Fig. 6.7. Heating copper powder

Observation:

- After heating, we observed that the copper powder becomes coated with black colour.
- This black coloured powder is copper(II) oxide.
- The black powder of copper(II) oxide turns brown.

Result:

After heating, copper becomes copper oxide because of oxidation; when a chemical reaction gains oxygen, it is said to be oxidised.

The black powder of copper(II) oxide turns brown again, and we get copper powder back. This reaction happens because of the reduction. When there is a loss of oxygen in a chemical reaction, it is said to be reduced.

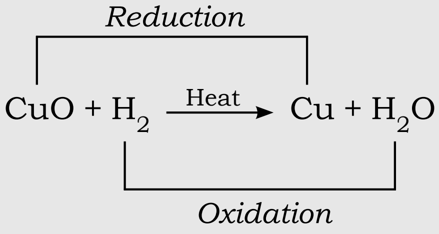
Reactions:

Oxidation reaction:

$$2\mathrm{Cu} + \mathrm{O_2} \xrightarrow{\triangle} 2\mathrm{CuO}$$

Redox reaction:

$$\mathrm{CuO} + \mathrm{H_2} \xrightarrow{\triangle} \mathrm{Cu} + \mathrm{H_2O}$$



$$\mathrm{CuO} + \mathrm{H_2} \xrightarrow{\text{Heat}} \mathrm{Cu} + \mathrm{H_2O}$$

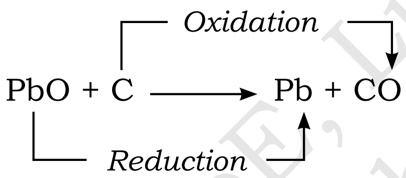
In this activity, the copper(II) oxide is loses oxygen during the reduction reaction. The copper(II) oxide is reduced to copper. Thus, ***reduction*** *is a process of losing oxygen.*

The hydrogen gains oxygen, and is being oxidised. Thus, ***oxidation*** *is the process of receiving oxygen.*

An oxidation-reduction reaction occurs when one reactant receives oxygen while the other reactant loses oxygen or when one reactant is oxidised while the other reactant is reduced at the same time. The oxidation-reduction reaction is also known as the redox reaction.

Here are some more examples of reduction reactions:

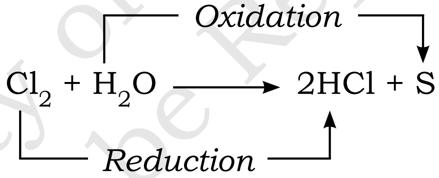
(i) *Reaction of PbO and carbon*:



$$PbO + C \longrightarrow Pb + CO$$

Here, oxygen is being removed from lead oxide (PbO) and is being added to carbon (C). Thus, PbO is reduced while C is oxidised.

(ii) *Reaction of* H_2S *and* Cl_2:



$$Cl_2 + H_2O \longrightarrow 2HCl + S$$

6.8. WRITING AND BALANCING CHEMICAL EQUATIONS

A chemical change can be represented by using symbols and formulae of various species involved in the change. Such a representation is known as chemical equation.

Representation of a chemical change in terms of symbols and formulae of the reactants and products is known as ***chemical equation*** *of the reaction.*

For example, when zinc carbonate is heated, zinc oxide and carbon dioxide are formed. This chemical change may be represented by the chemical equation given below:

$$ZnCO_3 \longrightarrow ZnO + CO_2$$

6.8.1. Writing Chemical Equation

Following are various steps of writing a chemical equation:

Step I: Writing Skeletal Equation

In order to write a chemical equation for a reaction, the first step is to write symbols and formulae of various reactants and products. For example, when zinc carbonate reacts with dilute hydrochloric acid it gives zinc chloride, carbon dioxide and water. This is represented by the following **'word equation'**.

In a word equation, we express the chemical change or chemical reaction by writing names of the reactants on left hand side and names of the products on right hand side. An arrow is put between reactants and products.

Zinc carbonate + Hydrochloric acid
$\longrightarrow$ Zinc chloride + Carbon dioxide + Water

Using symbols and formulae for various reactants and products we get,

$$ZnCO_3 + 2HCl \longrightarrow ZnCl_2 + CO_2 + H_2O$$

- Reactants are written on left hand side while products are written on right hand side.
- The symbols and formulae of various reacting substances (reactants) are separated by a plus sign (+) between them.
- The symbols and formulae of various substances formed (**products**) are also separated by a plus sign (+) between them.
- An arrow ($\longrightarrow$) is put between reactants and products.

Equations such as given above in which no attempt has been made to equalize the number of atoms of various elements on both the sides of the equation are known as **skeletal equations**. Thus:

*A **skeletal equation** is an equation in which various reactants and products are represented by their respective formulae but no attempt is made to equalize the number of atoms of various elements on both the sides of the equation.*

Step II : Balancing of Chemical Equation

We know from Dalton's atomic theory that atoms can neither be created nor destroyed. Moreover, from law of conservation of mass we know that mass is neither created nor destroyed during a chemical reaction.

Therefore, number of atoms of various elements on both the sides of the chemical equation must be equal.

After writing skeletal equation, the next step is to equalize the number of atoms of various elements on both the sides of the equation by multiplying various formulae by appropriate coefficients. This process is known as **balancing of chemical equations** which is discussed later in this section in details.

In order to equalize the number of atoms of various elements, the various species are multiplied by appropriate numbers. For example, formation of ammonia is represented by the balanced chemical equation given below:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

*An equation in which number of atoms of each element is equal on both the sides of the equation is known as **balanced chemical equation**.*

Sometimes the skeletal equation and the balanced equation may be identical. Some examples are given below:

$$CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$$

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

6.8.2. Balancing of Simple Chemical Equations

The balancing of chemical equations is based upon Dalton's atomic theory and law of conservation of mass. According to Dalton's atomic theory atom is the smallest unit of an element that takes part in chemical reactions and that during chemical reactions atoms are neither created nor destroyed. Therefore, **the number of atoms of each element should remain same before and after the reaction**.

To make the number of atoms of all the elements equal on both the sides in a skeletal equation is known as balancing.

During balancing, the symbols and formulae of various species given in the equation are multiplied with appropriate coefficients. A coefficient is a small whole number, like the coefficients used in algebraic equations. The simple equations can be balanced by **Hit and trial** method. The following steps may be followed while balancing the chemical equations by this method:

Step 1: Write the correct skeleton equation. The skeleton equation contains the formulae of only one molecule of each reactant and product. Once you are satisfied that the skeleton equation is correct, do not change the subscripts in any of the formulae.

Step 2: Start with the compound that has the maximum atoms or the maximum kinds of atoms and the atoms present in it are balanced first.

Step 3: Balance elements that appear only once on each side of the arrow first. Then balance elements that appear more than once on a side.

Step 4: Elementary substances are balanced last of all.

Step 5: If required the whole equation is multiplied by some suitable number in order to make all the coefficients whole numbers.

6.8.2.1. Balancing of Ionic Equations

A molecular equation is balanced on the basis of **mass or material balancing** by equating the number of atoms of various elements on both the sides of the chemical equation. However, in case of ionic equations we have to consider not only mass or material balancing but also **charge balancing**.

For example, consider the following equation:

$$Zn + Ag^{+} \longrightarrow Zn^{2+} + Ag$$

This equation is balanced as far as mass balancing is considered. However, there is + 1 charge on reactant side and + 2 charge on product side. In order to balance charges Ag^{+} on reactant side and Ag on product side are multiplied by 2. Thus, the correct balanced equation is

$$Zn + 2Ag^{+} \longrightarrow Zn^{2+} + 2Ag$$

Thus, **a balanced chemical equation must satisfy mass balance as well as charge balance**. The application of these rules is illustrated in the following solved examples.

Example 6.12: *Aluminium reacts with oxygen on heating to form aluminium oxide. Write balanced equation for the reaction.*

Solution: Aluminium + Oxygen ⟶ Aluminium oxide

The skeleton equation for the reaction is

$$Al + O_2 \longrightarrow Al_2O_3$$

Start with Al_2O_3 because it contains maximum number of atoms and most kinds of atoms. To balance aluminium atoms multiply Al by 2.

$$2Al + O_2 \longrightarrow Al_2O_3$$

In order to balance oxygen atoms, multiply O_2 by $\frac{3}{2}$ so that number of O atoms becomes three on both the sides.

$$2Al + \frac{3}{2}\ O_2 \longrightarrow Al_2O_3$$

In order to make all the coefficients whole numbers, the whole equation is multiplied by 2.

$$2 \times \left[2Al + \frac{3}{2}O_2 \longrightarrow Al_2O_3\right]$$

$$\mathbf{4Al + 3O_2 \longrightarrow 2Al_2O_3}$$

The equation is balanced. There are 4 aluminium atoms and 6 oxygen atoms on each side.

6.9. LIMITING REAGENTS/REACTANTS

In many situations during the chemical reactions, one of the reactants is taken in excess. This is to ensure the completion of reaction. However, on completion of reaction, some of the reactant, taken in excess, is left over. For example, consider the combustion of hydrogen

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

Suppose that 2 moles of H_2 and 2 moles of O_2 are available for reaction. It follows from the equation that only 1 mole of O_2 is required for complete combustion of 2 moles of H_2; 1 mole of O_2 will, therefore, be left over on completion of the reaction. The amount of the product obtained is determined by the amount of the reactant that is completely consumed in the reaction. This reactant is called the *limiting reagent*. Thus, **limiting reagent** may be defined as *the reactant which is completely consumed during the reaction*. The reactant which is not completely consumed is referred to as **excess reactant**.

In the above example, H_2 is the limiting reagent and O_2 is excess reactant. The amount of H_2O formed will, therefore, be determined by the amount of H_2. Since 2 moles of H_2 are taken, it will form 2 moles of H_2O on combustion.

Example 6.13: *How much magnesium sulphide can be obtained from 2.00 g of magnesium and 2.00 g of sulphur by the reaction Mg + S → MgS? Which is the limiting reagent? Calculate the amount of one of the reactants which remains unreacted. [Atomic masses: Mg = 24.3, S = 32.1]*

Solution: First of all each of the masses are expressed in moles :

$$2.00 \text{ g of Mg} = \frac{2.00}{24.3} = 0.0824 \text{ moles of Mg}$$

$$2.00 \text{ g of S} = \frac{2.00}{32.1} = 0.0624 \text{ moles of S}$$

From the equation, Mg + S → MgS, it follows that one mole of Mg reacts with one mole of S. We are given more moles of Mg than of S. Therefore, Mg is in excess and some of it will remain unreacted when the reaction is over. **S is the limiting reagent** and will control the amount of product. From the equation we note that one mole of S gives one mole of MgS, so 0.0624 mole of S will react with 0.0624 mole of Mg to form 0.0624 mole of MgS.

Molar mass of MgS = 56.4 g

∴ Mass of MgS formed

= 0.0624 × 56.4 g = **3.52 g of MgS**

Mole of Mg left unreacted

= 0.0824 – 0.0624 moles of Mg

= 0.0200 moles of Mg

Mass of Mg left unreacted

= moles of Mg × molar mass of Mg

= 0.0200 × 24.3 g of Mg = **0.486 g of Mg.**

EXPERIMENT 6.1

Aim: ***To explore a reaction with a limiting reactant.***

You will need: Zinc metal (powder), hydrochloric acid, 3 flasks and 3 balloons.

Safety: Goggles must be worn and hold the balloons on the test tubes tightly while the reaction takes place.

Procedure:

1. Take three flasks of equal sizes and label them as A, B and C.
2. Weigh the following three amounts of Zn powder: 7.000 grams, 3.270 grams, and 1.310 gram.
3. Take three balloons and label them as 1, 2 and 3. Put the three different masses of zinc powder into the balloons using a small plastic funnel.

   Note: Make sure the zinc powders goes to the bottom of the balloon.
4. Using the graduated cylinder and pipette, accurately measure and transfer 0.100 mole of hydrochloric acid into each of the three flasks.
5. Attach the filled balloons to the mouth of the flasks.

   Note: Make sure that the contents of the balloon and flasks are not mixed.

6. After the balloons are securely attached to the flasks, lift the balloons on the flasks so that the contents of the balloon mix with the contents of the flask.

   Note: Make sure the balloons are held on tightly to the flasks.

   What do you observe?

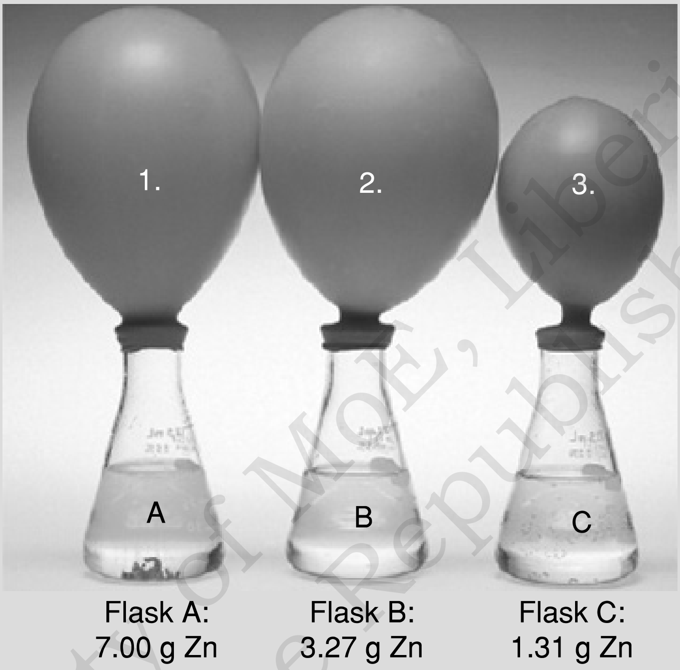


Fig. 6.8. Reaction with limiting reactant

Observation:

When the reactants are combined, the following reaction occurs:

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

The H_2 gas inflates the balloon attached to the flask. The results are as follows:

- **Flask A:** Balloon inflates completely, but some Zn remains when inflation ceases.
- **Flask B:** Balloon inflates completely. No Zn remains.
- **Flask C:** Balloon does not inflate completely. No Zn remains.

Explanation:

In flask A, there is four times the stoichiometric quantity of Zn present, so the balloon inflates to a certain extent because all of the HCl reacts to form H_2 gas; and excess Zn that did not react is visible in the bottom of the flask.

The flask B, contains stoichiometric equivalent quantities of both the reactants, so the balloon inflates to the same extent as the flask A because all of the HCl reacts to form H_2 gas; all of the Zn is used up. It takes longer for the balloon to inflate to the same extent as the first balloon because the reaction slows down considerable as the concentration of Zn and HCl approaches zero towards the end of this reaction.

In flask C, there is one quarter of the stoichiometric quantity of Zn, so the balloon is noticeably smaller than the other two since the Zn is used up before all the HCl is converted to H_2 gas.

GLOSSARY

- **Anion:** A negatively charged ion.
- **Atom:** The smallest particle of an element that takes part in chemical reactions.
- **Cation:** A positively charged ion.
- **Combustion reactions:** A combustion reaction is a reaction of a substance with oxygen and is generally accompanied by rapid release of heat to produce a flame.
- **Decomposition reactions:** The reaction in which a single compound breaks up into two or more simpler substances.
- **Displacement reactions:** The chemical reactions in which one element displaces another element from a compound and takes its place.
- **Endothermic:** A chemical reaction in which heat energy is absorbed.
- **Exothermic:** A chemical reaction in which heat energy is released.
- **Ion:** An atom or group of atoms carrying positive or negative charge.
- **Molar mass:** The mass of one mole, i.e., 6.023×10^{23} particles of that substance.
- **Molar volume:** The volume occupied by one mole molecules of a gaseous substance.
- **Molecule:** The smallest particle of a substance that has independent existence.
- **Oxidation:** It is the process of gain of oxygen or loss of hydrogen.
- **Reduction:** It is the process of loss of oxygen or gain of hydrogen.

SUMMARY

- Law of Conservation of Mass states that "In every chemical reaction, the total mass before and after the reaction remains constant."
- Law of Constant Composition states that "A pure chemical compound always contains same elements combined together in the same proportion by mass."
- Law of multiple proportion states that "When two elements combine with each other to form two or more than two compounds, the masses of one element which combine with fixed mass of the other, bear a simple whole number ratio to one another."
- The molecular formula of a compound may be defined as the formula which gives the actual number of atoms of various elements present in the molecule of the compound.
- The empirical formula of a compound may be defined as the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound.
- The reactions in which the products do not react under any condition to give back reactants are called irreversible reactions.
- The reactions in which products under suitable conditions react to give back reactants are called reversible reactions.
- An equation in which number of atoms of each element is equal on both the sides of the equation is known as balanced chemical equation.
- To make the number of atoms of all the elements equal on both the sides in a skeletal equation is known as balancing.
- Limiting reagent may be defined as the reactant which is completely consumed during the reaction.

EVALUATION

I. Multiple Choice Questions

1. Who proposed law of definite proportions?

(a) John Dalton
(b) Joseph Proust
(c) Robert Boyle
(d) None of these

2. Molar volume of all gases at S.T.P. is:

(a) 24.2 litres (b) 22.4 litres

(c) 22.5 litres (d) None of these

3. Which of the following contains one mole molecules of the substance?

(a) 16 g Oxygen (b) 7 g Nitrogen

(c) 2 g Hydrogen (d) 36 g Water

4. When a single product is produced from two or more reactants, the reaction is

(a) Metathesis reaction (b) Decomposition reaction

(c) Combination reaction (d) Displacement reaction

5. Combination reactions may involve

(a) Combination of two elements

(b) Combination of two compounds

(c) Combination of one element and one compound

(d) All of these

6. Name the reaction in which energy in the form of heat, light and electricity is required to complete the reaction.

(a) Combination (b) Decomposition

(c) Single replacement (d) Double replacement

7. These reactions are also known as Metathesis reaction.

(a) Combustion reaction (b) Decomposition reaction

(c) Combination reaction (d) Double displacement reaction

8. This is the most reactive metal.

(a) Potassium (b) Gold

(c) Aluminium (d) Iron

9. It is the process of gain of oxygen or loss of hydrogen.

(a) Oxidation (b) Reduction

(c) Oxidation-Reduction (d) None of these

10. Which of the following is a characteristic of a reversible reaction?

(a) It never proceeds to completion in a closed container.

(b) It proceeds only in forward direction.

(c) Number of moles of reactants and products are equal.

(d) It can be influenced by a catalyst.

II. State True or False

1. Law of conservation of mass is also known as law of indestructibility of matter.
2. Carbon dioxide contains carbon and oxygen in the ratio of 3 : 5.
3. Atomic radius is measured in nanometres.
4. The balancing of chemical equations is based upon Dalton's atomic theory and law of conservation of mass.
5. In a chemical reaction, the reactant which is completely consumed is called excess reactant.
6. Gold is the least reactive metal.
7. Reduction is the process of loss of oxygen or gain of hydrogen.
8. All combustion reactions are endothermic.
9. In a chemical equation, reactants are written on the right hand side.
10. Cations are also called basic radicals.

III. Answer the Following Questions

1. State law of conservation of mass.
2. On heating, potassium chlorate decomposes to potassium chloride and oxygen. In one experiment 30.0 g of potassium chlorate generates 14.9 g of potassium chloride and 9.6 g of oxygen. What mass of potassium chlorate remain sun decomposed?
3. State law of definite proportion.
4. 1.375 g of cupric oxide was reduced by heating in a current of hydrogen and the weight of copper that remained was 1.098 g. In another experiment, 1.179 g of copper was dissolved in the nitric acid and the resulting copper nitrate converted into cupric oxide by ignition. The weight of cupric oxide formed was 1.476 g. Show that these results illustrate the law of definite proportion.
5. State law of multiple proportion.
6. Carbon and oxygen are known to form two compounds. The carbon content in one of these is 42.9% while in the other it is 27.3%. Show that this data is in agreement with law of multiple proportions.
7. Define an atom.
8. Why is it not possible to see an atom with naked eyes?
9. Define molecule.
10. What is the difference between an atom and a molecule?
11. What is the difference between a molecule of an element and a molecule of a compound?
12. What is an ion?

13. Define mole.

14. Calculate the number of gram-atoms and gram-moles in 25.4 mg of iodine (I_2). Atomic mass of I = 127 u.

15. What is molecular formula? Give an example.

16. What are the steps for writing molecular formula?

17. What is empirical formula? Give an example.

18. What are the steps for writing empirical formula?

19. Find empirical formula for each of the compound having molecular formulae:

(i) H_2O (ii) Na_2CO_3

(iii) N_2O_4 (iv) Fe_2O_3

20. Explain reversible reaction citing an example.

21. Explain irreversible reaction citing an example.

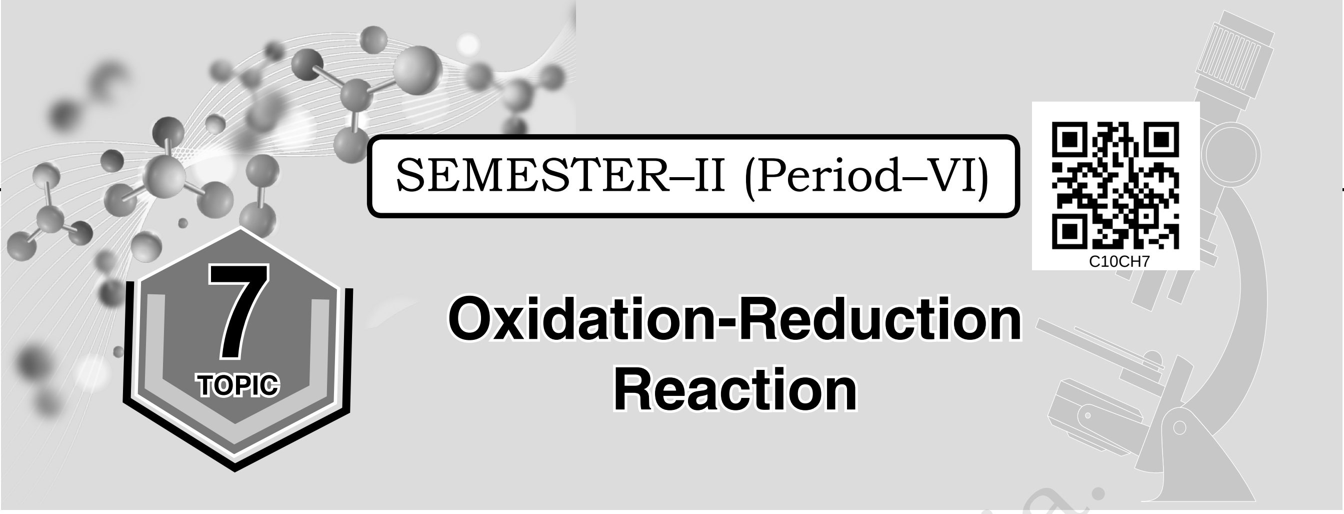
22. What is a chemical equation? Give an example.

23. What are the two steps to write a chemical equation?

24. What do you mean by skeletal equation?

25. What is a balanced chemical equation? Why should the chemical equations be balanced?

26. What do you mean by the term limiting reagents?



SEMESTER–II (Period–VI)

C10CH7

7

TOPIC

Oxidation-Reduction Reaction

Learning Objectives

Upon completion of this topic, learners will be able to:

- Discuss the terms oxidation and reduction
- Discuss the difference between oxidizing and reducing agents
- Apply the concept of calculating oxidation numbers and
- Apply the rules for balancing redox reactions.

Introduction

In our daily life we come across processes like rusting of objects made of iron, fading of the colour of the clothes, burning of the combustible substances such as cooking gas, wood, coal, etc. All such processes fall in the category of specific type of chemical reactions called ***oxidation-reduction reactions*** or ***redox reactions***. A large number of industrial processes like, electroplating, extraction of metals like aluminium and sodium, bleaching of wood pulp, manufacture of caustic soda, etc., are also based upon the redox reactions. Redox reactions also form the basis of electrochemical and electrolytic cells. In order to have proper understanding of redox reactions let us define oxidation and reduction.

7.1. OXIDATION AND REDUCTION

7.1.1. Classical Concept of Oxidation and Reduction

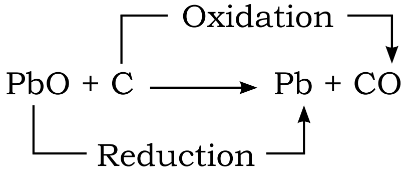
According to classical concept following definitions were proposed to explain the process of oxidation and reduction.

Oxidation *is a process of chemical addition of oxygen or any electronegative radical or removal of hydrogen or any electropositive radical.*

***Reduction** is a process of chemical addition of hydrogen or any electropositive radical or removal of oxygen or any electronegative radical.*

Some examples of oxidation and reduction reactions are given below:

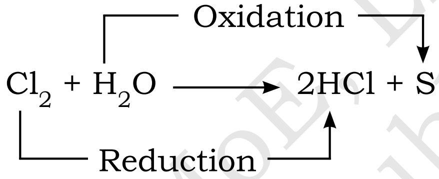
(i) Reaction of PbO and carbon



$$PbO + C \longrightarrow Pb + CO$$

Here, oxygen is being removed from lead oxide (PbO) and is being added to carbon (C). Thus PbO is *reduced* while C is *oxidised.*

(ii) Reaction of H_2S and Cl_2



$$Cl_2 + H_2O \longrightarrow 2HCl + S$$

Here, hydrogen is being removed hydrogen sulphide (H_2S) and is being added to chlorine (Cl_2). Thus H_2S is *oxidised* and Cl_2 is *reduced.*

ACTIVITY 7.1

Study of Oxidation of Copper Powder

- Take about 2 g of copper powder in a China dish and heat (Fig. 7.1)

  What do you observe?

  It is observed that the surface of copper powder becomes black.

  The surface of copper becomes black due to formation of copper (II) oxide by reaction between copper and oxygen present in the air

$$2Cu(s) + O_2(g) \longrightarrow \underset{\text{(Black)}}{\underset{\text{Copper (II) oxide}}{2CuO(s)}}$$

In this reaction copper is oxidised to copper (II) oxide.

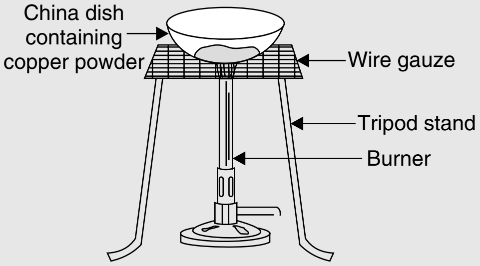


Fig. 7.1. Oxidation of copper to copper oxide

7.1.2. Electronic Concept of Oxidation and Reduction

According to electronic concept, the oxidation and reduction processes can be defined as follows oxidation. ***Oxidation*** *is a process in which an atom or a group of atoms taking part in chemical reaction loses one or more electrons*. The loss of electrons results in the increase of positive charge or decrease of negative charge of the species.

Some examples are as follows.

$$\underset{\text{Chloride ion}}{2Cl^-} \longrightarrow \underset{\text{Chlorine molecule}}{Cl_2 + e^-}$$

(*Decrease of negative charge*)

$$\underset{\text{Iron (II)}}{Fe^{2+}} \longrightarrow \underset{\text{Iron (III)}}{Fe^{3+} + e^-}$$

(*Increase of positive charge*)

$$\underset{\text{Copper}}{Cu} \longrightarrow \underset{\text{Copper (II)}}{Cu^{2+} + 2e^-}$$

(*Increase of positive charge*)

The species which undergo the loss of electrons during the reactions are called reducing agents or reductants. Cl^-, Fe^{2+} and Cu are reducing agents in the above examples.

Reduction *is a process in which an atom or a group of atoms taking part in chemical reaction gains one or more electrons*. The gain of electrons results in the decrease of positive charge or increase of negative charge of the species. Some examples are as follows.

$$Ag^+ + e^- \longrightarrow Ag$$

(*Decrease of positive charge*)

$$\underset{\text{Iron (III)}}{Fe^{2+} + e^-} \longrightarrow \underset{\text{Iron (II)}}{Fe^{2+}}$$

(*Decrease of positive charge*)

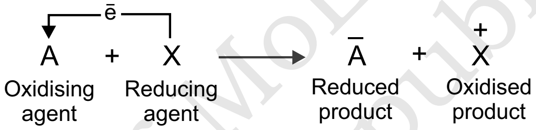
$$\underset{\text{Bromine molecule}}{Br_2 + 2e^-} \longrightarrow \underset{\text{Bromide ions}}{2Br^-}$$

(*Increase of negative charge*)

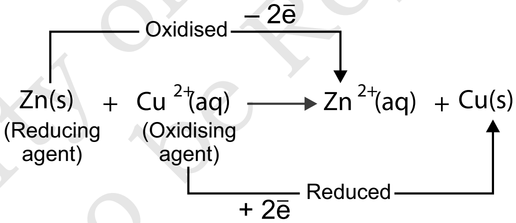
The species which undergo gain of electrons during the reactions are called **oxidising agents** or **oxidants**. In the above examples, Ag^+, Fe^{3+} ions, Br_2 molecule are oxidising agents.

7.1.3. Simultaneous Occurrence of Oxidation and Reduction

Since oxidation involves *loss* of electrons and reduction involves *gain* of electrons, it is evident that if one substance loses electrons, another substance at the same time must gain electrons because electrons cannot be the products in any chemical change. This means that in any process, oxidation can occur only if reduction is also taking place side by side and *vice versa*. Thus, neither oxidation, nor reduction can occur alone. Both the processes are complementary like *give* and *take* and proceed simultaneously. That is why chemical reactions involving reduction-oxidation are called **redox reactions**. In fact, during the redox reaction there is a ***transference of electrons from the reducing agent to the oxidising agent*** as shown below:



For example, consider a reaction between zinc and copper ions.



In this reaction, zinc atoms lose electrons and are oxidised to zinc ions (Zn^{2+}) whereas cupric ions (Cu^{2+}) gain electrons and are reduced to copper atoms. Thus, cupric ions act as *oxidising agent* and zinc atoms act as *reducing agent*. In fact, the oxidising agent gets reduced while reducing agent gets oxidised during the redox reactions.

7.1.4. Some Important Oxidising and Reducing Agents

Some important oxidising and reducing agents alongwith their corresponding reduction/oxidation half reactions are summarized below:

A. Oxidising Agents

1. **Potassium dichromate ($K_2Cr_2O_7$)**

   It is powerful oxidising agent. It acts as oxidising agent in acidic medium and gets reduced to Cr^{3+} ions

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$$

2. **Halogens.** Halogens are very good oxidising agents. They get reduced to halide ions.

$$X_2 + 2e^- \longrightarrow 2X^-$$

3. **Potassium iodate (KIO_3)**

It acts as oxidising agent in acidic medium and gets oxidised to I^- ions

$$IO_3^- + 6H^+ + 6e^- \longrightarrow I^- + 3H_2O$$

4. **Nitric acid (HNO_3)**

Both concentrated and dilute HNO_3 can act as oxidising agents. However, their reduction products are different.

Conc. HNO_3:

$$NO_3^- + 2H^+ + e^- \longrightarrow NO_2 + H_2O$$

Dil. HNO_3:

$$NO_3^- + 4H^+ + 3e^- \longrightarrow NO + 2H_2O$$

5. **Manganese dioxide (MnO_2)**

MnO_2 gets reduced to Mn^{2+} ions in acidic medium.

$$MnO_2 + 4H^+ + 2e^- \longrightarrow Mn^{2+} + 2H_2O$$

B. Reducing Agents

1. **Stannous chloride or Tin (II) chloride ($SnCl_2$)**

Sn increases its oxidation number from + 2 to + 4.

$$Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$$

2. **Hydrogen sulphide (H_2S)**

H_2S gets oxidised to sulphur

$$S^{2-} \longrightarrow S + 2e^-$$

3. **Sulphur dioxide (SO_2)**

In aqueous solutions SO_2 exists as H_2SO_3, which gets oxidised to SO_4^{2-} ions.

$$SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + 2H^+ + 2e^-$$

4. **Oxalic acid ($H_2C_2O_4$)**

Oxalic acid gets oxidised to carbon dioxide. The oxidation number of carbon increases from + 3 to + 4.

$$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-$$

5. **Metal atoms like Zn, Mg, etc.**

These are oxidised to their respective cations.

$$Zn \longrightarrow Zn^{2+} + 2e^-$$

7.1.5. Differences between Oxidising Agent and Reducing Agent

The following table illustrates some differences between oxidising and reducing agents:

Table 7.1. Differences between Oxidising Agent and Reducing Agent

	Oxidising Agent	**Reducing Agent**
Description	It is a process in which an atom or a group of atoms taking part in chemical reaction loses one or more electrons.	It is a process in which an atom or a group of atoms taking part in chemical reaction gains one or more electrons.
Nature	Acts as the electron acceptor	Acts as the electron donor
Examples	Concentrated H_2SO_4, $KMnO_4$, $K_2Cr_2O_7$, O_2, Cl_2, etc.	Concentrated HCl, pure metals, carbon, H_2, SO_2, H_2S, etc.

7.1.6. Testing for Oxidising and Reducing Agents

7.1.6.1. Testing for Presence of Oxidising Agents

ACTIVITY 7.2

Testing for Presence of Oxidising Agent

1. Add a reducing agent, e.g. aqueous potassium iodide (KI) to the oxidising agent.
2. Shake the mixture.

   What do you observe?

   A brown solution of iodine is produced.
3. To identify the presence of iodine, add starch solution to it.

   What do you observe?

   A dark blue coloration is obtained. It confirms the presence of Iodine (oxidising agent) because iodine reacts with starch.

The presence of oxidizing agent can also be detected using any of the following reagents.

(i) Sulphur (IV) oxides, SO_2 with acidified Barium trioxonitrate (V) solution
(ii) Iron (II) Chloride solution ($FeCl_2$)
(iii) Hydrogen sulphide gas (H_2S)

Table 7.2. Summary of Tests

S. No.	Test	Observation	Inference
1.	Oxidising Agent + $FeCl_2(aq)$	Green colour of Fe^{2+} solution turns to reddish-brown of Fe^{3+}	Oxidising agent is present
2.	Oxidising Agent + $H_2S(g)$	Formation of yellow deposits of sulphur	Oxidising agent is present
3.	Oxidising Agent + $SO_2(g)$ + dilute $HNO_3(aq)$ + $Ba(NO_3)_2(aq)$	White precipitate of insoluble $BaSO_4$ is formed	Oxidising agent is present

7.1.6.2. Testing for Presence of Reducing Agents

ACTIVITY 7.3

Testing for Presence of Reducing Agent

1. Add an oxidising agent, e.g. acidified potassium manganate (VII) to the reducing agent
2. Shake the mixture

   What do you observe?

   The acidified potassium manganate (VII) turns colourless on the addition of reducing agent. It confirms the presence of reducing agent.

The presence of reducing agents can also be detected using any of the following reagents.

(i) Acidified potassium dichromate(VI)
(ii) Chlorine (Cl_2)

Table 7.3. Summary of Tests

S. No.	Test	Observation	Inference
1.	Reducing Agent + acidified $K_2Cr_2O_7$	Orange solution of $K_2Cr_2O_7$ turns green on addition of reducing agent.	Reducing agent is present
2.	Reducing Agent + $Cl_2(g)$	Greenish-yellow coloured Cl_2 turns colourless on addition of reducing agent.	Reducing agent is present

Example 7.1: *Consider the combustion of magnesium in oxygen;*

$$2Mg + O_2 \longrightarrow 2MgO$$

Identify:

(i) *the substance undergoing oxidation*
(ii) *the substance undergoing reduction*
(iii) *the substance acting as oxidising agent*
(iv) *the substance acting as reducing agent.*

Solution:

(i) The given reaction may be written as:

$$2Mg + O_2 \longrightarrow 2Mg^{2+}\,O^{2-}$$

In this reaction Mg atom loses its two valence electrons and changes into Mg^{2+}.

$$Mg - 2e^- \longrightarrow Mg^{2+}$$

Thus, *Mg undergoes oxidation.*

(ii) Oxygen atom accepts or gains two electrons and changes to O^{2-}

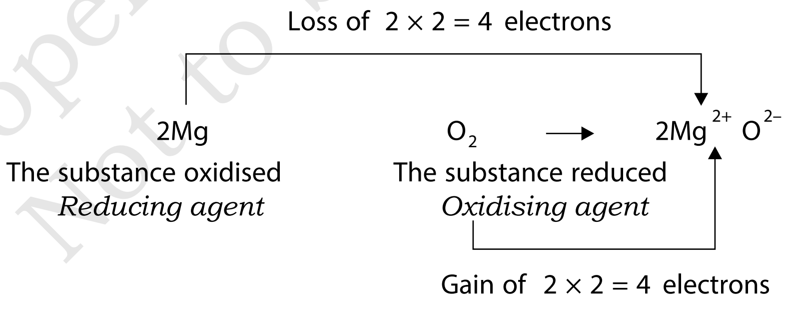
$$O + 2e^- \longrightarrow O^{2-}$$

$$\text{or} \quad O_2 + 4e^- \longrightarrow 2O^{2-}$$

Thus, *O_2 gains electrons and undergoes reduction.*

(iii) The electrons lost by Mg are gained by O_2 which undergoes reduction. Thus, *Mg is the reducing agent.*

(iv) Similarly, O_2 by accepting electrons brings about oxidation of Mg and hence *O_2 acts as oxidising agent.*



7.2. CALCULATING OXIDATION NUMBER

In many covalent reactions such as reaction between H_2 and Cl_2:

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$$

the loss and gain of electrons could not be easily explained. In order to explain transference of electrons in either of the species in a more convenient way, the concept of oxidation number has been introduced.

Oxidation number (O.N.): Oxidation number of the element in a compound is defined *as the residual charge which its atom has or appears to have when all other atoms from the molecule are assumed to be removed as ions by counting the shared electrons with more electronegative atom.*

For example, in hydrogen chloride molecule, chlorine is more electronegative than hydrogen. Therefore, the shared pair is counted towards chlorine atom as shown below:

As a result of this, chlorine gets one extra electron and acquires a unit negative charge. Hence, oxidation number of chlorine is –1. On the other hand, hydrogen atom without electron has a unit positive charge. Hence, oxidation number of hydrogen in hydrogen chloride is +1.

It may be noted that *electrons shared between two similar atoms are divided equally between the sharing atoms*. Hence in molecules like H_2, Cl_2, Br_2 the oxidation number of element is zero.

7.2.1. Rules for Assigning Oxidation Number to an Atom

The followings rules have been formulated on the basis of the assumption that electrons in a covalent bond belong entirely to the more electronegative atom.

1. *The oxidation number of the element in the free or elementary state is always zero irrespective of its allotropic form.* For example,
   Oxidation number of helium in He = 0
   Oxidation number of chlorine in Cl_2 = 0
   Oxidation number of sulphur in S_8 = 0
   Oxidation number of phosphorus in P_4 = 0.
2. *The oxidation number of the element in monoatomic ion is equal to the charge on the ion.* For example, in K^+Cl^-, the oxidation number of K is **+1** while that of Cl is **–1**. In the similar way, oxidation number of all the alkali metals is **+1** while those of alkaline earth metals is **+2** in their compounds.
3. *The oxidation number of fluorine is always **–1** in all its compounds.* Other halogens (Cl, Br and I) also have an oxidation number of **–1**, when they occur as halide ions in their compounds. However, in oxoacids and oxoanions they have positive oxidation numbers.
4. *Hydrogen is assigned oxidation number **+1** in all its compounds except in metal hydrides.* In metal hydrides like NaH, MgH_2, CaH_2, LiH, etc., the oxidation number of hydrogen is **–1**.

5. *Oxygen is assigned oxidation number* **–2** *in most of its compounds, however, in peroxides* (which contain O—O linkage) like H_2O_2, BaO_2, Na_2O_2, etc., its oxidation number is **–1**. Similarly, the exception also occurs in compounds of fluorine and oxygen like OF_2 (F—O—F) and O_2F_2 (F—O—O—F) in which the oxidation number of oxygen is **+2** and **+1** respectively.
6. *In accordance with principle of conservation of charge, the algebraic sum of the oxidation numbers of all the atoms in molecule is* ***zero****. But in case of polyatomic ion the sum of oxidation numbers of all its atoms is equal to the* **charge on the ion**.
7. *In binary compounds of metal and non-metal, the metal atom has positive oxidation number while the non-metal atom has negative oxidation number.* For example, O.N. of K in KI is +1 but O.N. of I is –1.
8. *In binary compounds of non-metals, the more electronegative atom has negative oxidation number, but less electronegative atom has positive oxidation number.* For example, O.N. of Cl in ClF_3 is positive (+3) while that in ICl is negative (–1).

REMEMBER

Oxidation Number of:

- Free elements = 0
- Flourine = –1
- Simple ions = Charge on them
- Oxygen = – 2 ; peroxides (–1); F_2O (+2); F_2O_2 (+1)
- Hydrogen = + 1 ; metal hydrides (–1)
- Sum of O.N. of atoms in molecules = 0
- Sum of O.N. of atoms in polyatomic ions = (Charge on them).

Example 7.2: *Determine oxidation number of carbon in CO_2 and CO_3^{2-}*

Solution:

(i) **C in CO_2:**

Let oxidation number of C be x

$$\text{O.N. of each O atom} = -2$$

$$\text{Sum of O.N. of all atoms} = x + 2(-2) = x - 4$$

As it is neutral molecule, the sum must be equal to zero.

$$\therefore \quad x - 4 = 0 \quad \text{or} \quad x = +4$$

(ii) **C in CO_3^{2-}**

Let O.N. of C be x

$$\begin{aligned} \text{O.N. of each O} &= -2 \\ \text{Sum of O.N. of all atoms} &= \text{Charge on ion} \\ \therefore \quad x + (-2) \times 3 &= -2 \\ \text{or} \quad x &= -2 + 6 = +4 \end{aligned}$$

Example 7.3: *Write the oxidation numbers of Cu, O and using these values calculate oxidation number of S in copper sulphate.*

Solution: In $CuSO_4$, O.N. of Cu = +2 as it exists as Cu^{2+} ion in the salt. O.N. of O is – 2.

Now, if O.N. of S is x, then sum of oxidation numbers of all the atoms = 0

$$\therefore \quad +2 + x + 4(-2) = 0 \text{ or } x = +6$$

Example 7.4: *Calculate the oxidation number of the underlined elements in the following species:*

$$\underline{Cr_2}O_7^{2-}, \underline{Pb_3}O_4, H\,\underline{N}\,O_3$$

Solution:

(i) **Cr in $Cr_2O_7^{2-}$:**

$$\begin{aligned} \text{Let oxidation number of Cr} &= x \\ \text{O.N. of each O atom} &= -2 \\ \text{Sum of O.N. of all atoms} &= 2x + 7(-2) \\ &= 2x - 14 \end{aligned}$$

Sum of O.N. must be equal to the charge on the ion.

$$\begin{aligned} \text{Thus,} \quad 2x - 14 &= -2 \\ x &= -\frac{2+14}{2} = \mathbf{+6}. \end{aligned}$$

(ii) **Pb in Pb_3O_4:**

Let O.N. of Pb be x

$$\begin{aligned} \text{O.N. of each O atom} &= -2 \\ \text{Sum of O.N. of all atoms} &= 3x + 4(-2) \\ &= 3x - 8 \end{aligned}$$

The sum of O.N. must be equal to zero

$$\therefore \quad 3x - 8 = 0 \quad \text{or} \quad x = \frac{8}{3} = \mathbf{+2\frac{2}{3}}.$$

(iii) **N in HNO_3:**

Let O.N. of N be x

O.N. of each O atom $= -2$

O.N. of each H atom $= +1$

Sum of O.N. of various atoms

$$= x + 1 + 3(-2) = x + 1 - 6$$
$$= x - 5$$

In molecule, sum must be equal to zero

$\therefore \quad x - 5 = 0 \quad \text{or} \quad x = \mathbf{+5}.$

ACTIVITY 7.4

1. In small groups, discuss about, redox reactions and oxidation number.
2. Make presentations highlighting the following:
    - Oxidation number.
    - Oxidation and Reduction in terms of electrons as well as in terms of oxidation number.
    - Oxidising and Reducing agents with appropriate examples.

7.3. BALANCING SIMPLE REDOX EQUATIONS

We have learnt that redox reactions involve change in oxidation number of the elements due to the transference of electrons. *In such reactions the number of electrons lost during oxidation must be equal to the number of electrons gained during reduction.* This forms the **basic principle** for balancing the redox equations. We shall use the concept of half reactions here, to balance the redox equations which can appear both as **simple equations and ionic equations**.

7.3.1. Steps Involved in Balancing Redox Equations

The various steps involved in the balancing of redox equations ion-electron method are as follows:

1. Indicate the oxidation number of each atom involved in the reaction. Identify the elements which undergo a change in the oxidation number.
2. Divide the skeleton redox equation into two half reactions; oxidation half and reduction half. In each half reaction equalise the atoms which undergo the change in oxidation number.

3. In order to make up for the difference in oxidation numbers, add electrons to left hand side to right hand side of the arrow in each half reaction.
4. Balance **oxygen** atoms by adding the proper number of H_2O molecules to the side which is falling short of O atoms in each half reactions.
5. This step is meant only for **equations in ionic form ionic equations**. It involves the balancing of **H atoms** in each half reaction as follows:
    - (i) *For acidic medium.* Add proper number of H^+ ions to the side falling short of H atoms.
    - (ii) *For basic medium.* Add proper number of H_2O molecules to the side falling short of H atoms and equal number of OH^- ions to the other side.
6. Equalise the number of electrons lost and gained by multiplying the half reactions with suitable integer. Finally add the two half reactions to get the final equation.

The application of various steps described above has been illustrated as follows by balancing the redox equation representing the reaction between iodine and nitric acid.

$$HNO_3 + I_2 \longrightarrow HIO_3 + NO_2 + H_2O$$

Step 1: Indication of oxidation numbers of each atom.

$$\overset{+1+5-2}{NHO_3} + \overset{0}{I_2} \longrightarrow \overset{+1+5-2}{HIO_3} + \overset{+4\,-2}{NO_2} + \overset{+1\,-2}{H_2O}$$

Thus, only nitrogen and iodine undergo change in oxidation number.

Step 2: Division into two half reactions and balancing the atoms undergoing change in O.N.

$$\overset{0}{I_2} \longrightarrow 2\overset{+5}{NHO_3} \qquad (\textit{Oxidation half reaction})$$

$$H\overset{+5}{NO_3} \longrightarrow \overset{+4}{NO_2} \qquad (\textit{Reduction half reaction})$$

Step 3: Addition of electrons to make up the difference in O.N.

$$\overset{0}{I_2} \longrightarrow 2H\overset{+5}{IO_3} + 10e^-$$

(Each I atom loses $5e^-$ therefore, two iodine atoms would lose $10e^-$)

$$H\overset{+5}{NO_3} \longrightarrow \overset{+4}{NO_2}$$

(Each N atom gains 1 electron).

$$\overset{+5}{\mathrm{HNO_3}} + e^- \longrightarrow \overset{+4}{\mathrm{NO_2}}$$

(Each N atom gains 1 electron).

Step 4: Balancing of O atoms by adding proper number of H_2O molecule to the side falling short of O atoms.

$$\begin{array}{rcl} I_2 + 6H_2O & \longrightarrow & 2HIO_3 + 10e^- \\ HNO_3 + e^- & \longrightarrow & NO_2 + H_2O. \end{array}$$

Step 5: Not required because the equation is not ionic.

Step 6: To equalise the electrons lost and gained, multiply reduction half reaction by 10 and add the two half reactions.

$$\begin{array}{rcl} I_2 + 6H_2O & \longrightarrow & 2HIO_3 + 10e^- \\ [HNO_3 + e^- & \longrightarrow & NO_2 + H_2O] \times 10 \\ \hline I_2 + 10HNO_3 & \longrightarrow & 2HIO_3 + 10NO_2 + 4H_2O \\ \hline \end{array}$$

Let us now understand the balancing of ionic equations in acidic and basic mediums by balancing the following skeleton equations.

(i) $Cr_2O_7^{2-} + Fe^{2+} \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$ (in acidic medium)

(ii) $Cr(OH)_3 + IO_3^- \longrightarrow I^- + CrO_4^{2-}$ (in basic medium)

(i) $(Cr_2O_7)^{2-} + Fe^{2+} \longrightarrow Cr^{3+} + Fe^{3+} + H_2O$

Step 1: Indication of O.N. of each atom.

$$\overset{+6}{Cr}_2\overset{-2}{O}_7)^{2-} + Fe^{2+} \longrightarrow Cr^{3+} + Fe^{3+} + \overset{+1}{H}_2\overset{-2}{O}$$

Thus, Cr in $Cr_2O_7^{2-}$ and Fe change their oxidation numbers.

Step 2: Writing the oxidation and reduction half reactions and equalising the atoms changing their oxidation numbers.

$$\overset{+6}{Cr}_2\overset{-2}{O}_7)^{2-} \longrightarrow 2Cr^{3+}$$ (*Reduction half*)

$$Fe^{+2} \longrightarrow Fe^{3+}$$ (*Oxidation half*)

Step 3: Addition of e– to make up the difference in O.N.

$$\overset{+6}{Cr}_2O_7)^{2-} + 6e^- \longrightarrow 2Cr^{3+}$$

(Each Cr atom gains $3e^-$. Thus, 2Cr atom will gain $6e^-$)

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

Step 4: Balance ‘O’ atoms by adding equal number of H_2O molecules to the side which is deficient in O atoms.

$$Cr_2O_7^{2-} + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$$
$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

Step 5: Balance H atoms by adding H^+ ions to the side which is deficient in H atoms.

$$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$$
$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

Step 6: Multiply oxidation half reaction by 6 to equalise the electrons lost and gained and add the two half reactions.

$$Cr_2O_7^{2-} + 6e^- + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$$
$$[Fe^{2+} \longrightarrow Fe^{3+} + e^-] \times 6$$

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$$

It may be remembered that in balanced equation the number of atoms and also the electrical charges must be equal on both the sides of arrow.

(ii) $Cr(OH)_3 + IO_3^- \longrightarrow I^- + CrO_4^{2-}$ (*in basic medium*)

Step 1: Indication of oxidation number of each element.

$$\overset{+3}{Cr}(\overset{-2}{O}\overset{+1}{H})_3 + (\overset{+5}{I}\overset{-2}{O}_3)^- \longrightarrow I- + (\overset{+6}{Cr}\overset{-2}{O}_4)^{2-}$$

Thus, we find that Cr in $Cr(OH)_3$ and iodine in IO_3^- undergo change in oxidation number.

Step 2: Writing oxidation and reduction half reactions.

$$\overset{+3}{Cr}(OH)_3 \longrightarrow (\overset{+6}{Cr}O_4)^{2-} \quad (\textit{Oxidation half})$$
$$(\overset{+5}{I}O_3)^- \longrightarrow I^- \quad (\textit{Reduction half})$$

Step 3: Addition of e^- to make up the difference in O.N.

$$\overset{+3}{Cr}(OH)_3 \longrightarrow (\overset{+6}{Cr}O_4)^{2-} + 3e^-$$
$$d(\overset{+5}{I}O_3)^- + 6e^- \longrightarrow I^-$$

Step 4: Balance O atoms by adding H_2O molecules to the side deficient in ‘O’ atoms.

$$Cr(OH)_3 + H_2O \longrightarrow CrO_4^{2-} + 3e^-$$
$$IO_3^- + 6e^- \longrightarrow I- + 3H_2O$$

Step 5: Balance H atoms. Since the medium is basic, therefore add proper number of H_2O molecules to the side falling short of H atoms and equal number of OH^- ions to the other side.

$$Cr(OH)_3 + H_2O + 5OH^- \longrightarrow CrO_4^{2-} + 3e^- + 5H_2O$$

$$IO_3^- + 6e^- + 6H_2O \longrightarrow I^- + 3H_2O + 6OH^-$$

Step 6: Equalise the electrons lost and gained by multiplying the oxidation half reaction with 2.

$$\begin{array}{rcl} [Cr(OH)_3 + H_2O + 5OH^- & \longrightarrow & CrO_4^{2-} + 3e^- + 5H_2O] \times 2 \\ IO_3^- + 6e^- + 6H_2O & \longrightarrow & I^- + 3H_2O + 6OH^- \\ \hline 2Cr(OH)_3 + 4OH^- + IO_3^- & \longrightarrow & 2CrO_4^{2-} + 5H_2O + I^- \\ \hline \end{array}$$

Example 7.5: *Write skeleton equation for the following process and balance it by half reaction method. permanganate ion oxidises oxalate ion in acidic medium to form carbon dioxide, water and dipositive manganese ion as product.*

Solution: The skeleton equation for the process is:

$$MnO_4^- + C_2O_4^{2-} + H^+ \longrightarrow Mn^{2+} + CO_2 + H_2O$$

Step 1: Indication of oxidation number

$$(\overset{+7}{Mn}\overset{-2}{O_4})^{2-} + (\overset{+3}{C_2}\overset{-2}{O_4})^{2-} \longrightarrow Mn^{2+} + \overset{+4}{C}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O}$$

Step 2: Writing oxidation and reduction half reactions

$$(\overset{+3}{C_2}O_4)^{2-} \longrightarrow 2\overset{+4}{C}O_2 \quad (\textit{Oxidation half reaction})$$

$$(\overset{+7}{Mn}O_4)^- \longrightarrow Mn^{2+} \quad (\textit{Reduction half reaction})$$

Step 3: Adding electrons to make up the difference in O.N.

$$(\overset{+3}{C_2}O_4)^{2-} \longrightarrow 2\overset{+4}{C}O_2 + 2e^-$$

(Each carbon atom loses one electron)

$$(\overset{+7}{Mn}O_4)^- + 5e^- \longrightarrow Mn^{2+}$$

(Each Mn atom gains $5e^-$)

Step 4: Balancing of O atom by adding H_2O molecules

$$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-$$

$$MnO_4^- + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

Step 5: Balancing of H atoms by adding H^+ ions

$$C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-$$

$$MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$$

Step 6: Multiply the oxidation half reactions by 2 and reduction half reaction by 5 to equalise the electrons lost and gained and add the two half reactions.

$$[C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-] \times 5$$

$$[MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{+2} + 4H_2O] \times 2$$

$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \longrightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$$

GLOSSARY

- **Oxidation Number:** A residual charge which an atom of the element appears to have when other atoms from the molecule are removed as ions by counting the electrons with more electronegative element.
- **Oxidation:** A process involving increase in oxidation number by the loss of electrons.
- **Oxidising Agent:** A substance which involves decrease in oxidation number of one or more of its elements.
- **Reducing Agent:** A substance which involves increase in oxidation number of one or more of its elements.
- **Reduction:** A process involving decrease in oxidation number by gain of electrons.

SUMMARY

- According to classical concept:
  - (i) Oxidation is a process of chemical addition of oxygen or any electronegative radical or removal of hydrogen or any electropositive radical.
  - (ii) Reduction is a process of chemical addition of hydrogen or any electropositive radical or removal of oxygen or any electronegative radical.
- According to electronic concept:
  - (i) Oxidation is a process in which an atom or a group of atoms taking part in chemical reaction loses one or more electrons.

(ii) Reduction is a process in which an atom or a group of atoms taking part in chemical reaction gains one or more electrons.

- Chemical reactions involving reduction-oxidation are called redox reactions.
- The species which undergo gain of electrons during the reactions are called oxidising agents or oxidants. Examples: halogens, potassium permanganate, potassium dichromate, nitic acid, manganese dioxide, etc.
- The species which undergo the loss of electrons during the reactions are called reducing agents or reductants. Examples: hydrogen sulphide, sulphur dioxide, oxalic acid, metals like Zn, Mg, etc.
- Oxidation number of the element in a compound is defined as the residual charge which its atom has or appears to have when all other atoms from the molecule are assumed to be removed as ions by counting the shared electrons with more electronegative atom.
- In a redox reactions the number of electrons lost during oxidation must be equal to the number of electrons gained during reduction.

EVALUATION

I. Multiple Choice Questions

1. The process of oxidation involves

(a) addition of oxygen
(b) addition of hydrogen
(c) removal of oxygen
(d) None of these

2. The process of reduction involves

(a) addition of oxygen
(b) addition of hydrogen
(c) removal of hydrogen
(d) None of these

3. This species undergoes the loss of electron during chemical reaction.

(a) oxidising agent
(b) reducing agent
(c) both (a) & (b)
(d) none of these

4. Which of the following is an oxidising agent?

(a) Nitric acid (HNO_3)
(b) Sulphur dioxide (SO_2)
(c) Hydrogen sulphide (H_2S)
(d) Oxalic acid ($H_2C_2O_4$)

5. In which of the following species oxidation number of C is –3?

(a) C_2H_2
(b) CO_2
(c) HCO_3^-
(d) None of these

6. Oxidation number of free electrons is

(a) Zero (b) +1

(c) –1 (d) –2

II. State True or False

1. Oxidising reagents undergo the gain of electron during chemical reaction.

2. Chemical reactions involving oxidation-reduction are called redox reactions.

3. Manganese dioxide is a reducing agent.

4. In redox reactions, the number of electrons lost during oxidation must be equal to the number of electrons gained during reduction.

5. Oxidising agent acts as electron donor.

6. Concentrated nitric acid is an oxidising agent.

III. Answer the Following Questions

1. Explain the terms: oxidation, reduction in terms of electrons. Give suitable examples.

2. In the formation of a compound AB, atoms of element A lost two electrons each while atoms of B gained two electrons each. Which of the elements A and B isoxidised in the formation of AB?

3. An element Y forms a chloride YCl. In terms of gain and loss of electrons, find out which atom is oxidised and which atom is reduced.

4. Explain the terms oxidising and reducing agents. Give some examples.

5. What are the differences between oxidising and reducing agents?

6. Explain the term oxidation number giving examples.

7. Write the rules for assigning oxidation number to an atom.

8. Assign oxidation number to the underlined elements in each of the following species:

(a) $NaH\underline{S}O_4$ (b) $H_4\underline{P}_2O_7$

(c) $K_2\underline{Mn}O_4$